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Question From the Classroom

By Bob Becker

Q. How do lasers work, and what is so special about laser light?

A. The answers can be found in one very excited group of electrons!

You may recall that an atom's electrons can only exist in very specific, discrete energy levels. When they absorb energy, they can become excited from their ground state up to a higher level. Being unstable there, however, they immediately drop back to a lower level, and when they do, they emit a photon of light.

The energy of this photon depends on the specific electron drop that occurred. For example, in a hydrogen atom, an electron dropping from level 3 to level 2 emits light with a wavelength of precisely 656 nm—a red band in the visible spectrum.

In a similar way, fluorescent lights make indirect use of gaseous mercury atoms whose electrons are excited by electrical current. Because the ground state is more stable, only a small fraction of the mercury atoms are in the excited state at any point in time. When an emitted photon strikes another mercury atom, it will most likely be in the ground state, so it will probably absorb the photon, only to reemit it immediately afterward.

The UV light emitted as the electrons fall back to their ground states is invisible—not exactly what you want in a light bulb. Then how do fluorescent bulbs light up your classroom? Visible light results when ultraviolet light emitted by the mercury strikes the phosphor coating on the inside of the bulb.

Laser devices also involve excited electrons, but there is an important difference. The sample inside the device is being constantly pumped with a steady stream of sufficiently high energy. Under this condition, a population inversion can occur. This means that there are more electrons in the excited state than there are in the ground state

at any point in time. In this high-energy environment, a remarkable chain of events can occur:

1. In an "inverted population" an emitted photon from one atom approaches another atom in which there is an excited electron.



2. The approaching photon stimulates the excited electron of the neighboring atom to drop. The second emitted photon heads off in the same direction—matching the first one crest for crest. The wavelengths match perfectly!



Should this pair of photons happen to approach yet another excited atom, a third photon will join their ranks, and so on. These synchronized photons are known as coherent light, for they do not tend to spread apart the way regular light does as it travels along.

Thus, the laser device simply consists of some medium to be excited (which can vary from a gas mixture to a dye molecule to a ruby crystal) and an "energy pump" pumping fast enough to cause a population inversion in that medium.

But there is one more important feature: a pair of mirrors facing one another on either end of the excited medium. It is important to

Lasers in Space

Look carefully at the science fiction-inspired space battle on the cover of this issue. Sci-fi movies are famous for scenes like this—brilliant "laser" strikes and booming, ear-splitting sound. But, are they accurate? Is this what a space battle would look like? Sound like? Use the information in this column to find at least one thing wrong with our cover scene. You might want to do more research on lasers to find even more flaws. We'll post the answers in our April Teacher's Guide. You can take a look. It's allowed!

www.chemistry.org/education/chemmatters.html

MIKE CIESIELSKI

remember that when the sample of atoms is excited and starts emitting light, the photons are emitted randomly in all different directions. In the laser device, most of these photons are lost as they get absorbed into the sidewalls. A small fraction, however, will just happen to emit their photons precisely perpendicular to one of the two mirrors. This beam of photons will effectively have an infinite path through the medium. Try positioning yourself between two perfectly parallel mirrors, and you'll witness this infinite pathway.

And as the photons bounce back and forth between the two mirrors, they stimulate more and more excited electrons to drop and recruit more and more coherent photons, amplifying the beam with each passage.

This light amplification by the stimulated emission of radiation goes by the familiar acronym "LASER". But this laser beam would be trapped inside the tube, bouncing back and forth forever, if it were not for the fact that one of the two mirrors is only partially reflective, allowing some of the coherent light to escape as a narrow beam. Because this beam does not tend to spread apart, its energy can be focused in ways that regular light cannot. This makes lasers much more powerful—and dangerous—than ordinary light.

From guided missiles to supermarket bar-code scanners, from CD players to fiber-optic phone connections, from tattoo removal to delicate eye surgery, there is no question that our world would be quite different if it were not for lasers. But without question, laser pointers should never be treated as toys. It's very likely your school district has banned them for nonclassroom use.



Question From the Classroom

2

How do lasers work, and what is so special about laser light?

The light source featured in sc-fi space battles also operates in much of our down-to-earth electronic gear. Find out what makes it special.

Student Chemist Gets Plants To Do the Dirty Work

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When soil is contaminated with heavy metals, cleanup is critical. But hold the bulldozers! One student hopes to show that plants can do the dirty work, causing fewer problems in the process.

Green Chemistry—Stopping Pollution Before It Starts

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Can we have our products without hurting the environment? It sounds too good to be true, but chemists and manufacturers are finding that with a little planning, many items can be "benign by design".

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How do you select the right energy source for a lab? Here's your chance to think like a green chemist.

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Chemistry of the Lightbulb—Still a Bright Idea

Around for so long, lightbulbs look ordinary and simple. But do you really know how they work? We'll give you a chance to enlighten yourself before explaining a little of the science behind this remarkable invention.

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Edible Wraps—Safe, Strong, and Delicious

Food chemists are about to give you another lunch-packing decision—the flavor of the wrapper! These plastic-like edible wraps will not only keep the food fresh but also improve nutrition and taste.

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Would you believe that lightbulbs don't give off light—they suck dark? No? Neither do we, but finding a way to disprove this well-argued theory makes an interesting April Fools' Day challenge.

TEACHERS!

FIND YOUR COMPLETE

TEACHER'S GUIDE FOR THIS ISSUE AT

www.chemistry.org/education/chemmatters.html

Think high school students can't make a difference? Think again. Marc Burrell spent his high school years researching some novel solutions for cleaning up our environment.

By Jay Withgott

Perhaps he's no ordinary high school student, but Marc Burrell's prize-winning research on how to get plants to remove toxic lead from contaminated soil shows what a student can accomplish given some smarts, some conviction, and a whole lot of persistence.

When soil becomes contaminated with heavy metals, it's not easy to clean up. Tens of thousands of sites worldwide are contaminated with toxic metals and organic pollutants from manufacturing industries, mining, oil extraction, and military ammunition. When these industries close up and move on, the land they leave behind is often too polluted for other uses.

The standard solution is to dig up tons of soil, piling it all into a hazardous waste dump. This is being done at many sites throughout the country by order of the Environmental Protection Agency (EPA). But bulldozing so much dirt can actually release harmful chemicals into the air, putting nearby residents at risk. It wastes soil and leaves ugly gaping holes. And it can cost huge amounts of money—often \$1–3 million per acre.

So some scientists are studying a better way to get the job done—using plants to do the “dirty work”.

Called phytoextraction, or *phytoremediation*, the approach uses plants (“phyto” means plant) to remediate, or detoxify, contaminated soils. “You use the plants to do the cleanup for you,” says Purdue University professor Peter Goldsbrough, a phytoremediation expert. “You end up with a less-expensive solution, and a greener solution, to the problem.”

One researcher making advances in phytoremediation is Marc Burrell. For his work, Burrell, who is 18 and a 2002 graduate of Nicolet High School in Glendale, WI, won fifth place and a \$25,000 scholarship in the prestigious Intel Science Talent Search last spring. He's also won prizes in other high-profile competitions that have taken him around the country, from Orlando to San Jose to New Orleans, across the Atlantic to London, and more recently, to Japan. He's appeared on “Good Morning America” and has been profiled in *Wired.com*, as well as in local TV and newspapers. Last fall, after sorting through the flood of college admission offers, Marc entered his freshman year at Rice University in Houston, Texas.

Dr. Goldsbrough, one of his mentors, says he was “just delighted” to work with Burrell. “What was really amazing about Marc is that as a high school student, he was very able to access the relevant scientific literature in technical journals and understand most of it without a whole lot of assistance from anyone else — *and*, he was able to pinpoint interesting questions. It was really quite remarkable.” How'd he do it? “He was very persistent, very dogged,” Goldsbrough says. “He really wanted to get an answer.”



ALL PHOTOS COURTESY MARC BURRELL

Student Chemist G To Do the

But Marc says, "I'm a pretty laid-back guy." *Really???* "... Okay, when I'm not in the lab or gone for science competitions," he admits. Nonetheless, between his scientific pursuits, Marc, who likes jazz, sports, and travel, also found time to play varsity basketball and football and to play saxophone in the jazz band.

Getting into some serious science

Marc got seriously interested in science in seventh grade. During his freshman year of high school, he began doing research and competing in science fairs. Enrolled in NASA's Sharp Plus program, he spent the summer of 2000 at Jackson State University in Mississippi. That's where he began work on phytoremediation.

Working under Jackson State professors Greg and Maria Begonia, he tested how well wheat grown in the lab could take up lead from soil under a variety of chemical conditions. Marc found that adding acetic acid and EDTA (ethylene diamine tetraacetic acid) to the soil increased the wheat's ability to draw up lead through its roots. Under this combined chemical treatment, the wheat accumulated lead to the tune of 1% of the plant's dry biomass. Marc found that when treated with EDTA alone, the plants took up only one-third of this amount. Without either EDTA or acetic acid, they took up only a millionth as much.

What was happening? Normally, when lead is in the soil, it's not accessible to plants. That's because it's tied up in compounds such as lead carbonate and lead oxide—neither of which is water-soluble. But some chemicals can bind to lead, making it water-soluble. The plant roots can then take up these water-soluble metal compounds and gradually join to form long polymers.

Organic chemicals capable of grabbing metals (EDTA is one) are called chelating agents or chelators (from the Greek word for "claw"). Adding acid enhances the effect of

the chelating agent, because the acid donates protons. And these hydrogen ions help break the bonds between lead and its former partners, freeing up more of it to bind to EDTA.

Of course, it's not quite so simple, as Marc points out. "Soil chemistry has very complex dynamics," he says, with factors such as temperature, humidity, soil moisture, and nutrients all playing a role. But his findings suggest one avenue by which the efficiency of phytoremediation could be improved.

Looking at plants from the inside

His findings were published in the May 2002 issue of the *Bulletin of Environmental Contamination and Toxicology*. But Marc wanted to better understand what was happening on the molecular level inside plants to enable them to take up toxic metals such as lead. In particular, he wanted to understand how genes and proteins affect lead uptake. "The long-term goal is to bioengineer plants to be able to accumulate lead," Marc says. "And you can bioengineer super-plants only when you understand the systems and their mechanisms."

So he sought out Goldsbrough, who had studied how cadmium and other metals are handled by *Arabidopsis*, a small mustard plant commonly used in laboratory genetic studies. Many plants, when they take up toxic metals, will send forth peptides called *phytochelatins* to bind to the metals. The phytochelatins drag the toxins to vacuoles, or empty cellular sacs. Here, the poisons are stashed away where they won't

harm the plant. Phytochelatins were thought to work with numerous types of metals, but no one had tested their interaction with lead.

So that's what Marc set out to do. Working under Dr. Heather Owen, director of the electron microscopy lab at the University of Wisconsin–Milwaukee, Marc tested a strain of *Arabidopsis* from transgenic seeds Goldsbrough had provided him. In this mutant strain, geneticists had knocked out the gene responsible for producing phytochelatins. Marc found that these mutant plants lacking the ability to produce phytochelatins were more susceptible to lead poisoning than normal plants and died sooner. So apparently,



Growing plants like these wheat plants for extracting metal ions from the soil is relatively new. Environmentalists find the approach interesting because it's cleaner, healthier, and cheaper than bulldozing up vast heaps of contaminated soil.

Gets Plants Dirty Work

phytochelatins are, in fact, functioning to squirrel away lead into vacuoles.

Arabidopsis is a great model organism for lab studies, but Marc wanted to work with something more applied, so he contacted Dr. Hillel Fromm of the University of Leeds in the United Kingdom, who works with tobacco plants. Remaining in Owen's lab, Marc obtained some of Fromm's mutant tobacco plants that overexpress the protein for producing phytochelatins, enabling the plants to take in more metals. Marc found that presenting the plants with lead induced them to produce even more phytochelatins, increasing the amount they could accumulate.

Marc Burrell entered Rice University in Houston in the fall of 2002. Selected as a member of the university Century Scholars Program, he was invited to join a bioengineering research team headed by Dr. Antonios G. Mikos (www.rufrice.edu/~mikosgrp/). Marc describes the work at Rice as "challenging and rewarding". He adds, "I'm still undecided about majors, but possibly bioengineering."

One of the highlights of his 2002 fall semester was a trip to Japan sponsored by Army Research Laboratories (ARL). As part of his prize for being one of the finalists in the May 2003 Intel International Science and Engineering Fair in Louisville, KY, he participated in the two-week Operation Cherry Blossom Program in Tokyo. While in Japan, Marc's group toured several ARL labs, visited the U.S. embassy, and toured cultural sites.

Before returning to the United States for the holidays with his family in Milwaukee, Marc was honored to represent the United States at the Japanese Student Science Awards Program, a ceremony that included his introduction to Prince and Princess Akishino of the Imperial Family—a tradition of the program.

Finally, Marc assisted Owen with her work on phytoferritins, proteins that help regulate the amount of iron in a plant's system. Plants need iron for proper photosynthesis, but too much of it is harmful. Any excess must be sequestered away into vacuoles. Marc wondered how plants differentiate between toxic metals like lead and helpful metals like iron—and whether the phytoferritins could also help sequester lead.

Marc and Dr. Owen's experiments showed that they could. Making iron scarce at

first and then resupplying it in excess to the plants caused the phytoferritins to take up both iron and lead. Owen says there's still work to be done to confirm the findings, but Marc's experiments gave them hope that they'd pinpointed phytoferritin's uptake of lead.

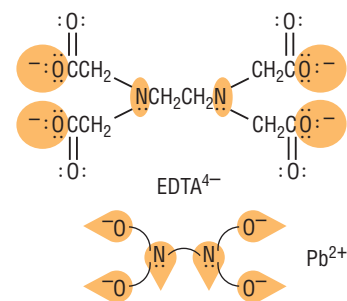
"A lot of my colleagues told me I was nuts" to take on a high-school student, she laughs. But the long hours she worked with him paid off, she says, with some valuable research results.



Can plants get the job done?

Phytoextraction procedures are still relatively new and undeveloped, but the approach is attractive because it's cleaner, healthier, and cheaper than bulldozing up vast heaps of contaminated soil. If the plants can concentrate the metals enough, they can be put through a smelting procedure to recover the metals. If they can't, then the plants are dried and disposed of in a hazardous waste site. The metal does not disappear this way, but its presence in the environment is more safely confined.

Phytoremediation faces some big hurdles, though. For one, it takes a long time. Plants can only take up so much. When one crop has done its job, it is harvested and another must be planted. It may take 5–20 years of repeated plantings before the soil's metal content is reduced to an acceptable level. Furthermore, the cleanup is limited to the depth of soil that the plants' roots will reach. And the metals need to be in a form that plants can take up (remember the need for water solubility discussed above). Finally,



EDTA is a chelating agent. Like a chemical "claw", it wraps around a metal ion to form multiple coordinate bonds. The result is a very stable water-soluble compound that is readily absorbed by plant roots.

plants that accumulate toxins can potentially harm insects that eat them. Birds and other animals eat those insects, passing toxins along the food chain.

Such obstacles help explain why real-world applications so far have been limited. But there are plenty of examples even now. For instance, the U.S. Army Corps of Engineers is using vegetation in artificial wetlands to clean up groundwater contamination from ammunition at military bases in Iowa, Tennessee, and Nebraska.

Edenspace, a private company based in Virginia, has used plants to extract lead from residential sites, to remove arsenic from military and energy facilities, to clean up zinc and cadmium at EPA Superfund sites, and to reclaim tungsten from abandoned mines. It has even used them to extract radioactive uranium, strontium, and cesium from U.S. military sites and the infamous Chernobyl nuclear reactor in the Ukraine. Another company, Ecolotree, in Iowa, has used poplar trees, legumes, and grasses at 55 landfills, wastewater treatment sites, agrochemical spill areas, and other locations in the United States and Europe.

Such promising efforts are at the forefront of modern environmental chemistry. Will these efforts be effective enough to replace the bulldozers when it comes to soil cleanup? Marc Burrell and other bright hard-working researchers are confident the answer is yes. ▲

Jay Withgott is a science writer and journalist based in San Francisco, CA. His most recent *ChemMatters* article, "Chemical Profiling—Tracking Down the Source", appeared in April 2002.

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Green Chemistry

STOPPING POLLUTION BEFORE IT STARTS

You've invited your friends over for videos and a half gallon of ice cream. Just for laughs, somebody reads the label and says, "Get this! Here's what we're eating!" They rattle off something that sounds like "dihixi-hexegazorniplatz hydroxy-megadogdoo". You make a mental note to buy something natural and chemical-free next time. Just to be safe.

Ha! Let's try an experiment. Put the magazine down and make a list of three things that

are free of chemicals. Easy? Take your time. We'll wait. So what did you come up with?



MIKE DIESELSKI

Organic herbal tea? Nope, lots of fragrant hydrocarbons there. Filtered water? Think again. It abounds in trace minerals along with plain old H₂O. Air? Need we mention oxygen, nitrogen, and carbon dioxide gases to name a few? Sure, you could be clever and name a vacuum, but that's more the absence of a thing.

The undeniable truth is this. We live in a chemical world. Any effort to rid ourselves of chemicals is futile. So, as we prepare to celebrate Earth Day 2003 this month, why are chemists and the chemical industries still looking like the bad guys?

Try this list. List all of the manufactured products on which your life and lifestyle depend—like antibiotics, processed foods, electricity at the touch of a button, silicon chips inside computers, perhaps even the latest fashion fad. The list is long, and few of us would want to revert to harsh frontier life, with most of our time tied up with chopping wood, hauling water, and scrubbing clothes on a washboard. But there is a nagging question: Are these advances coming at too high a price for the environment? Can we have our products and still sustain a healthy livable planet?

You're probably already bracing for the message. This is going to be all about smokestacks and acid rain. You've heard all you want to hear about that. Bad industries, bad polluters, bad consumers, bad vehicles—bad, bad, bad. Or maybe you think that because this is a chemistry magazine, we're going to defend all industrial practices, gloss over the obvious problems, and paint a rosy picture of a utopian world in which chemistry solves all problems while creating none.

Let's try a different approach. Common sense. Although your school chemistry lab is not a big manufacturing plant, you and your teacher make important process decisions every time you do an activity. And many of those decisions are a lot like those made by industrial chemists: What chemicals do we need? (Are they safe? Expensive? Can we use less and still get good results?) What solvent should we use? (Again, is it safe? Will water work as well?) How shall we heat the reaction vessel? (Will it go just as well at room temperature if we wait? Hot plate? Bunsen Burner? Microwave?) And what should we do with the wastes that accumulate? (Down the drain? In the trash can?) You're skeptical. Surely that little bit of chemistry doesn't make any difference.

continued on page 10



DIGITAL VISION

GREEN ENERGY—

When green chemists design the chemical pathway for making a product, energy use is high on their list of concerns. They look for alternative ways to make the same product, always considering whether the product would be more “benign” if made by a method requiring less energy. With many factors to consider—starting materials, solvents, and the like—the decision is seldom clear-cut.

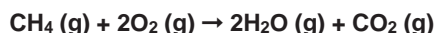
Once they’ve settled on the right chemistry, they think about finding the most efficient way to deliver the energy to the reaction. Does one

way conserve precious energy dollars and fuel resources better than other ways?

You and your teacher face similar choices when you conduct your activities in the lab. You probably decide whether to bring the reactants up to the required temperature by heating with a Bunsen burner, electric hot plate, or microwave oven. Have you asked which one is most efficient in terms of energy cost? How can you find out? We’ll get you started on the problem, and then we’ll leave it up to you to find the answer for the equipment in your lab.

The Bunsen burner

The typical laboratory burner requires natural gas—largely methane (CH_4)—that is piped in from a commercial supplier. Upon ignition, methane mixes with air in the burner to produce a flame. The complete combustion of methane in air is represented by the following equation:



This reaction is very exothermic, and the thorough mixing of the natural gas with the air by the Bunsen burner helps ensure that the reaction is relatively complete.

Of course, not all of the heat released by this reaction is actually absorbed by the materials to be heated. Some portion of the heat goes into heating up the containers or the immediate surroundings. The efficiency of the process is calculated as follows.

$$\% \text{ Efficiency} = \frac{\text{Heat absorbed by material}}{\text{Heat given off by burning natural gas}} \times 100$$

In this activity, we will measure how efficient it is to heat a sample of water using a Bunsen burner.

Be sure to wear safety goggles when operating the Bunsen burner. Conduct this activity only under teacher supervision.

- Using the balanced equation, calculate the amount of energy (ΔH_f) for the combustion of methane. Use a handbook or textbook to look up the standard heats of formation. (Assume STP).



$$\Delta H_r = (\sum \Delta H_f \text{ products}) - (\sum \Delta H_f \text{ reactants})$$

- Next, you need to know how *much* natural gas is delivered from the connection on your desktop over time. Suppose you had the following apparatus: **1 meter of hose, a tub or bucket for holding water, a stopwatch or clock with second hand, a 2-L soda bottle, and a variety of volumetric measuring containers such as large graduated cylinders.** How would you measure the amount of methane delivered every second? Since you’ll need to open the gas valve on your desk and the gas control valve on your Bunsen burner to the optimum settings for producing a flame of reasonable size, you’ll need to begin with some burner tests.

- Now, design a plan for collecting and measuring gas, paying particular attention to safety and how you will dispose of the natural gas you are measuring. If you are actually going to carry out the plan, **be sure you have your teacher’s approval and careful supervision before proceeding. It is very important that no open flames be present during this phase of the procedure.** After you have finished measuring, record your results for use in the completion of this activity. Or, obtain the figure from your teacher if this measurement is already available.
- After all collected methane has been safely disposed, set up a **250-mL beaker of water containing 200 mL of water on a ring stand** or support suitable for heating with the Bunsen burner. Using a **thermometer**, measure the initial temperature of the water. Then, light the burner and begin to heat the water. Start tracking the time.
- Heat the water until the temperature increases about 30–50 °C over the starting temperature. Measure the final temperature to the nearest 0.1 °C. Note the time when your heating has finished.
- Calculate the amount of heat absorbed by the water, the amount of heat released by combustion, and the percent efficiency of the heating process. Use the following equations in your calculations. (Assume STP.)

IT'S YOUR DECISION

$$\text{Heat absorbed by the water} = \text{mass of water (g)} \times \Delta T \times 4.18 \frac{\text{J}}{\text{g} \times ^\circ\text{C}}$$

$$\text{Heat released by combustion} = \frac{\text{L (gas)}}{\text{sec}} \times \frac{\text{mol (gas)}}{22.4 \text{ L (gas)}} \times \frac{\text{J}}{\text{mol (gas)}} \times \text{sec}$$

$$\% \text{ Efficiency} = \frac{\text{Heat absorbed by material}}{\text{Heat given off by combustion of natural gas}} \times 100$$

The electric hot plate

Another typical means of heating in the laboratory is the use of an electric hot plate. Design an experiment, similar to the one described on page 8, to determine the efficiency of an electric hot plate in heating.

1. The energy released by the hot plate depends on its energy rating. Look at the bottom or sides of the hot plate and look for its power rating, in watts.



2. Because a watt = 1J/sec, we can calculate the total amount of energy released from the hot plate using the equation below.

$$\text{Energy released by a hot plate} = \text{total wattage of hot plate} \times \text{time} \times \frac{1\text{J/sec}}{1 \text{ watt}}$$

The microwave oven

Although microwave ovens have been more typically used in the home, they are becoming more widely used in scientific laboratories. Use the experience and information from the previous two activities to determine the efficiency of heating a sample of water using a microwave.

1. The energy released by the microwave depends on its energy rating. Look at the bottom or sides of the microwave and look for its power rating, in watts.
2. Because a watt = 1J/sec, we can calculate the total amount of energy released from



the microwave using the equation below.

3. Using a procedure similar to that described on page 8, heat a sample of water and use the results to calculate the efficiency of the microwave in heating the water.

$$\text{Energy released by a microwave} = \text{total wattage of microwave} \times \text{time} \times \frac{1\text{J/sec}}{1 \text{ watt}}$$

Discussion

You now have what you need to compare the efficiency of the three ways to supply heat to a reaction in your lab. But is that all you need to know? What if you wanted to think about overall costs—and we're talking about more than just dollars. For example, it does little good to choose electricity over natural gas if the power plant that generates the electricity contributes significantly more pollution than does the natural gas producer.

Here are some things to consider before making a final decision.

1. Use your household electricity and natural gas bills to calculate the cost in dollars for each of the experimental trials you did above. Calculate the cost for heating a sample of water by 10 °C for each method. If natural gas is not available in your community, use the national average price or

substitute the cost of propane. Although the cost in dollars is not a perfect reflection of the total cost to the environment, it does reflect how difficult it is to bring the energy source to market. Which energy source was most expensive? Which was least expensive?

2. There are many ways to generate electricity and produce natural gas. Some require less energy to produce than others. Using the Web and other sources, investigate which power source tends to require less energy and contributes less total pollution to your local environment.

Finally, using the information you have collected regarding efficiency, dollar cost, and environmental cost, make a recommendation for how best to minimize the energy used to heat substances in your school lab.

This activity is adapted from Ryan, M. A., Tinnensand, M., Eds. *Introduction to Green Chemistry: Instructional Activities for Introductory Chemistry*; American Chemical Society: Washington, DC, 2002; pp 45–53.

continued from page 7

Now scale those decisions up to manufacturing plant level. History shows that as our nation's industries grew, manufacturing processes were largely about one thing—product. Why not? Manufacturers could tap vast supplies of starting materials to be chopped, pumped, or mined. Exhausting one vein or field, they found others just waiting to be developed. There were streams and wide-open spaces to sweep away the wastes, and the public eagerly awaited the finished products. Few, except for those living downwind or downstream from the plant, complained, and even fewer worried about whether we could have our products and a healthy planet, too.

Stopping pollution before it happens

Since it began on April 22, 1970, Earth Day has served as a day for checking the vital signs of our natural environment. For the last few decades, local Earth Day observances have been about community recycling and cleanups.

This April, although many organizations are planning more clean-up efforts, some imaginative chemists are working on a better plan. What if, they ask, we could manufacture products in such a way that we wouldn't need to haul away waste products? What if we could make products that didn't deplete the planet of precious resources, didn't take a lot of energy or dangerous solvents to make, and didn't leave unwanted byproducts behind? And what if the products themselves would degrade harmlessly when we were finished with them?

Scientists and manufacturers working toward these goals call their movement Green Chemistry. In short, their aim is to prevent pollution before it happens.

Economizing on atoms

Like all good ideas for making new and better products, the so-called green products must pass an important test. Economists call it “the bottom line”. Not only must these green products do their jobs as well as traditional products, they must be profitable for companies and comparably priced for the consumers. If they cannot meet these two tests they probably have little chance for success.

One important principle of green manufacturing is something called “atom economy”—a simple calculation of how many of the atoms found in the starting materials actually end up in the final product. Think about it. If the atoms aren't going into the product, then where are they going? Sometimes they're going up smokestacks or into groundwater! And, if you're the manufacturer, you are not only risking fines and costly cleanups, you're also throwing away valuable atoms that you've just paid for in your starting materials.

Green chemists work hard to build in manufacturing techniques with high atom economy whenever creating new products and processes. When a pharmaceutical manufacturer designed a new process for making ibuprofen, the pain reliever found in over-the-counter medications like Advil and Motrin, they nearly doubled the atom economy from 40 to 77%. The new process decreased waste and increased profits for the German pharmaceutical company Badische Anilin & Soda Fabrik (BASF), the company that developed the new synthesis.

Getting off to a safe start

In some manufacturing processes, not only the waste products but also the starting material can pose hazards. Risky exposures to danger-

ous chemicals are routinely controlled by supplying workers with protective equipment like gloves, respirators, and fume hoods. Green chemists argue that many of these exposure risks can be eliminated by simply choosing safer starting materials.

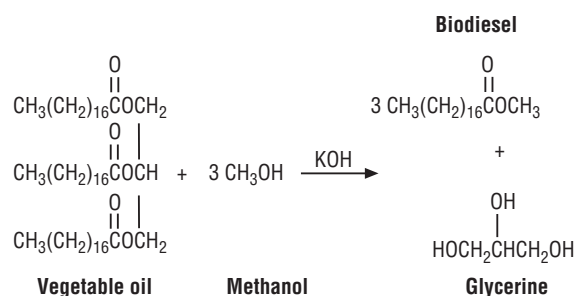
Large quantities of adipic acid ($\text{HOOC}(\text{CH}_2)_4\text{COOH}$) are needed every year for the industrial production of nylon, polyurethane, lubricants, and plasticizers. The typical starting material for making adipic acid is benzene, a known cancer-causing agent. Recently, Karen M. Draths and John W. Frost, chemists at Michigan State University, have developed a greener synthesis of this acid using a starting material that is far less hazardous than benzene, and one that is almost infinitely available in nature—glucose. Using an enzyme found in genetically altered bacteria, glucose can be converted to adipic acid without exposing workers or the environment to hazardous substances.

Using renewable resources that never run out

Green chemistry also means using renewable resources as starting materials whenever possible. Burning a fuel made from readily replaceable sources would be preferable to burning more of our ever-dwindling supply of fossil fuels. For example, many vehicles in the United States operate on diesel fuel. One promising possibility is the manufacture of *biodiesel*. As the name suggests, biodiesel is diesel fuel made from renewable resources like oils derived from farm crops such as soybeans. It is synthesized by removing glycerine—useful for making soap—from soybean or other vegetable oil. Biodiesel can even be made from recycled vegetable oils—like those left over from making fries at fast food restaurants. In the process, a potential waste product is converted to a valuable fuel. In fact, burning biodiesel smells like French fries!

The advantages are clear. Biodiesel, unlike fossil fuel-derived petroleum diesel, is a renewable source of energy. Unlike burning petroleum diesel, burning biodiesel does not emit sulfur, nor does it increase the overall amount of carbon dioxide (CO_2) in the atmosphere.

Does that mean that biodiesel doesn't release any CO_2 when it burns? No, it just means that the CO_2 released is balanced by the carbon dioxide the plants removed from the atmosphere as they grew.



Making green decisions

You may not be a green chemist, but your part is important. Green chemists hope you'll become informed and then choose those products that are, in their words, “benign by design”. 🌱

Michele La Merrill, Kathryn Parent, and Mary Kirchhoff of the Green Chemistry Institute at the American Chemical Society and the editors of *ChemMatters* contributed to this article.
www.chemistry.org/education/greenchem



Chemistry of the Lightbulb

Still a Bright Idea

By Brian Rohrig

With the flick of a switch, we are instantly bathed in visible light. When the power goes out and we scramble for candles and a flashlight, we realize how much we rely on the electric lightbulb.

Few inventions have changed our lives as much as this remarkably simple, yet ingenious device.

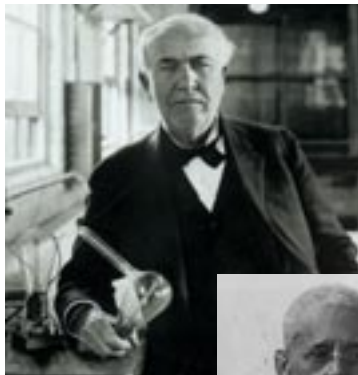
The fact is, we rely on the electric lightbulb. Few inventions have changed our lives as much as this deceptively simple, yet ingenious device.

With the flick of a switch, we are instantly bathed in visible light. When the power goes out, we scramble for candles and matches. But the charm of glowing firelight soon wears off, especially when sports events or homework are in the plans.

Edison's invention

Although Thomas Edison was not the first person to patent an electric lightbulb, he made so many improvements on its design that history generally gives him the credit. Without Edison's improvements, lightbulbs would last about as long as candles. In 1879, he constructed a lightbulb that glowed continuously

for 40 hours, but he was determined to do better than that. By the end of 1880, his 1500-hour lightbulb was ready for public sale. A mere 25 years later, his electric bulb had forever transformed human life by illuminating homes and cities all over the world.



Thomas Edison

© NATIONAL PARK SERVICE



Lewis Howard Latimer

© NEW YORK PUBLIC LIBRARY

These first lightbulbs were *incandescent*, as are most of the lightbulbs in your home today. Incandescence is the property of giving off visible light when heated. The hotter the object, the more energetic the light that is given off.

As you have probably observed, a piece of metal glows if its temperature gets high enough. As the temperature increases, the color of the emitted light changes from dull red, to orange, and at about 5800 °C, to white. A typical incandescent lightbulb operates at a temperature of about 2500 °C, where it glows with a yellow-white light.

Finding just the right filament was not an easy task. Using a simple trial-and-error approach, Edison experimented with thousands of different types of filaments. He was looking for one that would be both long-lasting and affordable.



ALL PHOTOS MIKE CIESIELSKI

Although we talk about glowing lightbulbs, the actual glow comes from only one bulb component—a very thin wire called the filament. A typical 60-watt bulb contains about 2 meters of very thin tungsten wire only about 25 micrometers (1/1000 inch) thick. Look very closely at a clear unlit bulb, and you'll see that the filament is tightly wound into a double coil.

The first breakthrough came when Lewis Howard Latimer, the only African-American member of the Edison research team, developed an improved carbon-based filament that yielded an extended glow. He went on to design efficient production methods for manufacturing them in commercial quantities at reasonable cost.

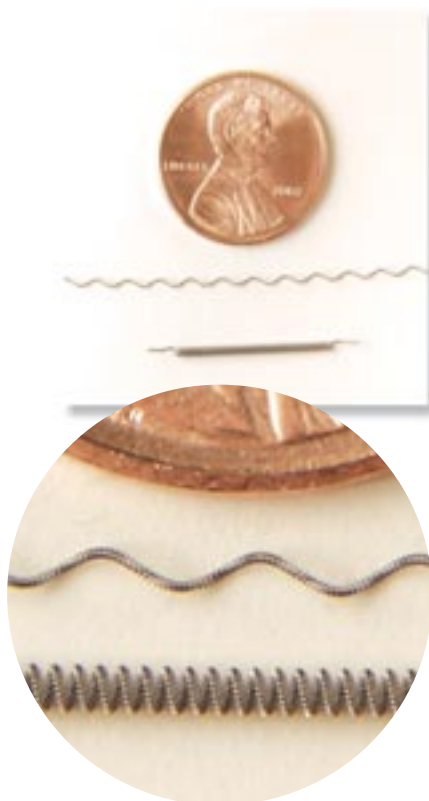
Cost was a significant consideration for the lightbulb to catch on with the public. For example, Edison's team experimented with long-lasting platinum but rejected it for its price tag.

In 1910, William Coolidge of the General Electric Company developed a tungsten filament that is still in use today. Tungsten metal, with a melting point of about 3680 °C, proved to be both affordable and long-lasting.

How does a lightbulb light?

When electric current follows the metal pathway through an incandescent lightbulb, a tremendous amount of resistance is encountered as the electrons make their way through the very thin wire filament. Similar to students in a crowded lunch line, closely packed electrons generate friction—and that, in turn, produces heat. The thinner the wire, the more resistance the electrons encounter as they attempt to pass through. More resistance generates more heat until the wire reaches a high enough temperature to produce visible light.

Lightbulbs “burn out” when the filament breaks. At high temperatures, the tungsten metal, like other metals, *sublimes*. Sublimation is a process by which a solid is converted directly into a gas without first passing through a liquid phase—like old trays of ice cubes shrinking in your freezer.



In lightbulbs, the sublimated tungsten atoms are still sealed inside with nowhere else to go. Examine a burned-out lightbulb and you'll see a tungsten deposit on the interior surface of the glass—clearly visible as a black spot on top of the bulb. Even with its low sublimation rate at high temperatures, the tungsten filament finally wears thin and breaks. As a result, the circuit is broken. When the metal pathway breaks, electrons stop flowing, and the lightbulb ceases to glow.

Finding the right filament was only one hurdle for the early inventors. The second big problem came from the surrounding air—specifically oxygen.

We talk about “burning” lightbulbs, but the fact is that the filaments inside successful bulbs glow *without* burning. Edison found that for many filaments, the presence of oxygen caused the hot material to rapidly combust and break. Since glowing doesn't require any oxygen, Edison's early lightbulbs consisted of filaments mounted in a vacuum.

Although the vacuum solved one problem, it created another. Without any gas exerting pressure on the filament, the rate of sublimation increased. The tungsten atoms of the solid filament readily entered the gas phase at the high temperatures within the glowing bulb.

Argon to the rescue

With the discovery of the noble gas argon in 1894, it became possible to lengthen the life of a lightbulb by filling it with this very unreactive gas. Not only did argon offer an oxygen-free filler, it also controlled the sublimation rate of the filament by transferring some of the excess heat away from the glowing metal.

The transfer process is called convection. As atoms of argon bump into the hot filament, some of the kinetic energy of the tungsten atoms is transferred to the argon atoms. This transfer cools the filament and heats the argon. The argon atoms then speed off to collide with the inner wall of the glass bulb. Upon impact, the argon atoms transfer some of their kinetic energy to the glass, raising its temperature.

In the process, glowing incandescent light bulbs become blistering hot. In fact, about 90% of the electrical energy consumed by an incandescent lightbulb is dissipated as heat.

Krypton would be a better noble gas to use in a lightbulb than argon, since it is a poorer conductor of heat. But krypton is very expensive. The only place to get krypton is from the atmosphere, where its concentration is only about 1 part per million (ppm). By contrast, argon comprises nearly 1% of the atmosphere, making it 10,000 times more abundant than krypton.



Krypton is usually used to fill small flashlight bulbs. Because argon-filled bulbs readily transfer energy away from the glowing filament to the glass, batteries drain rapidly in the process. But when poorer-heat-conducting krypton fills the space, the bulb feels cool to the touch even after extended use, and batteries last longer.

Extended-life bulb

An extended-life lightbulb contains a much longer filament. As a result, there is more surface area to dissipate the heat. The filament does not get as hot and does not sublime as quickly.

The drawback is that the light produced by the extended-life bulb is dimmer and redder. It takes a higher wattage to produce the

same amount of light as a typical incandescent bulb. Wattage refers to how much energy is used per second. So, even though they last longer, extended-life bulbs may not really be a bargain. They actually cost more money due to the greater wattage required to operate them.

Heat lamp bulb

A heat lamp is essentially an incandescent bulb with a very long filament. The filament has a long life and glows a dull red color. The cooler filament emits most of its energy as longer-wavelength infrared light, which cannot be seen but will still heat up objects on which it falls. The globe of the bulb is very large; this increased surface area allows it to radiate more heat. This design makes the heat lamp very practical, since its intended purpose is to give off heat, not light.

MIKE CIESIELSKI



Fluorescent lamps

Since all incandescent bulbs give off a great deal of heat, cool fluorescent lamps offer a much more efficient alternative. Introduced in the 1950s, they soon became widely accepted for nearly all schools, offices, and commercial buildings.

Fluorescent lamps consist of a sealed glass tube containing a mixture of noble gases and a few drops of mercury that vaporize within the tube. When an electric current passes through the gas in the tube, some of the electrons of mercury become excited. Excitation occurs when electrons absorb energy and temporarily achieve a higher energy level. As they return to ground state, the energy previously absorbed by the electrons is primarily released as ultraviolet (UV) light, a light with shorter wavelength and greater energy per photon than visible light.

Because UV light is invisible to humans, the fluorescent lamp must convert it into visible light. This is accomplished by the white phosphor coating on the inside of the bulb. When UV light strikes this phosphor coating, it is converted into visible light. Thus, unlike incandescent bulbs, fluorescent lamps do not give off light by heating any of their components. That makes them much more energy-efficient.

If you have a UV-sanitizing cabinet for

goggles in your classroom, examine the bulb inside. It will resemble a typical fluorescent bulb, but with one big difference. It's transparent. The lack of a phosphor coating means that the bulb emits UV light, with very little visible light. It is this UV light that kills microorganisms and sterilizes your goggles.

Halogen bulb

Halogen bulbs, another type of incandescent bulb, produce intense white light. They are commonly used in car headlights, floodlights, and other applications where very bright light is needed. Halogen bulbs, as their name implies, contain the vapor of a halogen (group 17 on the periodic table), usually bromine or iodine. The halogen molecules act as chemical scavengers, picking up stray tungsten atoms that have sublimed and depositing them back on the filament. The unique ability of the halogen atoms to combine with tungsten atoms means you're not likely to find black spots on the inside of the bulb. However, the filament of the bulb eventually breaks due to uneven deposition of tungsten atoms on the filament.

Since the filament lasts much longer, halogen bulbs are designed to glow several hundred degrees hotter than a typical incan-

TRY THIS!

Can you assemble these three familiar items correctly so that the light bulb lights? You'll need a 12-inch piece of insulated wire with exposed metal at each end, a C or D-cell power source, and a flashlight bulb.

If you succeed on the first try, you are a shining example to us all! A well-known video distributed to educators by the Private Universe Project in 1989 shows recent Harvard and Massachusetts Institute of Technology (MIT) grads struggling with the challenge. In the video, one frustrated individual exclaims, "I'm a mechanical engineer, not an electrical engineer!" (Private Universe Project, *A Private Universe* [Videotape], Harvard-Smithsonian Center for Astrophysics: Cambridge, MA, 1989)

The answer appears on page 20.



MIKE CIESIELSKI

descent lightbulb. Generating intense heat, halogen lamps have been known to cause fires. Homeowners must take care to keep the lamps away from draperies and other combustibles. Furthermore, since halogen bulbs give off so much heat, their energy efficiency gets low marks.

Neon lights

Neon signs are similar to fluorescent lamps, except that they contain no mercury or phosphor coatings. A mixture of neon and other gases within the tube gives off colored light when the electrons are excited by an electric current.



Light-emitting diodes

Eventually, both incandescent and fluorescent bulbs may give way to light-emitting diodes (LEDs). The indicator lights on computers and the numbers on digital alarm clocks utilize LEDs—light sources based on the properties of semiconductors such as silicon. For an explanation of how these durable devices operate, see "Light-Emitting Diodes—Tune in to the Blues" in the April 2001 issue of *ChemMatters*. 🎵

Brian Rohrig teaches chemistry at the Eastmoor Academy in Columbus, OH. His most recent article for *ChemMatters*, "Matches—Striking Chemistry at Your Fingertips", appeared in the December 2002 issue.

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EDIBLE WRAPS

Safe, Strong, and Delicious

Packing your lunch means making the same old choices. Peanut butter? Tuna? Ham? Mayo? Cheese? But get ready. Before long, you may be adding the flavor of the wrapper to that list.



By Mary Pat Mahoney

Wrapping food to keep it fresh is important. In some parts of the world, banana leaves or cornhusks do the job. After all, nobody wants to eat dry, discolored foods.

Today, we rely on a variety of plastic wraps to keep air and moisture from spoiling our lunch. They're thin, flexible, lightweight, and cheap.

But from the "green chemistry" perspective, plastic has problems. Its manufacture normally requires the use of petroleum, a nonrenewable resource. Furthermore, the final product isn't biodegradable. In other words, your plastic sandwich bag could still be sitting around in a landfill when you're packing lunches for your own kids!

Edible food wraps have many of the same properties we look for in plastic wraps. And the best news is that they are a lot easier on the environment. Not that there aren't some obvious drawbacks. You might not care to eat a wrap that's been sitting in your locker or at the bottom of a lunch bag for most of the day. Nevertheless, Tara McHugh, a food scientist at the United States Department of Agriculture (USDA) believes that the idea has potential.



In a January interview for *FoodTech Cyber Magazine*, McHugh explained that the wraps are not intended to replace the synthetic packaging necessary for market displays. "You still want to cover it with packaging because if you want to consume the wrap, you wouldn't want everyone in the store handling it." But, within the package, wrappers could safely be of the edible variety.

She's developed a variety of fruit and vegetable-based wraps for munching right along with your lunch. She makes the edible wraps by using concentrated, unstrained fruit and vegetable purees that look a bit like

MIKE DIESELSKI



MIKE GIESIELSKI

thinned-down baby food. By a process called “casting”, it’s poured onto Teflon-coated trays that are placed in a dehydrator. Water evaporates from the puree, leaving behind brightly colored flexible films, each about as thick as a piece of paper. As for flavor, the film retains the intense taste of the original fruit or veggie.

Polymer power

Both edible and plastic wraps owe their strength and flexibility to long chainlike molecules called *polymers*. In Greek-derived scientific terms, *poly* means “many” and *mer* refers to a unit. A polymer is a compound made from several smaller units or monomers (*mono* means “one”). Think of a polymer as the line of cars on a train or the beads on a string. The monomer is the individual car or bead, and the polymer is the entire unit.

The plant-derived polymers used to make the edible wraps are called polysaccharides. Plant cells take sugar (saccharide) molecules like glucose ($C_6H_{12}O_6$) and join them end-to-end, subtracting water molecules in the process. The resulting long polysaccharide chains give the wraps flexibility. One polysaccharide, cellulose, found in tough plant cell walls, gives the wrap strength.

A wrap made entirely from fruits or vegetables may be good at preventing air from getting to food, but it isn’t very water-resistant. It quickly dissolves in your mouth (good news) or in water (bad news). Polysaccharides are *hydrophilic* (water loving), a term describing a natural attraction to water molecules. Surrounded by water, neighboring chains dissociate from one another, and the wrap falls apart in the process.

The hydrophilic property may be an advantage or a disadvantage, depending on how you want to use the wrap. If you want to take it off your sandwich and eat it, you’ll appreciate how it melts (actually dissolves) in your mouth. If you want a tomato wrap to cover your frozen pizza, or a carrot wrap to seal a casserole to save in the fridge for later, you’ll need a more water-resistant wrap.

Dr. Tara McHugh creates wraps from a range of produce including apples, peaches, mangos, guavas, papaya, apricots, tomatoes, carrots, and even broccoli. Her role, as she sees it, is to develop technologies for enhancing the value of the nation’s food supply. It will be up to commercial companies to further refine the process to make marketable products.

To improve water resistance, Dr. McHugh adds lipids—like vegetable oils—to the puree. Lipids, a category including all fats and oils, are hydrophobic (water hating), which means they resist water. The lipid-enhanced wraps are more water-resistant than those without, but they still aren’t as moisture-sealing as plastic.

COURTESY FOOD RESEARCH CENTER, USDA/ARS/WESTERN REGIONAL RESEARCH CENTER

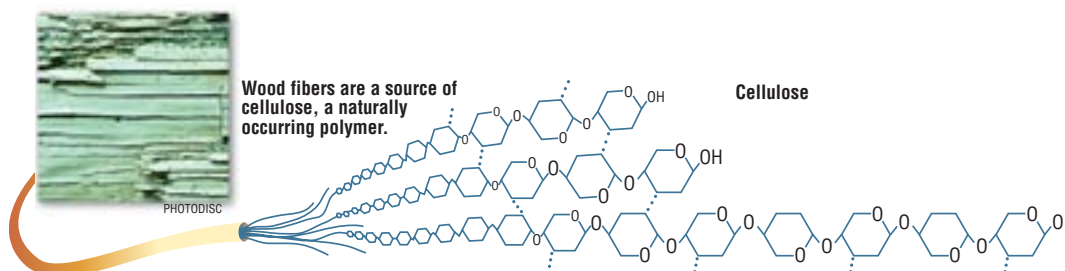


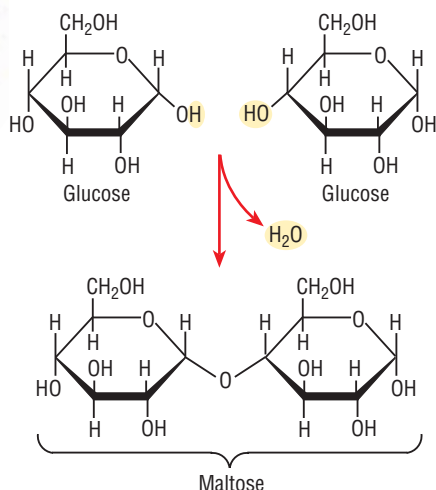
Green wrappings

Besides offering an alternative to plastic wrap, the edible wraps have additional advantages. First of all, they’re green. Well, actually, only the broccoli wraps are green, but they’re “green” as in environmentally friendly. A wrap made out of vegetables and fruit is biodegradable. It won’t end up in a landfill like plastic wraps and containers. Even if the wrap is thrown away, it breaks down into simple sugars.

Early in 2003, we read the latest information on what constitutes a healthy diet. The experts differed on some points, but on one, they were clear: People should eat five to nine servings of fruits and vegetables a day. Look around the lunch table, and you’ll find most people eating less than half that much. Wrappers to the rescue? Made from highly concentrated purees, one wrapper is equal to a serving of a fruit or vegetable. Putting away a few at lunch might make good diet sense.

The next time you’re in the grocery store, take a look at the produce. Nearly every apple is picture perfect. Consumers look for only certain shapes, sizes, and colors. But what about the not-so-perfect-looking apples that are still good to eat? Farmers need markets for all of their produce—even what isn’t perfect. Made from concentrated purees, edible wraps take lots of apples, peaches, mangos, carrots, or tomatoes. Their processing plants will provide farmers with another place to sell their goods.





Plant cells take sugar molecules like glucose and join them end-to-end, subtracting water in the process.

Plastics everywhere

From a practical standpoint, plastic wraps are hard to beat. For water resistance and strength, plastics are a superior choice.

It's hard to imagine a world without the vast variety of plastic products. Look around you. Your backpack, pen, clothing, desktop, and watch are probably either entirely or partially made from plastic. You might even have some plastic in your mouth if you wear a retainer or in your eyes if you wear contacts.

Plastics are amazingly versatile. You want thin? Plastics can be as thin as a piece of food wrap. How about strong? Use it to create bullet-proof vests. You like your plastic clear or neon green? No problem. Rigid enough to form car bumpers? Delicate enough for a human heart valve? It is woven into clothing, stuffed into sleeping bags, even made into soft and squeezable baby toys.

Both edible wraps and plastics are made with polymers. But unlike the edible wraps, the polymers in plastics are created in a laboratory. Plastics begin with raw materials like ethylene (C_2H_4), propylene (C_3H_6), or benzene (C_6H_6)—all derived from fossil fuels. To make polymers, the starting compounds are taken apart and rearranged. Additional compounds might be added to synthesize a custom-made polymer that meets the specifications for a product. Versatile plastics can be extruded, molded, rolled, spun like thread, or foamed.

They can be clear, opaque, tinted, brittle, soft, flexible, or rigid. The possibilities are nearly endless.

To create a simple plastic milk jug made from polyethylene, often stamped #2, HDPE, chemists use the process called *addition polymerization*. Ethylene (C_2H_4) from crude oil serves as the monomer.

A catalyst starts polymerization, soon forming a lengthening chain as molecule after molecule of ethylene is added. In the addition process, the double bond of each ethylene is replaced with bonds to the preceding monomer and to the following monomer.

As more and more units are added, the polymer grows like beads on a string.

Some polymers are synthesized by a process called *condensation polymerization*. This time, instead of adding one monomer after another like beads on a string, the monomers clump together to form short chains. As the process continues, shorter chains clump together to form longer chains. Each time two smaller chains combine, a small molecule is usually split off to allow the previously separated ends of the chains to bond to one another. In this way the monomers gradually join to form long polymers.

It's a wrap

Plastics have many appealing qualities. But one of their most useful features is also one of their biggest drawbacks—durability.

Plastics last a long time—a *very* long time. What's wrong with that? You want your plastic lunch bag to last through lunch and your pen to survive a few months of school. But neither needs to last forever. Unlike a leaf, a slice of bread, or a nail, a plastic item doesn't decompose when exposed to air, water, and microorganisms. In fact, a plastic sandwich bag can last up to twenty years or more in a landfill, and plastic bottles last considerably longer.

So what do you want in your polymer? Edible wraps can't match plastic wraps for durability and water resistance. Yet, plastic wraps are not as "green" as edible wraps no matter what the color.

It won't be long before you'll have another choice to make at the grocery store. The edible wraps are expected to be on market shelves some time this year. Will they be on your sandwich? That's up to you. 🍌

Mary Pat Mahoney is a science writer from Colleyville, TX.

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2028 Use Your

Chemagination

By Michael Tinneland, Marisa Burgener,
and Sharon Worthy

Imagine you have just received the latest issue of *ChemMatters* magazine. The cover article catches your eye. It is an article on a new automobile paint that can change colors. You think, "Cool, I wonder why I haven't heard about this before?"

You flip to the next article and read about "Digisniff", a new accessory for computers and televisions that allows you to smell what you are seeing on the screen. Your eyebrow slowly begins to rise. Something doesn't seem quite right, but at the same time it all sounds so ... so *possible*. By the time you reach the article on "Endless Chew", the reusable bubble gum, you know something is up. You start looking over your shoulder for the hidden camera! Finally, looking back at

the cover you notice the publication date is April
2028—25
years

into
the
future!

What are you reading? You are looking at some of the winning articles from Chemagination, a science essay and poster contest sponsored by the American Chemical Society (ACS) and hosted by ACS membership groups as part of their local, regional, and national meetings. Local ACS sections invite area high schools to participate and supply them with copies of *ChemMatters* for inspiration.

The contest poses an interesting challenge to high school students: Write an article for *ChemMatters* magazine as it would appear 25 years in the future. Rules state that the article must answer the question: "*What breakthrough or innovation related to chemistry and/or its applications, will improve the quality of a teenager's life 25 years in the future?*" Categories include biotechnology,

medicine/healthcare, new materials, transportation, and environment.

But there is one stipulation that helps focus the challenge. The idea is to come up with innovations that are possible

**What if you could
read a magazine
dated 25 years
into the future?
What would
ChemMatters
look like?**

in 25 years. The featured technology should be a logical extension of today's knowledge, rather than wild fantasies with little chance of coming true. With this added feature, the contest becomes a great way for students to apply what they know, with a creative twist thrown in.

And there is plenty of creativity in the projects of a recent group of national contest winners. How about this for starters? A waterproof microchip which, when suspended in the toilet, provides instantaneous urinalysis data. That was the inspiration of eleventh graders Ashlee Hillerud, Kate Mashburn, and Kaylee Nuckolls from Mansfield, TX. According to their design, wires run through the bathroom wall from the chip to a touch-screen monitor hidden in a cabinet or mounted on a wall. The microchip takes a reading of the urine each time the bathroom is used. The results—analyzed and displayed on the monitor—tell what diseases or abnormalities might be present and suggest immediate modifications for the user's diet.

Then there was the winning project dealing with new materials from Beaumont, TX, entitled "New Cars Save Lives". West Brook High School students Tarannum Jaleel, Muniba Riaz, and Elizabeth Leary wrote about a newly created metal polymer, more shock-absorbent than the materials currently used in cars. This extra protection was designed to keep accident-prone teenage drivers safe in case of a collision—effectively making the new cars more like today's bumper cars.



But the contest involves more than writing. Participating students are expected to create a table-top display, much like the ones for a science fair, showing the students' design for the cover of the future issue of *ChemMatters*, as well as the text of their article.

Communication skills are important too. At least one member of the team attends the contest to present the display in an interview with two or three judges. The judges then

assign points based on the display and the student's ability to convey the science behind their ideas.

Although students are having a good time with the creative challenge, teachers are enthusiastic about the

educational aspects of the contest. Most schools are making curriculum decisions based on a set of National Science Education Standards—standards that encourage teachers to help their students think creatively and to appreciate how knowledge is acquired. For example, there are the History and Nature of Science Standards. These state that students need to understand that “science reflects its history and is an ongoing, changing enterprise.”

That's what Chemagination is all about—changes in the future that reflect what is happening in science laboratories right now.

Working on a Chemagination project, students see science as a human endeavor—one that involves asking questions, thinking critically and logically, making decisions based on data, and communicating scientific arguments.

While it's not easy to predict where real technology is headed, students involved in Chemagination projects haven't hesitated to try. Winning projects take recent breakthroughs in science and then give them additional ‘twists’ to show what might result from them in the not too distant future.

Are you starting to “chemagine” an idea for your own project? To find out if a contest is scheduled in your area in the near future, check the Chemagination Web page at <http://chemistry.org/chemagination>. 📖

Michael Tinnesand is an Associate Director of the Education and International Activities Division and **Marisa Burgener** heads the Chemagination project of the Office of Community Activities at the American Chemical Society. **Sharon Worthy** is a Senior Communications Officer at ACS.

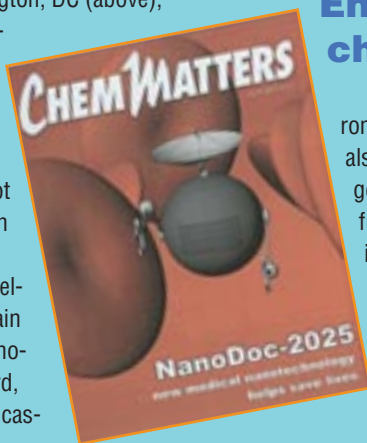


Just chemagine these possibilities!



NanoDoc-2025

This winning project from 10th grader Andrei Munteanu and 12th grader Iheanyi Umez-Eronini of Benjamin Banneker High School in Washington, DC (above), envisions a breakthrough for making medical diagnoses. The revolutionary device—a nanobot weighing less than 50 nanograms—works at the subcellular level. The main component of NanoDoc-2025 is a hard, smooth spherical cas-



ing enclosing a control center and an onboard medicinal drug dispenser. Eight different drugs, each compressed under pressures of up to 15 atm, can be loaded for treating patients for more than two years without recharging.

Since the nanobots would be targeted by the body's immune system as invaders, the designers selected smooth, chemically inert diamond for fashioning the outer armor. Powered by the body's own process of glycolysis, the nanobots consume sugars from the diet. Oxygen, needed for the metabolic fuel cell to function, would be carried on board under extreme pressure.

Endless chewing gum

Not only does the gum keep the environment free of unsightly discards, but it also provides welcome relief for the budgets of cash-impaired teenagers! (Your future kids seem to have the same funding problems you have!) Imagine this. You buy one piece of durable gum base and keep the flavor flowing by purchasing tiny flavor cartridges, each of which is guaranteed to last for 30 days. Each time you finish a chewing session, you

simply replace the gum in its own special case. Here it is dried, reconditioned, sanitized with a special ionizer, and reflavored for your next enjoyment.

Would it sell? Project designers Andrea Tegeler and Lindsay Blair of Lake Howell High School in Winter Park, FL, think the answer is yes. According to their research, “The average American in the year 2000 chewed at least 300 sticks of gum. Total sales reached over 2 billion dollars.”

Remember, you saw it here first!



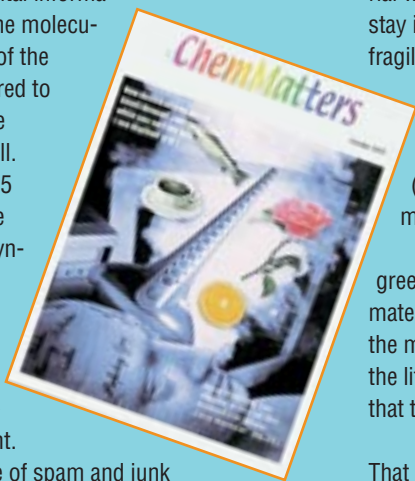
Digisniff

Tyler Miller, Jordan O'Keefe, and Phil True, students at Mechanicsburg Senior High School in Pennsylvania entered their invention, The New Digisniff Machine, in the New Materials category.

Their technological breakthrough is a

device that allows you to smell what you see on your computer screens instead of just watching and hearing it. Operating much like a sound card in a computer, it makes scents rather than sounds. The source site gives the machine digital information about the molecular makeup of the esters required to duplicate the desired smell. Within 10–15 seconds, the esters are synthesized in your own terminal to provide you with the desired scent.

Beware of spam and junk sites! Things could get ugly!



Self-healing plastics

Matthew Teachworth and Diana Tehrani of La Jolla High School in California ask you to imagine the savings for high school athletic budgets if damaged sports equipment repaired itself! Their technological breakthrough of the future is a line of plastics that would do just that.



They admit that their biggest design hurdle was finding a way to imbed resin capsules throughout the material so that they would only break when the surrounding matrix was significantly damaged. The capsule wall material was the key. It had to be strong enough to stay intact during normal wear and tear, but fragile enough to break when the material required repair. They eventually designed some high-strength fibers to be embedded in a polymer matrix. (Actual formulations are proprietary information!)

The designers have thought about the green chemistry aspects of the self-healing materials. While the embedded resins render the material almost impossible to recycle, the lifespan of the original object is so long that the item is unlikely to be discarded.

Just make sure you like the color. That helmet is going to be around for a long time!

Shaping lives forever: Prosthetics

Another team from La Jolla High School went to work envisioning a medical breakthrough for making the lives of amputees more normal than was ever thought to be possible. Rachael McDonald, Azlynn Hare, Joshua Shoreinstein, and Kylise Hare designed a line of prosthetic limbs which would not only look like normal body parts, but feel normal to the person fitted with them. A lightweight titanium alloy would ensure proper support after being attached to the residual natural bone. Then the limb is fashioned with lifelike synthetics covered with a latex “skin”. An elaborate biofeedback system gives the owner full control and sensory input. Power is supplied by a tiny, well-shielded nuclear reactor. The only disadvantage? The new limbs are so sensitive that accidental impacts and pinches are once



again the painful experience nature intended. “Ouch” is back in the wearer’s vocabulary.

Chocolate—A guilty pleasure no more

We may have saved the best news for last. Shivani Patnaik, Abhik Saha, and Natalie Nicolas of Damascus High School in Maryland bring us CocoaVeggie, the chocolate substitute that is not only delicious, it’s good for us. Packed with vitamins and

minerals, it replaces one of our daily recommended servings of vegetables. A genetic engineering breakthrough enables the designers to make a cocoa plant variety in which selected genes for nutritious vegetables are brought

on board. Even the taste is superior!

And if all else fails, you can always wait for the innovation thought up by these 10th graders from Damascus High School to come true. Their project, entitled, “A Guilty Pleasure No More—A Healthy Treat for Us To Eat—Genetically Modified Chocolate” should solve all our problems. CocoaVeggie is the chocolate replacement that is actually good for you. Packed with loads of vitamins and minerals, it can replace one of your five daily recommended servings of vegetables.

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Chemagination



Sometimes “invention” means looking at a familiar object in a different way. “Different” may be an understatement for this view of the ordinary lightbulb.

The *Inventor's Digest* (www.inventorsdigest.com) permits us to share this interesting “theory”, which first appeared in its January/February 1997 issue. For your April Fools’ Day enjoyment, *ChemMatters* presents **The Dark Sucker Theory**. For years, we’ve assumed that electric bulbs emit light, but recent information proves otherwise. Electric bulbs don’t emit light, *they suck dark!* Thus, we propose changing the name of these bulbs to the more accurate term—Dark Suckers. Furthermore, the existence of Dark Suckers proves that dark has mass and is heavier than light. (We’ll get to that later.)

Electric bulbs suck dark.

Look at a Dark Sucker in the room you’re in. Note that there is much less dark right next to it than there is elsewhere. The larger the Dark Sucker, the greater its capacity to suck dark. Notice that the Dark Suckers in the parking lot have much greater capacities than the ones in your room.

There are also portable Dark Suckers. In these, the small bulbs cannot handle all the dark by themselves and must be aided by a Dark Storage Unit (DSU). When the DSU is full, it must be either emptied in a

recharger or replaced by a fresh DSU before the portable Dark Sucker can operate again.

Even the sturdiest Dark Suckers don’t last forever. Once they are filled to capacity with dark, they can no longer suck in more. The dark spot appearing on a full Dark Sucker is irrefutable evidence that this is true.

A candle is actually a primitive Dark Sucker. Look at a new candle. It has a white wick. But after its first use, the wick turns black from all the dark it has absorbed. If you put a pencil next to the wick of an operating candle, it will also turn black. That’s because it got in the way of the dark flowing into the candle.

There are other statements that naturally relate to the original true statement:

Dark has mass.

When dark goes into a Dark Sucker, friction from the moving mass generates heat. Thus it is not wise to touch an operating Dark Sucker. Candles present a special problem since the mass must travel directly into a solid wick instead of first passing through glass. This generates a great amount of heat, and therefore it’s especially unwise to touch an operating candle.



Dark is heavier than light.

If you were to swim just below the surface of a lake, you would see a lot of light. If you were to slowly swim deeper and deeper, you would notice it getting darker and darker. If you were in a deep-sea diving vessel, you might submerge to the point of nearly total darkness. This is because the heavier dark sinks to the bottom of the lake or ocean and the lighter light floats to the top. In fact, now you know why we call it “light”.

Dark is faster than light.

Stand in a lighted room in front of a closed, dark closet. Slowly open the closet door. Do you see the light slowly entering the closet? But, since dark is so *fast*, you couldn’t see the dark leaving the closet!

A challenge: Stump Dr. Darkbulb!

So, you aren’t buying any of this? We didn’t think you would. But, you must

admit that arguing with any of the points poses an interesting challenge.

How would you go about proving that lightbulbs emit light? That dark is not “heavier” than light? What kind of evidence would you need to back up your argument? What fair tests would you set up?

We would like to hear from you.

ChemMatters has enlisted the services of our own “Dr. Darkbulb”. Send your arguments and theory challenges to him at chemmatters@acs.org. We will select a few for his rebuttal. Winners will be posted on our Web page at www.chemistry.org/education/chemmatters.html.

Good luck, and may the force be with you!



ACS CLIP ART



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Here’s the enlightening answer to our challenge on page 13. The secret behind the glow is this: The lightbulb must be part of a circuit—a continuous metal pathway through which electrons move from the power source, through the wire, through the bulb filament, and back to the power source.

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