

CHEM MATTERS

OCTOBER 1984



CHEMISTRY IN OUR UNIVERSE

From the big bang
to the little pop!

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Available

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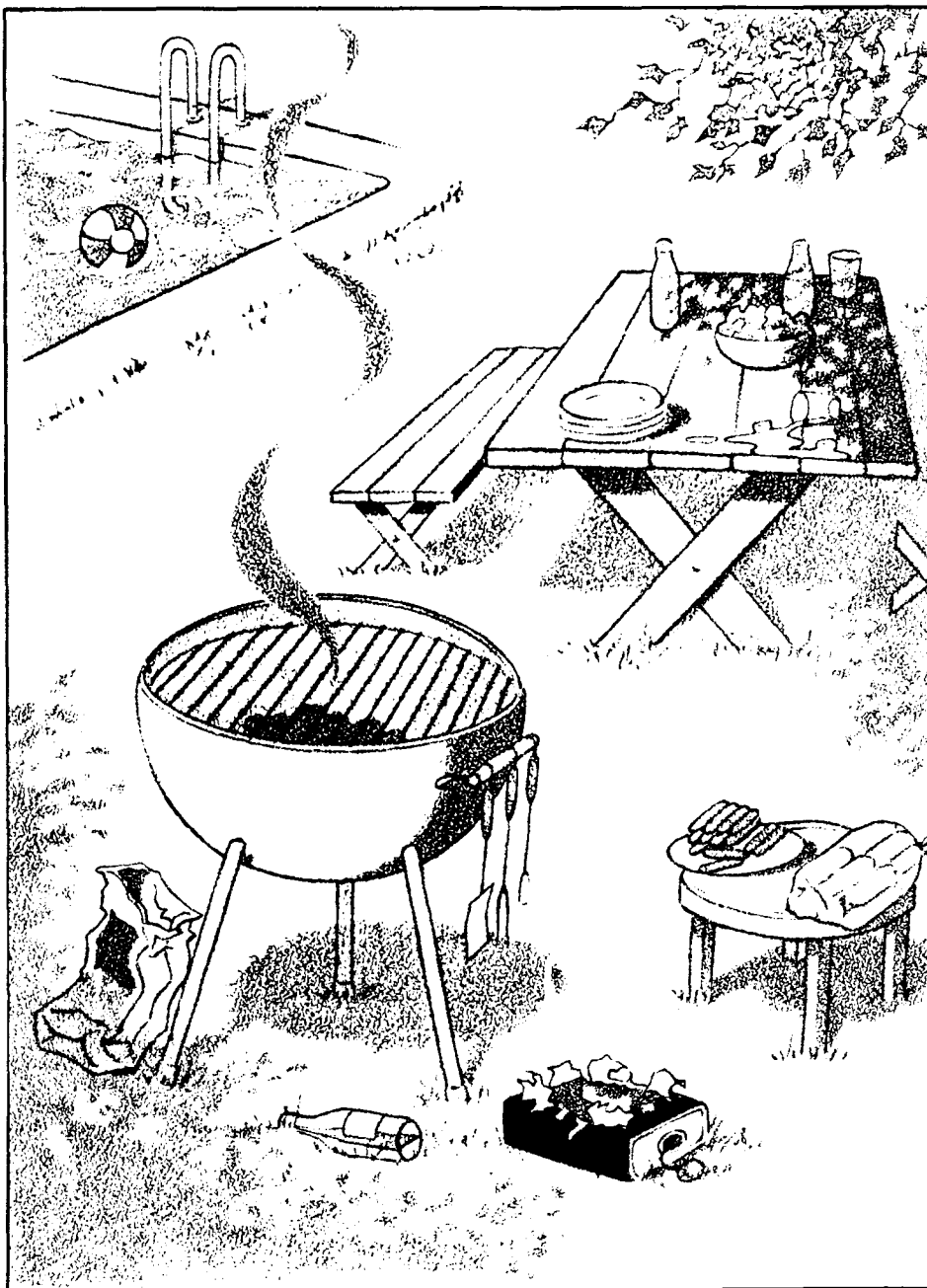
Steve, a salesman for a well-known manufacturer, was a hard worker. He had just made the biggest sale in the history of his company. To celebrate, he invited his friends to a barbecue party at his backyard pool.

The party started well, with everyone enjoying the bright sun and cool water. When his guests had worked up an appetite, Steve turned his attention to starting a fire in the charcoal grill. He made a cone-shaped pile of briquettes and squirted on some charcoal lighter fluid. He waited a few minutes, then lit the pile with a match. As the flames began to ignite the briquettes, Steve was called inside to the telephone. The conversation was longer than he expected, and when he came outside again the fire had gone out—or so it appeared.

Steve retrieved the can of lighter fluid, pointed it at the charcoal and squeezed the sides of the can. There was a loud explosion . . . the can ripped open . . . he was splattered with burning fluid. Steve was on fire! Thinking quickly, he ran to the pool and jumped in. The water extinguished the flames. The ambulance arrived, and Steve was rushed to the burn treatment unit at the local hospital.

I was asked to investigate the accident and try to determine the cause of the explosion. When I interviewed the guests (I was unable to interview Steve due to his serious condition), they offered a simple explanation. The fire had not really gone out. Steve did not notice that the coals were glowing. When he squirted the fluid again, it caught fire immediately, the flames ran up the stream of liquid, entered the can, and triggered the explosion.

This was a simple theory, but I believed it was impossible. Three conditions must be met before there can be an explosion of this kind: 1) A fuel must be present. 2) There must be a



source of ignition. 3) Sufficient oxygen must be present. It was the third condition that concerned me. Where did the oxygen come from? There is oxygen in the air, of course, but there would be no air inside the can. When Steve squeezed the sides of the can to squirt the liquid, this would have forced out most, if not all, of the air inside the can. Not enough oxygen would have been present.

I decided to conduct some experi-

ments to try to resolve the mystery. This would mean intentionally squirting charcoal lighter fluid on burning coals. The trick would be to do it safely. I selected an isolated, paved area which was fenced so no other persons could enter. I wore safety glasses and a face shield. My clothing was fire resistant and well padded. I had two fire extinguishers.

The lighter fluid was placed in an open container mounted 8 ft above the

pavement. With an open container there could be a fire, but no explosion. A long hose, with a clamp valve, led from the bottom of the container. A spout from a charcoal lighter fluid can was attached to the other end of the hose. In this way, the container of fluid was kept 10 ft away from the spout.

I positioned a barbecue grill where the stream of fluid would strike it and started a charcoal fire in the grill. From a safe distance, I opened the clamp, watched, and took pictures. I made repeated experiments—with the fire barely burning and with the grill in full flame.

There was never any explosion. In fact, the flames never traveled up the stream of lighter fluid. If the charcoal was not burning vigorously when I opened the clamp, large clouds of white smoke would form (see box). However, if the charcoal was burning vigorously, the instant the stream of lighter fluid hit the coals, a very large flame jumped out from the grill, sometimes reaching as far as 8 or 10 ft from the grill. Now, the real explanation became clear.

Smoke or Flame?

Charcoal lighter fluid is a mixture of hydrocarbons (molecules containing just hydrogen and carbon). A typical molecule might be C_7H_{16} . When hydrocarbons are oxidized completely (usually burning with a bright flame) they form carbon dioxide and water, as in this unbalanced equation:



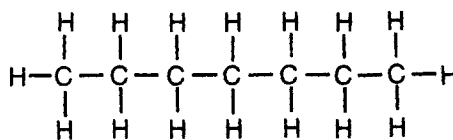
When hydrocarbons are *partially* oxidized, they form *aldehydes* and *ketones*, in which oxygen attaches to the hydrocarbon molecules without converting them to CO_2 and H_2O . These molecules may appear as white smoke.

When Steve added lighter fluid the second time, the coals were in fact burning vigorously, but the glowing of the coals was not visible in the bright sunlight. When the fluid hit the coals, a big flame leaped toward him and startled him. He let go of the can of lighter fluid and the sides of the can relaxed, first sucking in some air and then a tongue of flame. Now, inside the can, there was lighter fluid vapor, flame, and oxygen. The explosion followed.

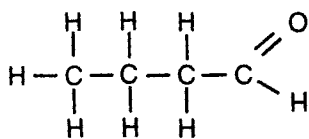
I later received a report that Steve had died. He had been burned severely over 70% of his body.

A special note of caution. Just because we know what caused *this* explosion does not mean that we can add lighter fluid to charcoal twice and get away with it. It means that we have all the more reason to add charcoal lighter fluid *only once* to start the fire safely.

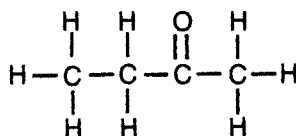
Jay A. Young, a former professor, has written chemistry textbooks and articles. Currently, he is a consultant in litigation matters involving chemical health and safety.



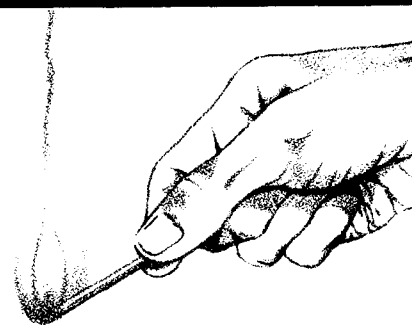
A hydrocarbon



An aldehyde



A ketone



Starting a barbecue safely

1. Remove the cover and the grid from the grill; empty the grill.
2. Using fresh charcoal, arrange briquettes in a cone-shaped pile.
3. Be sure there are no sources of ignition nearby. Do not overlook hot surfaces, candle or lantern flames, lighted cigarettes, cigars, pipes, or other sources of sparks. Squirt the charcoal lighter fluid liberally and evenly on the briquettes. Close the cap securely and place the can at least 15 ft from the grill and away from other sources of ignition.
4. Wait 1 min for the fluid to soak in, but no longer than 5 min. Light the charcoal carefully with a match.
5. If, after 15–30 min, the briquettes are not burning well, douse the fire with a bucket of water and discard the briquettes in a fireproof location. Begin again with fresh briquettes.
6. When briquettes are burning well and covered with white ash, use tongs to spread them evenly. Replace the grid and cover, and enjoy your barbecue.
7. *Never* substitute other liquids for charcoal lighter fluid. Some of these liquids, such as gasoline, kerosene, paint thinner, or alcohol, *will* allow a flame to travel up the stream of squirting liquid or they may explode for other reasons not described in this article.

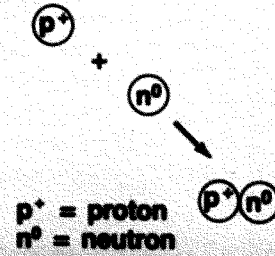
Simplified a bit, the story of elements begins like this . . .



TOO HOT

WAIT
3 MIN
46 SEC

FORM DEUTERIUM



TOO HOT
WAIT 34 MIN



FORM TRITIUM

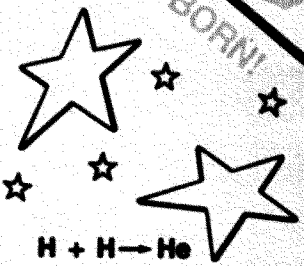
THEN HELIUM



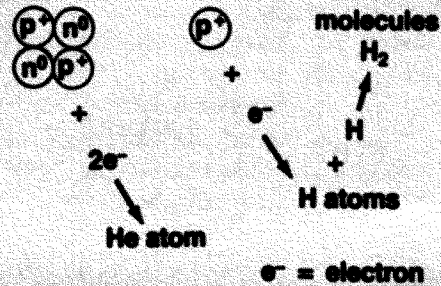
WAIT MILLIONS OF YEARS FOR EFFECTS OF GRAVITY *



STARS BORN!

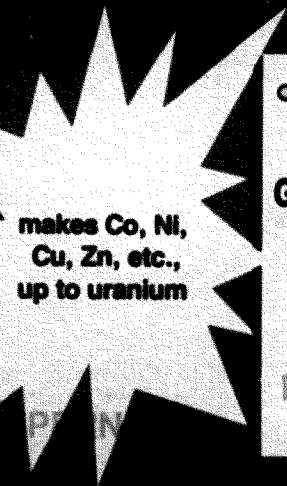


BEGIN CHEMISTRY



STILL TOO HOT

WAIT
100,000
YEARS



GO BACK TO

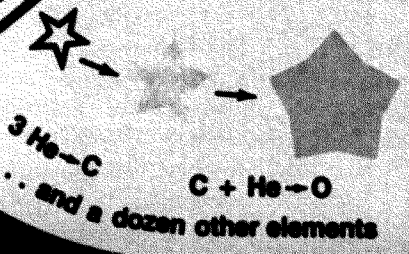


STARS REBORN



WAIT MILLIONS TO BILLIONS OF YEARS

STAR EXPANDS → RED GIANT



HIGH MASS STAR

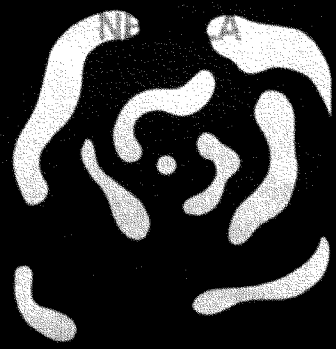
CHOOSE

LOW MASS STAR

GO BACK TO



STARS REBORN



The Origin of the Elements by Ann Finkbeiner

An ancient explosion created the universe and, with it, the elements hydrogen and helium. The stars made all the rest, every last atom of every other element. They made the gold in your neck chain, silver in your ring, iron in your car, aluminum in pans, neon in lights, and silicon in sand at the beach. This, according to the latest theories, is how the elements began.

Modern theories of the universe have become quite refined. Remarkably, they permit scientists to calculate the time—often to a fraction of a second—when certain events are believed to have occurred. The range of time is astounding. The big bang made hydrogen and helium in a few minutes. The stars made some of the other elements, such as carbon and oxygen, over millions or billions of years. Still other elements, like iron, were made during the deaths of stars, in just a couple of seconds.

Climbing a ladder

When scientists first considered the origin of the elements, they suspected that all elements were somehow constructed from hydrogen. A hydrogen nucleus is simple—just one proton. Add another proton, that is, add a hydrogen, throw in a couple of neutrons, and you get helium. In the simplest sense, helium plus a hydrogen is lithium, lithium plus a hydrogen is beryllium, beryllium plus a hydrogen is boron, and so on. Building the elements was thought to be like climbing the rungs of a ladder.

The *idea* seemed fine. The problem was getting nature to agree. Some nuclei are easily formed and hold together well. They make strong rungs on the ladder and allow the process to advance another step. Other nuclei are fragile—overstressed by the repulsion of proton for proton—and may exist for only a fraction of a second be-

fore ripping themselves apart or disintegrating in a collision with another nucleus. These nuclei are weak rungs which give way immediately, bringing the process to a halt.

It seems that, during the creation of the elements, nature encountered several weak rungs. The buildup of the elements got just so far, then encountered an obstacle that brought the process to a halt. After a delay, which could last billions of years, conditions changed and the synthesis of elements continued on . . . until it encountered the next weak rung.

The first step

In the beginning, shortly after the big bang, there were protons and neutrons. The protons and neutrons collided with enough energy to stick, forming an *isotope* of hydrogen called *deuterium* (see box, "Hydrogen 1, 2, 3"). But deuterium is a loosely bound nucleus and is blown apart in collisions with other particles before it can form *tritium*, another isotope of hydrogen and the next step up the ladder. For a short while, the creation of elements stopped at the deuterium roadblock.

But the infant universe was expanding and cooling rapidly. Three minutes and 46 seconds after its beginning, according to theoretical calculations, the temperature fell to one billion degrees, and collisions became less likely. Deuterium stuck together, and the roadblock broke. Deuterium captured a neutron to form hydrogen three, tritium. Tritium captured a proton and changed to helium. The second element was formed.



Helium nucleus (alpha particle)

Nature repaired this broken rung quickly. Within 34 min after the big bang, the universe consisted of 22% to 28% helium. The rest of the universe remained on the first rung, hydrogen.

Stuck again

But here nucleosynthesis—the formation of new elements—stopped at another missing rung. Helium with its two protons and two neutrons has a *mass number* of four. (The mass number is simply the sum of the number of protons and neutrons in the nucleus.) No nucleus with a mass of five exists in nature. All mass fives are unstable. Physicists have artificially produced helium five in their laboratories, but it breaks apart instantly into helium four.

Looking back on it now, some 15 billion years later, we can see that a universe of hydrogen and helium nuclei is fairly useless. You can't *make* anything—no compounds, or proteins, or life. Nuclei need electrons to attach to each other . . . to form molecules. Electrons were around at the beginning, but at 34 min, the universe was still too hot for electrons to bind to nuclei and form atoms. Not until 100,000 years later was the temperature low enough that atoms could combine by sharing and exchanging their electrons—and chemistry began.

Hydrogen 1, 2, 3

All atoms of the same element have the same number of protons (p^+), but may have different numbers of neutrons (n^0). Varieties of the same element with different neutrons are called *isotopes*. Hydrogen has three isotopes (their nuclei are shown here), which can be identified by name or by *mass number*, the sum of the protons and neutrons.



Common hydrogen (H-1)



Deuterium (H-2)



Tritium (H-3)

A star is born

The universe would still be a collection of hydrogen and helium atoms were it not for gravity. Millions of years later—and astrophysicists argue about how many millions—gravity pulled the thin, diffuse clouds of hydrogen and helium into increasingly thicker clouds. The clouds heated as they contracted until, at about ten million degrees, stars were born.

In the cores of stars, nucleosynthesis began again. In what seems like a step backward, the stars once again began converting hydrogen into helium. However, due to a shortage of neutrons, the stars used a different process, called the proton-proton chain. During the big bang the young universe contained lots of neutrons. However, free neutrons (outside a nucleus) live an average of only 16 min. Millions of years later, when stars were formed, few neutrons were around to collide with protons. The stars found another way to make helium, which re-

leased a steady stream of heat and starlight.

Push vs. pull

A normal star continues “burning” hydrogen into helium for most of its life, giving off energy. The temperatures are high, and the hot star wants to expand. But the push outward is countered by the pull of gravity inward, and for millions and billions of years, the star remains stable.

Sooner or later, however, the hydrogen supply runs out, with dramatic results. Nucleosynthesis in the core stops, and the balance tips in favor of the inward pull of gravity. The core contracts, then heats up. The outer envelope of the star expands violently, and the star becomes a red giant. When the contracting core reaches 100 million degrees, the second roadblock breaks, and nucleosynthesis begins again. For the first time, in what’s called the helium flash, helium ignites.

In helium burning, two helium nu-

clei, also called *alpha* particles, collide to create the element beryllium. Beryllium lives for 10^{-16}

(0.0000000000000001) seconds before it disintegrates, but during that time, it can combine with another helium to make carbon. In essence, three heliums collide practically simultaneously, in what’s called the triple alpha process, to make carbon. Carbon combines with helium to make oxygen, but the temperature is not high enough to create any elements heavier than oxygen—the third broken rung.

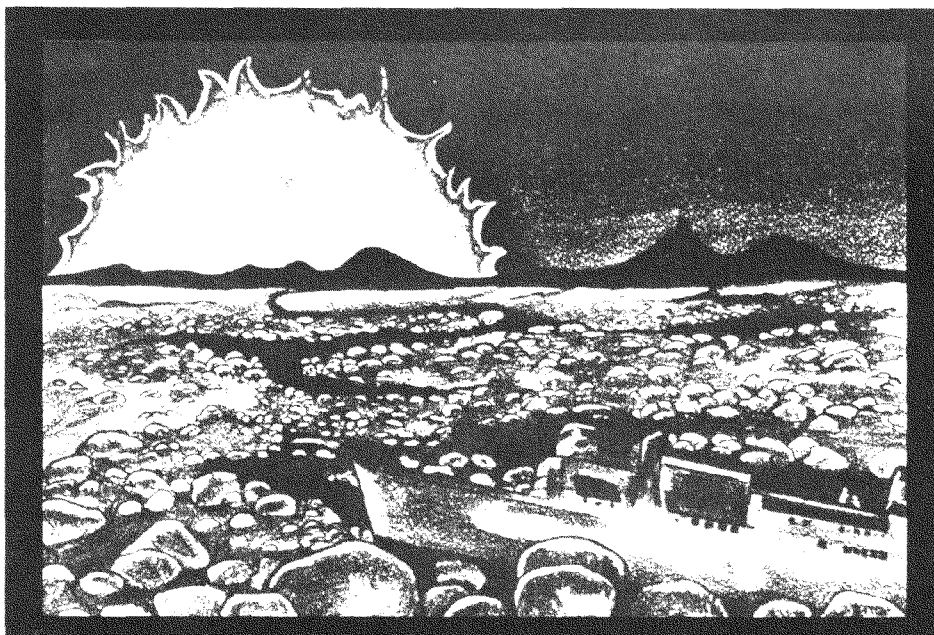
Gravity and temperature are once again in balance, and the star will remain stable, creating carbon and oxygen from helium for a million years or so. Then, when the helium supply in the core runs out, history repeats. The core will again contract, then heat up to 600 million degrees, hot enough for carbon to begin burning. The core is now layered like an onion: Carbon and oxygen burn in the center; helium just outside; and hydrogen outside that.

Then anything goes. Carbon combines with carbon, or oxygen with oxygen, or carbon with oxygen to create neon, sodium, magnesium, aluminum, and silicon. At one billion degrees, silicon burns, combining with helium to make sulfur, argon, calcium, manganese, and everything in between, in layer upon layer of heavier and heavier elements, all the way up to iron. But at iron, at five billion degrees, nucleosynthesis stops: The next rung is missing completely.

The iron barrier

This obstacle is special. The problem is that of all nuclei, iron is the most tightly bound, the hardest to pull apart. All the other processes of nucleosynthesis, from the proton-proton chain to silicon burning, liberated huge amounts of energy. Adding another nucleus to iron will, instead of freeing energy, require it.

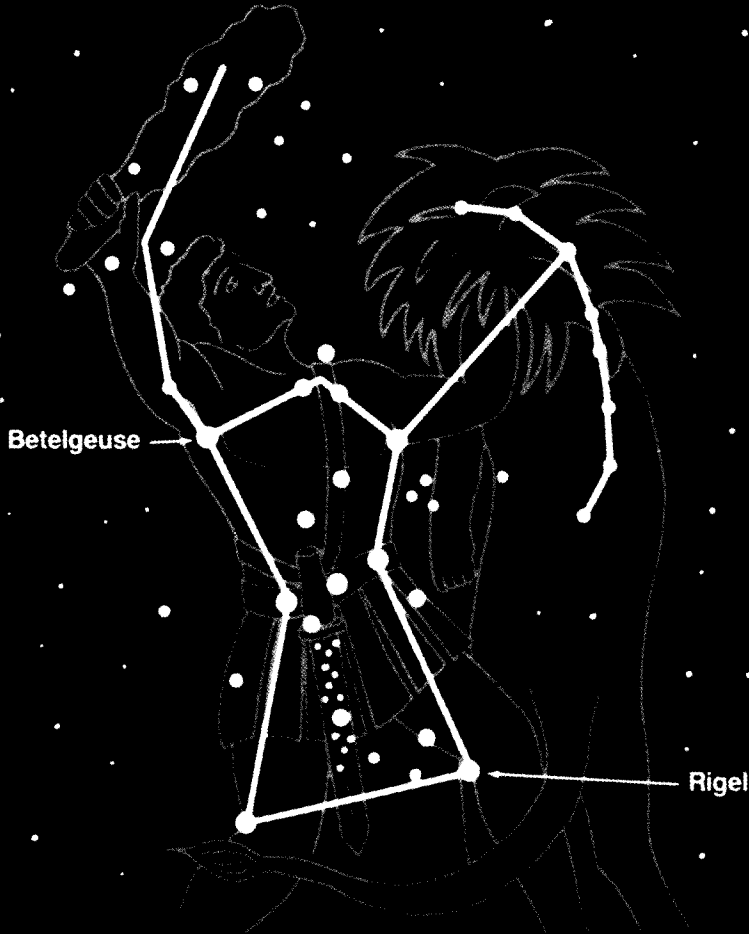
The star has two choices. If the star is 1.44 times the mass of our sun or less, the expanded outer layers will simply drift off in a wispy ring called a planetary nebula. The remnant of the star, called a white dwarf, will burn more and more quietly, get dimmer



Last one turn out the light . . .

Our sun has been a stable star for about five billion years, balancing temperature against gravity, burning hydrogen into helium. About five billion years from now, as its fuel runs low, it will become a hugely distended *red giant*. The monstrous star will scorch the

Earth and dry up the oceans before it engulfs our once friendly planet. Finally, the ravaging star will die as a *white dwarf*. Its outer envelope will become a planetary nebula; its core will burn lower and dimmer until it quietly goes out.



Constellation of elements

Out in the sky, the formation of stars and elements continues. The constellation Orion is a good place to look. The sword dangling from Orion's belt is a brilliant, cloudy birthplace of stars. Rigel, the bluish

star at Orion's left knee, is a huge, hot, young star. Betelgeuse, in Orion's right shoulder, is a red supergiant, unstable, probably burning helium, and possibly on the verge of becoming a supernova.

and dimmer, and eventually go out. "But add that $1/100$ of a solar mass," explains an astrophysicist, "and get out of the way." Gravity again steps in. A star with mass greater than 1.44 times the mass of the sun will contract so violently that the core implodes. In a matter of seconds, the star reaches ten billion degrees. A massive shock wave tears through the star, and it explodes as a supernova.

In this gigantic explosion, dozens of new elements are created—approximately the bottom half of the periodic table. At the extreme temperatures of the supernova, some of the iron breaks up, producing huge numbers of neutrons. The iron nuclei capture the neutrons, and other neutrons inside the nucleus decay into protons, forming cobalt. In the same process of capture and decay, called the "r" or

rapid process, cobalt creates nickel, nickel creates copper, and so on to zinc, gallium, everything up to uranium. Uranium and anything heavier are likely to undergo fission, simply splitting apart into lighter elements. Therefore, no element heavier than uranium, with 92 protons, occurs in nature.

The supernova creates all elements from iron to uranium during its explosion and takes about a second to do so. It throws all its elements—hydrogen through uranium—to drift in veils and filaments through space.

Children of stars

The material cools, mixes with the rest of the matter between stars, and gravity pulls it into clumps. The clumps thicken, condense, heat up, and ignite. Another star, of a second generation,

is born. Because it is formed from remnants of a first-generation star, a second-generation star has different starting material, with traces of heavy elements. This permits the star to combine nuclei in new ways, forming a different mix of isotopes.

Each time a star dies, the neighborhood of the star—the interstellar medium—has more heavy elements than before. Gradually another star will condense, ignite, live and die—once again enriching the interstellar medium. This cycle has been running for billions of years, and, because the universe is still composed mostly of starting ingredients (98%), it will run for billions more.

And what of our own neighborhood, our sun and solar system? The sun is probably a second- or third-generation star; its parent, a nearby supernova. The planets of our solar system probably condensed from clouds of matter, the Earth from a rich mixture of heavy elements. On our planet, in a complex and little-understood process, carbon, nitrogen, oxygen, and hydrogen combined to form proteins and, eventually, life. In a very real sense, we can say that we are all made of stardust.

Ann K. Finkbeiner is a Baltimore freelance writer, specializing in cosmology.

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Weinberg, Steven. "The First Three Minutes: A Modern View of the Origin of the Universe"; Basic Books: New York, N.Y., 1976.

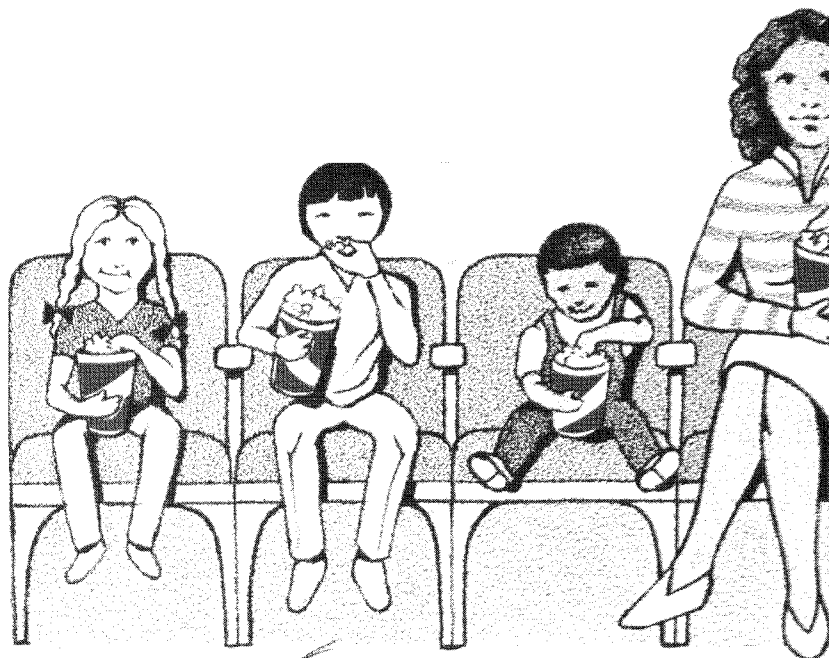
WE BUY OLD GOLD AND SILVER



"What do you mean, it's not antique? This gold is 5 billion years old!"

by Lynn K. Sibley

You're hungry and the cupboard is bare. How about some popcorn? You get out the popper, pour in the oil, and drop in a test kernel. The oil sizzles . . . suddenly the kernel bursts into white fluff and jumps out of the pan. The oil is now at the right temperature. You pour in more corn, just covering the bottom, and wait. A slow tapping builds into a crescendo of happy popping. The kernels beat against the lid, lifting it slightly, and some vapor puffs out. Oh, what an aroma! You open the popper and toss a few of the fluffy mounds into your mouth. But wait a minute! Did you ever wonder why popcorn pops? Just what is that white stuff? And what was the vapor that escaped from the popper? There is a lot of interesting chemistry in a kernel of popcorn. Fill up your bowl and read on.



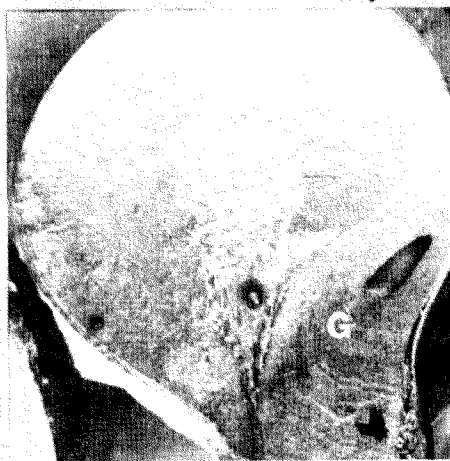
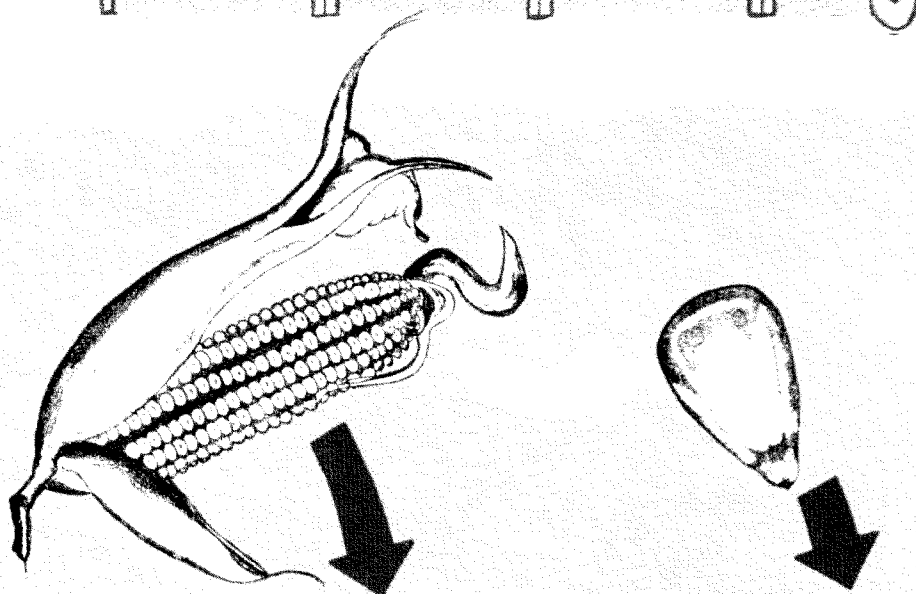
An American heritage

Popcorn, a New World grain, has been an American favorite for centuries. Columbus found the natives in the West Indies not only eating popcorn, but wearing it as jewelry. The English colonists were introduced to popcorn at the first Thanksgiving feast when an Iroquois Indian brought a deerskin bag of popped corn as a gift. Quick to put this unusual food to use, the colonial families ate popcorn with cream, sugar, and fruit for breakfast—the first “puffed” cereal.

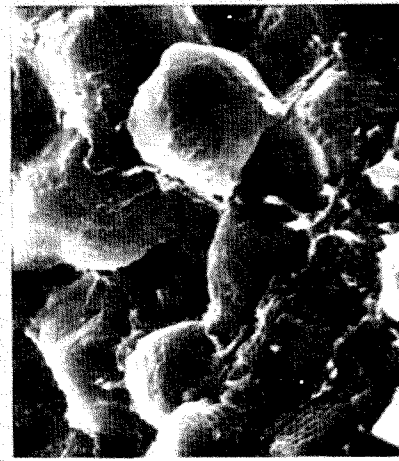
Popcorn took to the streets in 1885 when Charles Cretors invented the steam-powered popping machine. In 1893 our first “junk food” was introduced at the Chicago World’s Fair: a combination of popcorn, peanuts, and molasses called Cracker Jack. Popcorn became an irresistible attraction at baseball games and circuses. Then, a new fad called “movies” took popcorn inside, and sales exploded.

Steam power

Popcorn, a cereal grain like wheat or oats, is about three-fourths carbohy-



An unpopped kernel contains translucent endosperm (TE), opaque endosperm (OE), and germ (G).



Higher magnification shows unpopped starch grains in the translucent endosperm.

CORN



A-maize-ing facts

- Americans eat 9.9 billion qt of popped popcorn annually—that's 42 qt per man, woman, and child (except in Minneapolis–St. Paul, where the champion popcorn eaters consume twice that amount).
- 70% of the popcorn is eaten at home. Most of the rest is munched at theatres, stadiums, and schools.
- Real butter topping disappeared from theatres years ago. That golden liquid squirted on your popcorn is typically partially hydrogenated soybean oil, artificial butter flavor, TBHQ (preservative), and artificial color.
- Until a few years ago, the volume of popped corn was about 25 times as great as the volume of the unpopped kernels. Today's hybrid popcorns expand even more.
- The word *corn* has been used for more than one kind of grain. The biblical "corn" was probably barley. The English "Corn Laws" of 1436 referred to wheat. In Scotland and Ireland, "corn" meant oats. In short, the predominant grain crop grown in an area was usually called "corn." The crop called corn in the U.S. is more properly called *maize*.



After popping, the kernel is mostly a foamlke structure of expanded starch.

drate (starch), with smaller amounts of protein, fat, minerals, and water. The water plays a critical role in the popping process. When heated, the moisture inside the kernel turns into steam. As the pressure increases, the starch expands and the kernel explodes. We like popped corn that is large and tender. This requires just the right amount of water in each kernel. Farmers harvest popcorn when the moisture content is 16% to 19% by mass. To ensure maximum popping expansion, the corn is then carefully cured (dried) until it reaches 13% to 14.5% moisture.

Like other cereals, popcorn kernels consist of three main parts: the *pericarp* (the hull or outer covering), the *germ* (the part that sprouts), and the *endosperm* (the starch that expands). Popcorn acts the way it does because of the special construction of the pericarp and the microscopic structure of the endosperm.

Popcorn has an extra strong pericarp. This tough, protective layer acts like a seal, holding in the steam until the pressure builds up high enough and the kernel explodes. If the pericarp has been cut or cracked during processing, the steam will be vented and the kernel will not pop properly. You may have heard of "hull-less" popcorn. Despite the name, all popcorn has hulls—it could not pop without them. In the so-called hull-less varieties, the pericarp simply separates itself better from the puffed endosperm and has less chance of sticking in your teeth.

Corn has two kinds of endosperm, *translucent* and *opaque*, which are named according to their appearance. The *expansion*, or popping, takes place in the tightly packed translucent endosperm. Therefore, the proportion of translucent and opaque endosperm in a kernel accounts for a corn's "popability." Popcorn contains mostly translucent endosperm, which is better at popping.

Where the action is

When you first put popcorn in the pan, the pressure inside and outside the kernel is the same. As the kernel heats (in oil which is typically 260 °C), the moisture turns to steam, and the internal pressure rises. When the temperature inside the kernel climbs above 100 °C, you might expect that all the water would turn to steam. In fact, only a small amount vaporizes because the tough pericarp acts like a pressure cooker. The boiling point of water is 100 °C at normal atmospheric pressure, but is higher at greater pressures.

The high-pressure steam penetrates the starch granules and transforms them into hot, gelatinized globules. Finally at about 175 °C, when the pressure inside the kernel is about 9 atmospheres (135 lb per sq in.), the pericarp ruptures.

The steam and superheated water (water above 100 °C), now surrounded by normal-pressure air, become the driving force that expands the kernel. The gelatinized starch granules do not explode, but *expand* into thin, jellylike bubbles. Neighboring bubbles fuse together and solidify, forming a three-dimensional network much like a sink full of soapsuds. This is the white fluffy solid we eat. The moisture content of the kernel is now about 1–2%, and the popcorn is transformed into a tender, fluffy morsel.

Popcorn comes in two natural col-

Just for fun

Today, "popcorn boutiques" offer as many as 50 flavors, including root beer, piña colada, and chocolate. Here's how you can make some of these treats:

Pizza popcorn

Pop about 3 qt of popcorn. In a large bowl, mix the popcorn with 1/2 cup of melted butter or margarine. Sprinkle on 3 tablespoons of dry spaghetti sauce mix, 3 tablespoons of oregano, and some parmesan cheese. Mix the ingredients gently—and enjoy.

Pops up years later

Fossil corn pollen, 80,000 years old, has been found in excavations under Mexico City. It is very similar to modern corn. Kernels of popcorn discovered in Bat Cave, N.M., (radiocarbon dated at 2000 B.C.) still pop! Pre-Incan pottery poppers date back 1500 years. According to Paul C. Mangelsdorf, Harvard popcorn researcher, "All races of wild corn were popcorns, their kernels small and flinty in texture, almost impossible to chew, and difficult to grind. Their real usefulness was discovered when kernels accidentally exposed to the heat of glowing coals exploded to become tender, tasty morsels with the nutritional value of whole grain bread. The current pop-

ors, white and yellow, which are located in the endosperm. Artificial dyes color only the pericarp, which is blown off during expansion. In the past, white popcorn was favored by home poppers as the most tender. Today, 90% of all popcorn is yellow—its larger kernels are more flavorful and it has an appealing already-buttered look.

Pop art

You can pop your corn several ways. One is the dry method. Used first by the Indians who popped corn in ceramic pots over open fires, this method has been revived in modern hot-air poppers. More common is the wet method, using oil. Although any vegetable oil will do, pop experts pre-

Fruity popcorn

Place about 3 qt of popped corn in a large, *buttered* bowl. In a saucepan, combine 1/2 cup of sugar with 1 1/2 cups of light corn syrup. Heat to a boil while stirring constantly. When it bubbles, remove from heat and stir in a 3-oz package of fruit-flavored gelatin mix—any flavor you like. When the gelatin has dissolved, *quickly* dribble the mixture over the popcorn. Cool the popcorn on a greased cookie sheet; then break into bite-sized pieces.



Pre-Incan popper, courtesy of Field Museum of Natural History, Chicago.

ularity of popcorn in the home and at the cinema represents a counterpart of corn's first use by the American Indians."

fer coconut oil. The newest method is popping by microwave. The popcorn is sold in a special bag, complete with oil and seasoning. When placed in a high-wattage microwave oven, the bag expands as the popcorn pops.

To achieve maximum poppability and expansion, you must treat your popcorn with care. To maintain the ideal moisture content of 13% to 14.5%, store popcorn in tightly covered containers at a cool temperature. Unpopped corn can lose as much as 3% moisture in five to six days if it is not sealed tightly or if it is too hot. To restore dry popcorn: Fill a quart jar 3/4 full of popcorn, add 1 tablespoon of water, cover tightly, and shake thoroughly. In about four days, your popcorn will be ready to give you its best—it will be tender, fluffy and, oh, so delicious.

Lynn K. Sibley is editorial assistant for Chem Matters. A popcorn enthusiast, she raises gourmet varieties in her backyard.

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EXPERIMENTER'S NOTEBOOK

EXPERIMENTER'S NOTEBOOK

EFFECTS OF POPCORN MOISTURE

HOW IMPORTANT IS MOISTURE TO THE POPPING PROCESS? TO FIND OUT, REDUCE THE MOISTURE CONTENT OF POPCORN KERNELS BY HEATING THEM.

1. MAKE A CORN WARMER BY WRAPPING A PIECE OF ALUMINUM FOIL OVER A LIGHT BULB (40 or 60W) AND BENDING THE EDGES UP. MARK ABOUT A DOZEN KERNELS WITH A PERMANENT MARKER; PLACE THEM IN THE CREASE OF THE FOIL. TURN ON THE BULB AND HEAT THE CORN OVERNIGHT.

2. PLACE ONE OF THE HEATED KERNELS IN A PAN WITH A NORMAL KERNEL. ADD OIL TO JUST COVER THE KERNELS. HEAT THE PAN WHILE WATCHING THE CORN. WHICH KERNEL POPS FIRST? WHICH PRODUCES THE FLUFFIEST POPCORN? DOES THE HEAT TREATMENT AFFECT POPPABILITY?



PERCENT MOISTURE

IF YOU HAVE ACCESS TO LAB EQUIPMENT, YOU CAN DETERMINE THE PERCENT MOISTURE IN YOUR POPCORN.

1. FROM ALUMINUM FOIL, FASHION A LID FOR A BEAKER. WEIGH THE EMPTY BEAKER AND LID. ADD ABOUT 20 KERNELS OF POPCORN; WEIGH THE BEAKER AND LID AGAIN. FIND THE MASS OF THE CORN BY SUBTRACTING THE FIRST WEIGHT FROM THE SECOND.

2. ADD OIL TO JUST COVER THE KERNELS AND WEIGH ONCE AGAIN. PLACE THE LID SNUGLY ON THE BEAKER; PIERCE THE FOIL WITH A FEW SMALL HOLES; THEN HEAT THE BEAKER TO POP THE CORN. REMOVE THE BEAKER FROM THE HEAT AND TAKE OFF THE FOIL. LET THE BEAKER STAND FOR A FEW MINUTES FOR STEAM TO ESCAPE.

3. WEIGH THE BEAKER WITH LID AND POPPED CORN. THE DIFFERENCE BETWEEN THIS MASS FIGURE AND THE ONE FROM STEP 2 IS THE MASS OF WATER LOST. CALCULATE THE MASS OF WATER LOST AS A PERCENT OF THE ORIGINAL MASS OF THE KERNELS.

Robert Bunsen . . . more than a burner design

by Derek Davenport

Some men are fated to be remembered for a single achievement—and not always their most glorious. Who recalls Custer's *first* stand? In chemistry, Erlenmeyer lives on as a flask, Claisen as a condenser, Gooch as a crucible and, above all, Bunsen as a burner. In remembering him thus, we do scant justice to Robert Bunsen, one of the greatest and most influential of 19th-century chemists. Besides, the design of the famous burner was not entirely new and Bunsen's instrument maker played as large a part as he in its perfection.

Robert Bunsen (1811–1899) was born in Göttingen, Germany, the son of the chief librarian and professor of modern languages at the university. A precocious student, he received his doctorate degree at the age of 19. For the next three years he traveled and studied in Germany, France, and England before embarking on his long and illustrious academic career.

His first major research was on *cacodyls*, compounds containing carbon, hydrogen, and arsenic. The most important of these proved to possess the molecular formula, $C_4H_{12}As_2$. These compounds were all highly poisonous, often spontaneously flammable, and were possessed of smells to make a billy goat blush. Bunsen described one of these compounds thus: "The smell of this body produces instantaneous tingling of the hands and feet, and even giddiness and insensibility. . . It is remarkable that when one is exposed to the smell of these compounds the tongue becomes covered with a black coating, even when no further evil effects are noticeable."

From such unlovable compounds Bunsen was able to establish one of the most important principles of organic chemistry—that in some chemical reaction sequences a particular group of atoms (in this case a CH_3 - or



methyl group) may survive totally untouched. This principle was one of the keys to the developing science of organic chemistry, which was shortly to enter upon its golden age. But when an explosion of cacodyl cyanide caused loss of sight in his right eye, Bunsen decided, golden age or not, to swear off organic chemistry for life. His remaining labors were, in the broadest possible sense, concerned with chemical analysis.

An expert glass blower, Bunsen perfected methods for manipulating and analyzing gases in a quantitative way. He then applied these methods to the improvement of blast furnaces, to the analysis of gases from Icelandic volcanoes, and to the photochemical and combustion behavior of mixtures of gases. A minor by-product of this research was the perfection of the immortal burner. He never patented his invention and indeed mentioned it only in passing in some of his papers.

In 1855, the city of Heidelberg was first illuminated by coal gas. Bunsen was in the process of designing his new laboratory (much of his previous work had been done in an abandoned monastery), and he was eager to em-

ploy the gas, not so much to illuminate the laboratory, but to replace the old, inconvenient, and sooty furnaces then in use for heating reaction vessels. He drew on earlier work by Humphry Davy and Michael Faraday in England and on designs by Aimé Argand and others in France, all of whom had tried to perfect a hot and smokeless gas flame. Working with his instrument maker, Peter Degasa, Bunsen devised the burner we know today in which the essential premixing of air and gas is controlled by a rotating sleeve. The trick was to premix the gases in such a way that the flame did not "strike back" even in the absence of the wire gauze used in the Davy–Faraday miner's safety lamp. Bunsen and Degasa succeeded admirably.

Bunsen was the first to use carbon (graphite) cathodes and zinc anodes in a battery, with chromic acid serving as electrolyte and depolarizer. With a bank of these batteries, he purified many of the known metals in bulk—chromium, magnesium, barium, strontium, cerium, and lanthanum. And in the early 1860s he discovered the remaining alkali metals, cesium and rubidium, in certain mineral waters.

Forty tons of mineral water had to be processed to produce 16.5 g of cesium and rubidium chlorides. The names of these two new elements—cesium from the Latin *caesium* or sky blue and rubidium for the Latin *rubidus* or dark red, refer to the most prominent lines in their flame spectra. This in turn brings us to Robert Bunsen's greatest contribution to chemistry—chemical spectroscopy.

About 1860, Bunsen, in collaboration with his physicist colleague, Gustav Kirchhoff, carried out a series of investigations which founded the science of chemical spectroscopy. Every element has its spectroscopic fingerprint. The bluish appearance of the cesium flame is traceable to an emission of radiation of wavelength 455.5 nm; the deep red color of the rubidium flame to an emission of radiation of wavelength 780.0 nm. Furthermore, Bunsen and Kirchhoff showed that "cold" atoms absorb light at exactly those wavelengths which they emit when "hot." Thus the dark lines in the spectrum of the sun (Fraunhofer lines) were explained by selective absorption in the "cold" outer layers of the sun of the continuous radiation produced by the nuclear reactions in the interior (see "Star Born," p. 6). All our present knowledge of the chemical constitution of the stars has come, and will probably always come, from spectroscopic observation, a science we owe to Bunsen and Kirchhoff.

And what part did Bunsen's burner play in all these marvels? Very little, though it did provide a convenient way of generating the high temperature needed to produce the flame emission spectra. Would Bunsen mind being remembered for a minor practical invention? Probably not, for he was an eminently practical man. Better to be remembered for a living burner than a dead theory.

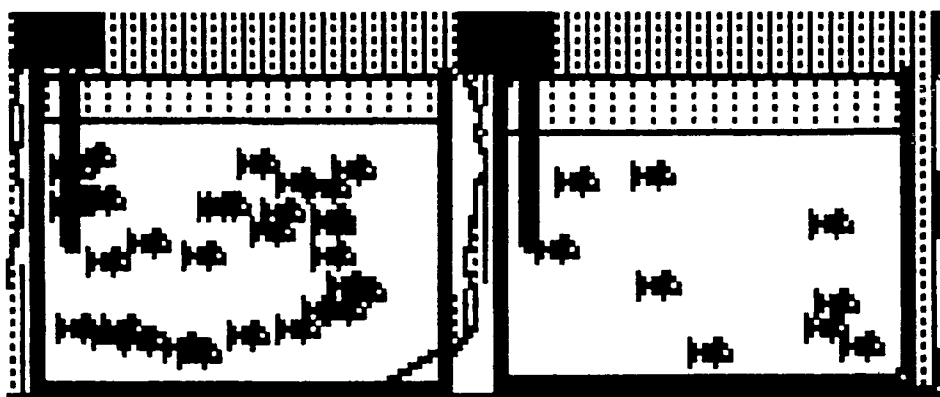
Derek Davenport is a professor of chemistry at Purdue University in West Lafayette, Ind.

You are an ecologist, working for the Fish and Wildlife Service. Your boss gives you a challenging new assignment: Determine why the fish are dying in a certain mountain lake. She is concerned because the problem may threaten a fish hatchery on the same lake. This is a simulation called "Lake Study," running on a personal computer, but it is more than a game because it combines the principles of science with real-world uncertainties and foul-ups. There is no "right" way to approach the fish-kill problem, and you must decide how to proceed.

Suppose you decide to collect samples of lake water. Press the proper key and the computer displays a bird's-eye view of the lake with a boat at a dock. You must steer the boat to the locations where you wish to take samples. Then, you may choose to go to the laboratory. A keystroke later you are in the lab, which has a flame spectrometer, a gas chromatograph, and (thankfully!) a technician to operate them. The lab tests tell you the concentrations of pesticides, heavy metals, and oxygen. Are these levels too high . . . too low? A trip to the library (press another key) provides some of the answers. Now you can

form a hypothesis about what is killing the fish. But the boss demands *proof* and asks for some lab experiments with fish. For each fish tank you must specify the temperature, number of fish, and concentration of pollutants. Later you decide to consult a colleague who points out that you forgot to test the water for mercury. Rats! . . . Head back to the lake for more samples.

"Lake Study" requires no prior knowledge of environmental chemistry, but rewards users who organize data, form concise hypotheses, and design good experiments. The program runs on a 48K Apple II+ or IIe. "Lake Study" was written by David Whisnant, associate professor of chemistry at Wofford College, Spartanburg, S.C. A copy may be obtained by ordering Apple Disk #18 from Project SERAPHIM/Chem Matters, John Moore, Department of Chemistry, Eastern Michigan University, Ypsilanti, Mich. 48197. Send a self-addressed mailing label (or file card) and a check for \$4 made out to Project SERAPHIM. You may wish to ask for the free SERAPHIM catalog, which lists inexpensive chemistry software for several types of computers.



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Final Conditions
32 Fish          9 Fish
20 Deg C        20 Deg C
1 ppm Lead      10 ppm Lead
    
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THE PUZZLE PAGE

Element Search

by *Oliveros McLloyd*

Hidden in the puzzle below are the names of the first 50 elements. Find them forward, backward, up, down, or diagonally. Circle each name as you find it. Some names overlap, using letters more than once; not all of the letters in the puzzle will be used.

Element names (atomic number 1-50)

- | | |
|----------------|----------------|
| 1. Hydrogen | 26. Iron |
| 2. Helium | 27. Cobalt |
| 3. Lithium | 28. Nickel |
| 4. Beryllium | 29. Copper |
| 5. Boron | 30. Zinc |
| 6. Carbon | 31. Gallium |
| 7. Nitrogen | 32. Germanium |
| 8. Oxygen | 33. Arsenic |
| 9. Fluorine | 34. Selenium |
| 10. Neon | 35. Bromine |
| 11. Sodium | 36. Krypton |
| 12. Magnesium | 37. Rubidium |
| 13. Aluminum | 38. Strontium |
| 14. Silicon | 39. Yttrium |
| 15. Phosphorus | 40. Zirconium |
| 16. Sulfur | 41. Niobium |
| 17. Chlorine | 42. Molybdenum |
| 18. Argon | 43. Technetium |
| 19. Potassium | 44. Ruthenium |
| 20. Calcium | 45. Rhodium |
| 21. Scandium | 46. Palladium |
| 22. Titanium | 47. Silver |
| 23. Vanadium | 48. Cadmium |
| 24. Chromium | 49. Indium |
| 25. Manganese | 50. Tin |

T C H L O R I N E O N Y S E P T M O T H D
J M B E C A I U N Z S Q X J N U L S L M Y
R I U S L T J E I U E B V Z I I M A A E O
P U V I R I G F R U T H E N I U M N B Q L
S W B O T Y U O O W Z X A R I L G O S O L
C E G I X N H M U V E T I D Y A A U R S C
A E L O D P O T L C I M O H N L L M A B M
N L N E S I M R F T I H U E S F L U W G S
D C E O N E U U T O R N S N U N I I Y E O
I L H K B I N M I S P E E R I N U H U R R
U P O R C R U O K C A N E S E M M T V M E
M Z M X O I A M G P L S O G R P U I O A U
J S V A D M N C O R L A O R K A P L C N I
B C I A G I I T I N A R C L O R Y O A I L
W Y N L O N A U C A D M I U M B Y X C U O
T A T B I S E Z M Y I V O I D I J P R M D
V E I T S C A S H M U I T E N H C E T S P
Z U Z I R C O N I U M X N O Q D V W L O I
M O U I V I E N R U R U I Z N L I T X R N
F M P L N O U U E J M V C A I Y B U O P Q
W J E V N C O M U I D O S S L T S N M I X