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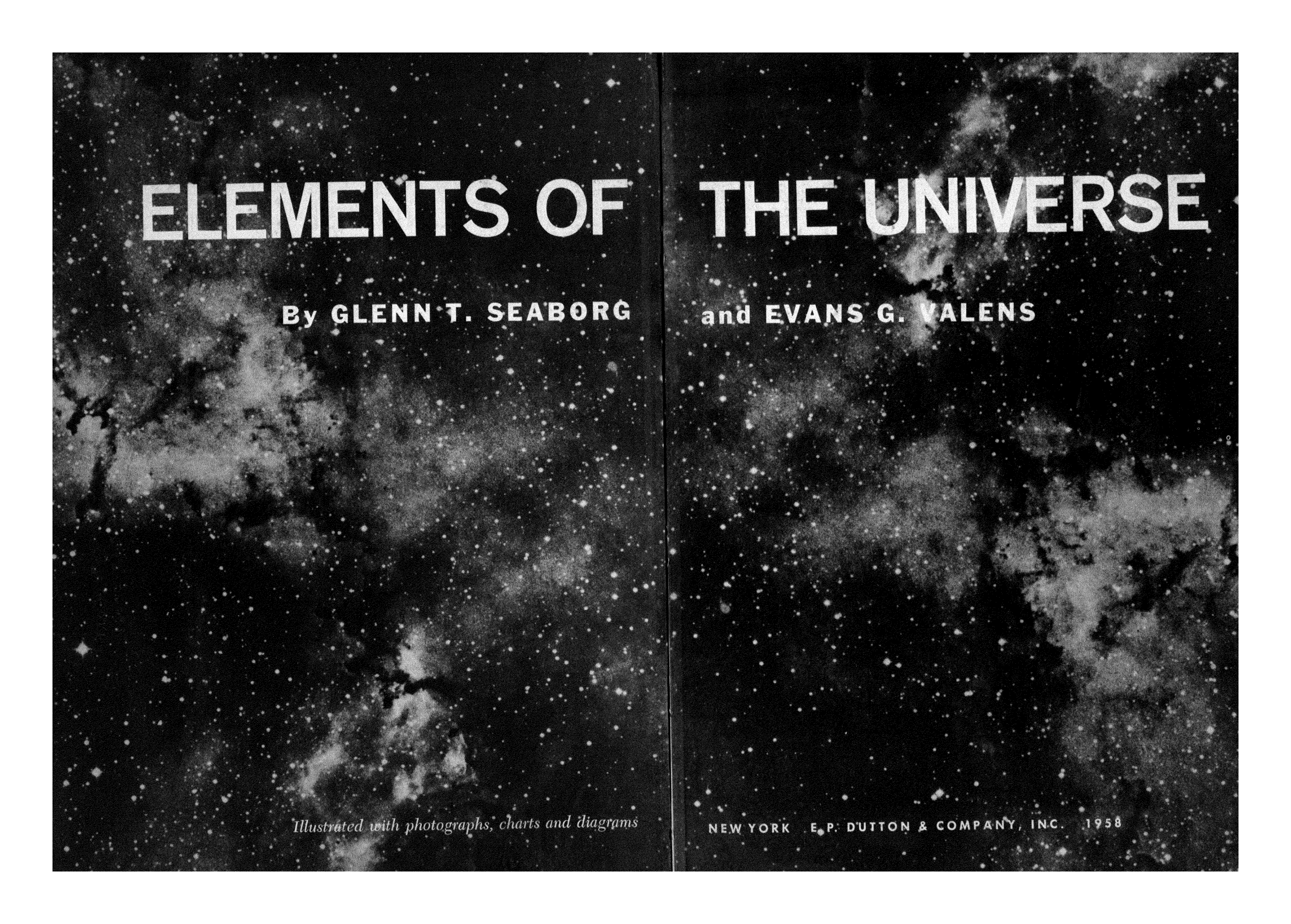
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ELEMENTS OF THE UNIVERSE



ELEMENTS OF

By **GLENN T. SEABORG**

THE UNIVERSE

and **EVANS G. VALENS**

Illustrated with photographs, charts and diagrams

NEW YORK E. P. DUTTON & COMPANY, INC. 1958

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To our children

Peter, Lynne, David, Stephen and Eric Seaborg

Tom, Marc, Jo Anne and Dan Valens

PUBLISHER'S NOTE

Glenn T. Seaborg, Nobel Laureate chemist and Chancellor of the University of California at Berkeley, is the co-discoverer of nine of the 102 chemical elements. He is the only man since Dmitri Mendeleev to have made a major change in the Periodic Table of the Elements—a change which led immediately to the discovery of the first element heavier than plutonium and subsequently to the discovery of another half dozen transuranium elements.

This book is the outgrowth of a series of ten half-hour films on the chemical elements made by educational television station KQED, in San Francisco, for the Educational Television and Radio Center. The series, featuring Dr. Seaborg, was filmed in the "cave room" of the University of California's Radiation Laboratory in Berkeley, California.

The "cave room," an underground section of the nuclear chemistry building high in the Berkeley hills, is the alchemist's dream come true. Here, with remote control equipment operating through heavy lead shielding, many of the synthetic elements known to the world today have been isolated and identified.

The authors are indebted to the staff of the Radiation Laboratory and in particular to Senior Chemist Bernard G. Harvey and to Daniel M. Wilkes, who were responsible for research and script writing in connection with the film series.

FOREWORD

We all know, of course, the important part that science plays in our defense effort, but perhaps we don't realize how important it is in our everyday peacetime lives.

Our entire economy is coming to depend upon continued technological advance. Eighty thousand employees of the General Electric Company, to mention one outstanding example, work today on products that did not exist before the end of World War II.

We also know of important medicines which science has given the physician, with which we are able to go a long way toward eradicating disease and prolonging life.

In spite of all these advances, we have a serious shortage of scientists and engineers in our country today. And as a country, one of the most critical problems we face is that of interesting more and more young people in science as a career.

The responsibility for this problem rests largely on the shoulders of scientists themselves. It is not enough to emphasize the importance of advanced intellectual training; we must also present a clearer picture of what a scientist actually does and try to convey some appreciation for the fun, the thrill, and the inner satisfaction which a scientist usually experiences in his work. And this fact is a primary motive for this book.

Personally, we feel that the population as a whole should learn more about science. In fact we feel that science should be a part of the repertoire of a cultured man today. We need great leadership in all fields by men who are cognizant of the values of our civilization and the factors which influence it. The liberal education that prepares men for such leadership must include science as an integral part, for science is too central a part of our modern culture to be ignored.

Naturally, no one book could accomplish the broad objective of communicating both the excitement and the satisfying logic of science. We have chosen one segment of the broad spectrum of science, the chemical elements, and we hope to make a sound beginning by saying something about these essential building blocks of nature.

GLENN T. SEABORG
EVANS G. VALENS
Berkeley, 1958

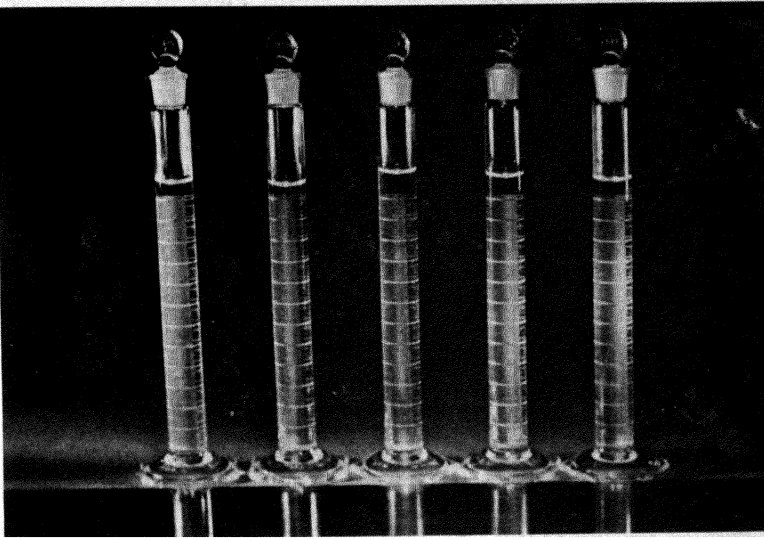
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ELEMENTS OF THE UNIVERSE

THE ATOM



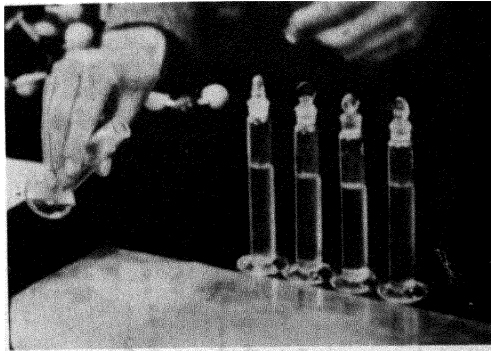
Here, in test tubes, are five colorless liquids.

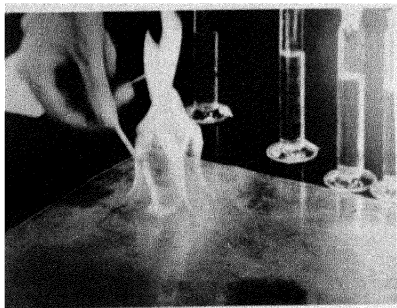
They look alike. They might be water.

But it would be as dangerous to drink one as it would be to play Russian Roulette with all but one chamber of your gun loaded.

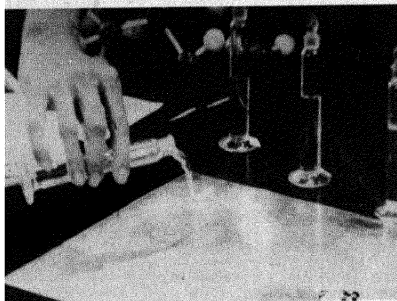
If we pour the first liquid onto a copper plate and put a match to it . . .

nothing happens.

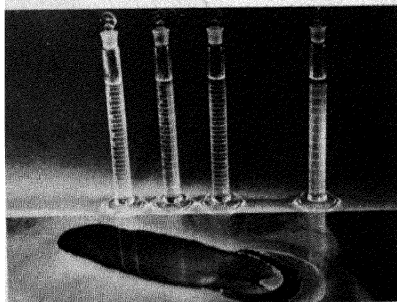




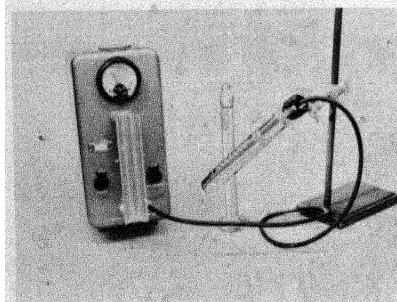
The second . . .
burns fiercely.



The third . . .
quenches the flame.



The fourth . . .
etches the copper.

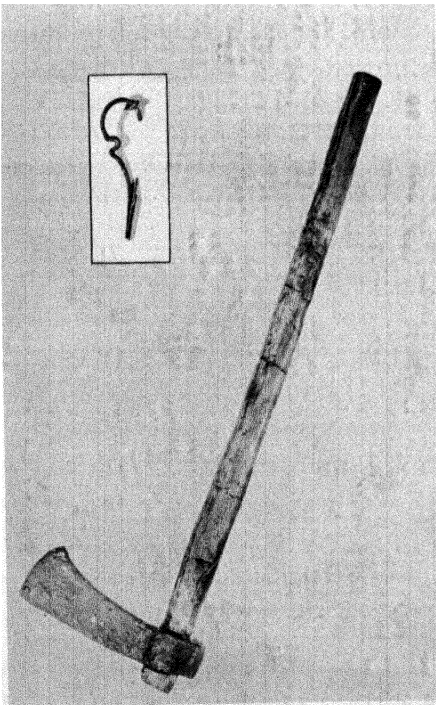
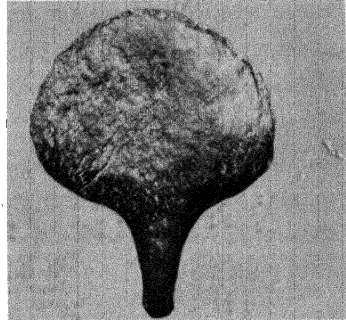


The fifth . . .
we dare not pour,
for it is highly
radioactive.

How can five liquids look so similar and yet react so differently?

We shall have a partial answer to the question as soon as we can discover just what an element is. Actually, the liquids are simple compounds, none involving more than three or four elements.

They are simpler, chemically, than wood and stone, the first materials man learned to re-shape for his own ends.



In the Bronze Age, long after the days of the stone ax, man began to use distinct elements to fashion weapons, tools, and even safety pins.

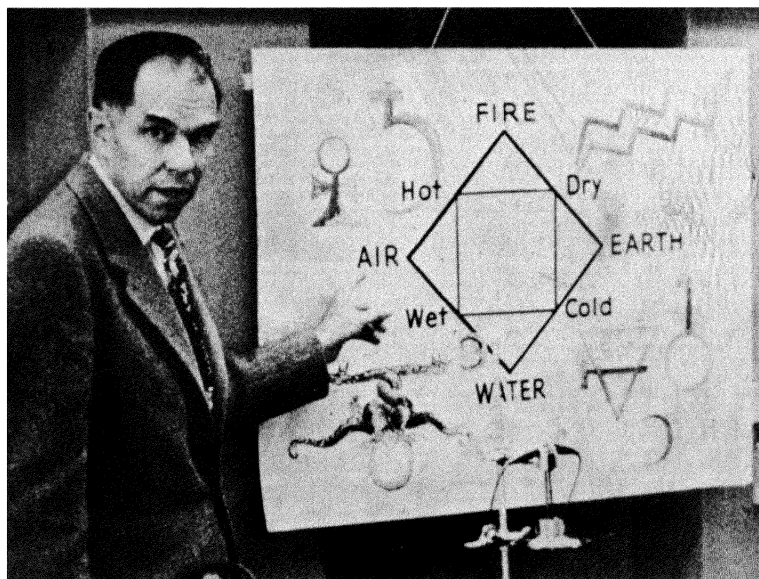
By the time of the Iron Age, he had become something of a master at extracting metal from ore and beating it into useful forms like this iron ax-head.

Texts written in Biblical times tell us that men were then working with gold, silver, copper, iron, tin, carbon, lead, sulfur, and mercury—without, of course, understanding that these were chemical elements.

In the Middle Ages, the alchemists tried to make some orderly sense out of the fragmentary and primitive knowledge they had of chemistry.

Although crude and visionary, the best of them were certainly “scientists.” They wanted to discover what the fundamental building blocks of matter were, and they performed experiments trying to find out.

They thought the basic elements were Fire, Earth, Air, and Water, and their first “Table of Elements” strikes us today as rather amusing. It even indicated what the alchemists thought were chemical reactions. For example, wood produced heat when burned. This proved—they believed—that wood (dry) is made of earth (ash) and fire.



H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	A
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac-103	(104)	(105)	(106)	(107)	(108)										

Lanthanides	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinides	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	E	Fm	Mv	102	(103)

Today, we know that everything on earth and in the universe is made up of the chemical elements. We know pretty well what an element is, and our "Periodic Table of the Elements" looks somewhat different from that of the alchemists.

It packs into one chart most of the secrets the alchemists spent their lives searching for in vain, a vast amount of information that any man today can discover as he learns to read it. It lists over a hundred elements instead of just four. It tells the weight and structure of each element and shows at a glance which elements are similar to one another.

With its help, we too can describe the nature of fire, earth, air, and water. Fire is the light and heat given off when certain other elements combine with oxygen. Earth is a complex combination of many dozens of elements. Air is a mixture of at least eight elements, plus the compound carbon dioxide. Water is a compound of two elements, oxygen and hydrogen.

Our current Table of Elements shows all the elements listed in order of their atomic numbers. Each name is abbreviated to a chemical symbol for convenience, as on pages 242-243.

The first eight elements, for example, are H (hydrogen), He (helium), Li (lithium), Be (beryllium), B (boron), C (carbon), N (nitrogen), and O (oxygen).

The symbols for some of the older elements are derived from former names which are no longer used—for instance, hydrargyrum for mercury (Hg) and argentum for silver (Ag). The origin of all the elements' names and symbols is given in the appendix.

On the Periodic Table, the number in the upper left corner of each rectangle is the element's atomic number.

For carbon, this is 6, and it tells us that a carbon atom has 6 protons in its nucleus. This also means that each carbon atom has 6 electrons, and this in turn indicates how carbon will combine—or refuse to combine—with other elements.

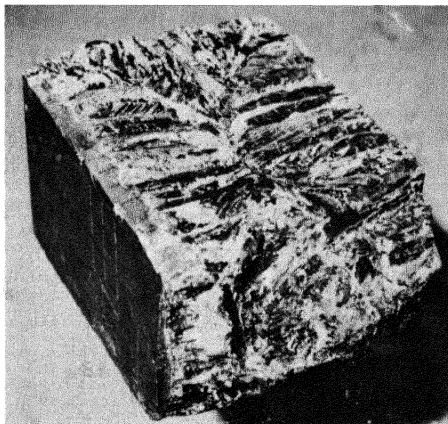
The number at the bottom—slightly more than 12 for carbon—is the weight of the element's atom as compared with the number 16, assigned to oxygen. An oxygen atom is thus almost sixteen times as heavy as a hydrogen atom.

The two numbers together indicate the composition of the element's nucleus—6 protons and a total of 12 protons-plus-neutrons for carbon.

6	Carbon
C	
12.011	

WHAT IS AN ELEMENT?

An element is a substance made up entirely of atoms of one single kind. This chunk of bismuth, for example, contains only bismuth atoms, all of them chemically identical. We can break it in two, smash it with a sledgehammer, and then grind it to dust. Each piece, each



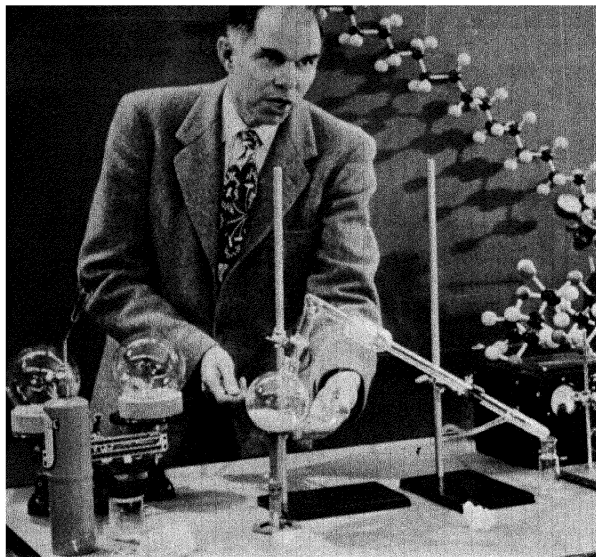
separate flake of dust, is still bismuth. We can heat it until it flows like water, or until it boils and every atom flies away into the air as a gas. We still have identical bismuth atoms and nothing else.

Most atoms can combine with others to form molecules.

An atom of one element may combine with other identical atoms, as when two oxygen atoms form an oxygen molecule. An atom may also combine with one or more atoms of other, different, elements to form molecules or chemical compounds.

An astonishing thing about chemistry is that an element usually becomes completely unrecognizable when combined with another element. A highly flammable gas, hydrogen, combines with oxygen to form the chemical compound, water.

Two poisons, chlorine gas and the soft, silver-white metal, sodium, combine to form table salt.

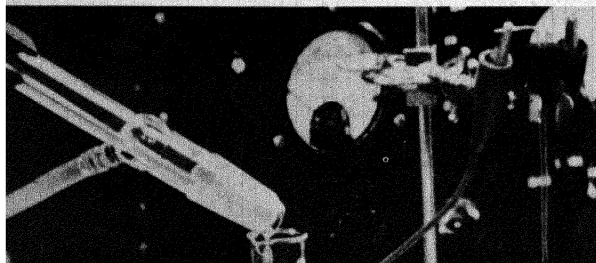
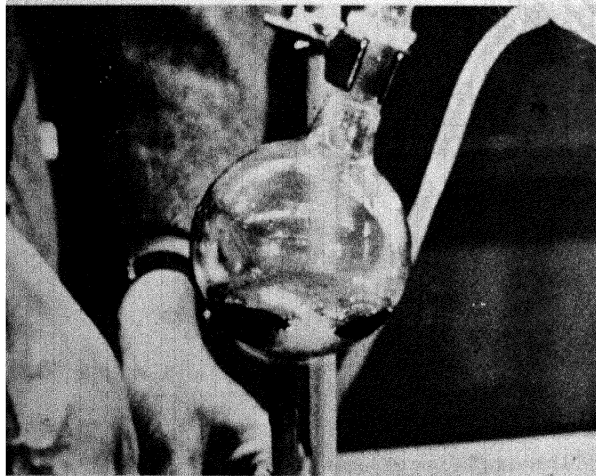


Sugar is another common compound, and we can rip apart its molecules rather easily. All we do is put some sugar in a retort and heat it.

Soon the molecules of sugar begin to break down. The black residue shows that sugar is composed partly of carbon. The other atoms which had been tied up in the sugar molecule recombine, steam off, and condense in droplets which spill into a beaker. These drops prove to be water.

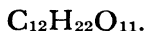
We can go further by passing an electric current through the water with an electrolysis apparatus. This breaks down the water molecules into hydrogen and oxygen, which come off separately as bubbles.

Sugar, then, is made of the elements carbon, hydrogen, and oxygen. Each sugar molecule contains 12 atoms of carbon,



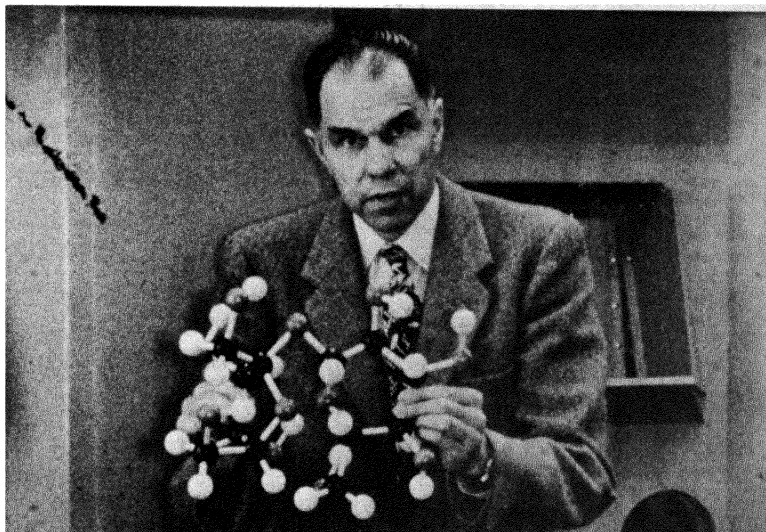
22 atoms of hydrogen, and 11 atoms of oxygen.

So the chemical formula for sugar is simply:



We were handling many billion billions of molecules of sugar in the retort, but we can imagine what must have happened to each molecule individually.

The easiest way to imagine this is to make a model.

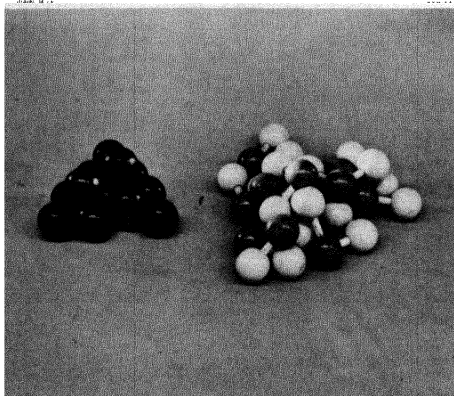
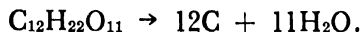


The black balls represent carbon atoms. The white ones represent hydrogen atoms and the gray, oxygen. The sticks indicate the chemical bonds which hold the separate atoms together in the molecule.

This of course is not what a sugar molecule would look like if we could see one, but it is a good way of showing the general arrangement of the atoms within the molecule.

When heated, the molecule breaks down. The 12 carbon

atoms are left behind and the 11 molecules of water boil off as steam. To abbreviate this result, we write:

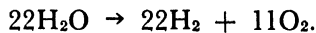


In other words, 1 molecule of sugar yields 12 atoms of carbon plus 11 molecules of water.

Finally, when the water molecules are split up, they release 22 atoms of hydrogen and 11 atoms of oxygen.

These combine in pairs, giving 11 molecules of hydrogen (H_2) and $5\frac{1}{2}$ molecules of oxygen (O_2) for each sugar molecule we've broken down.

We can write this reaction and make it come out in whole numbers by doubling the number of water molecules involved in the formula:



In another experiment, we can heat the compound mercuric oxide, a red powder. The name tells us what elements the compound is made of—mercury and oxygen.

First the mercuric oxide changes color. Then it boils and begins to steam up inside the flask.

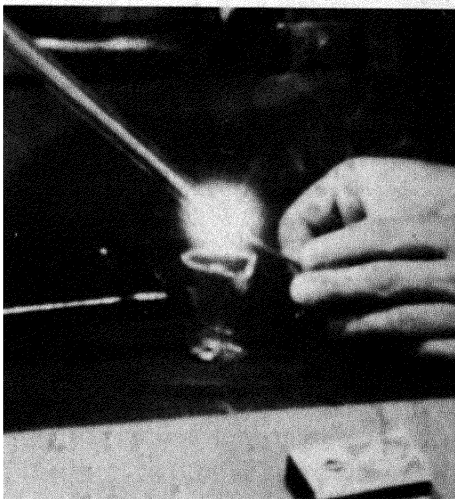
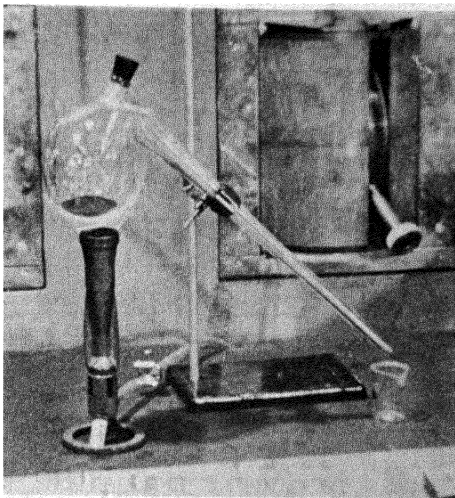
These gases cool as they are forced out of the neck of the retort.

The mercury forms into drops and falls into the beaker.

The oxygen also should be flowing out of the retort. We can't see it, but we can test for oxygen by holding out a glowing splint of wood. Fed by the nearly pure oxygen, the wood bursts into flame.

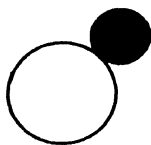
So we find that a dry red powder is made of a shiny liquid metal and a gas that ably supports a flame.

A molecule of mercuric oxide is far more simple than



a molecule of sugar. It consists of just 2 atoms, one of mercury (which has the symbol "Hg") and one of oxygen. This we could tell easily from the formula for mercuric oxide: HgO.

A diagram of the HgO molecule might look like this.



The white circle represents an atom of oxygen and the smaller black circle an atom of mercury. A model of the whole experiment would show the mercuric oxide molecules dashing around when heated, slamming into one another and into the sides of the retort.

Each molecule at length breaks apart, the oxygen atom springing away from the mercury atom. The single mercury atoms are forced out the neck of the retort as a gas and condense into droplets as they cool. The oxygens combine in pairs and escape as a gas.

The process can be written quite simply with chemical symbols:

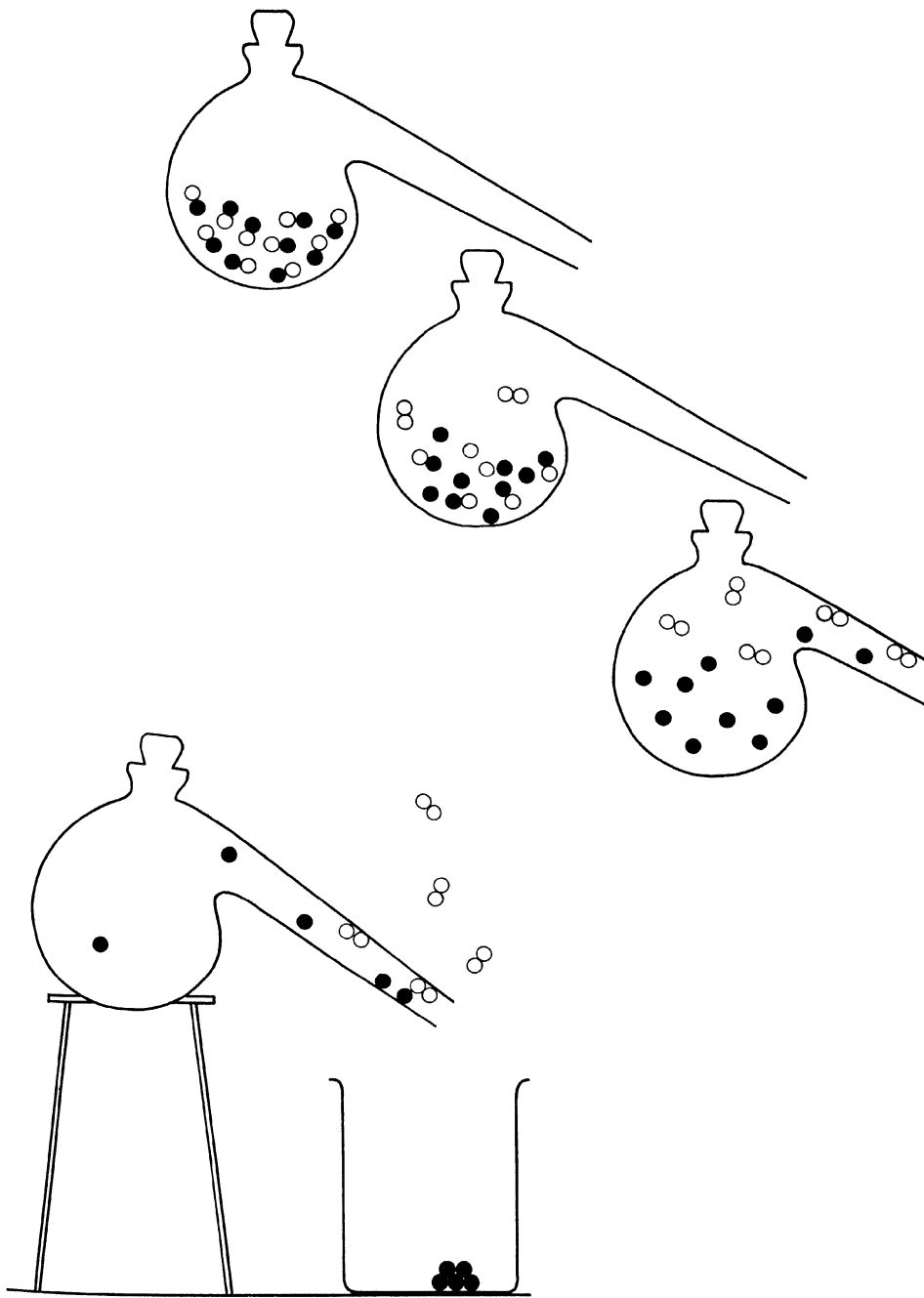


However, because oxygen molecules consist of pairs of atoms, the symbol for free oxygen—that is, for the oxygen molecule—is "O₂." So we can change the formula by considering 2 molecules of mercuric oxide instead of 1:



This means: 2 molecules of HgO yield 2 atoms of mercury and 1 molecule of oxygen, which of course contains 2 oxygen atoms.

Mercuric oxide then, like sugar, is a compound—a collection of similar molecules, each of which is made of atoms of more than one kind.



The five clear liquids we tested earlier were also simple compounds, although the last was a solution of two compounds.

All of the five liquids, incidentally, were made from gases only or from gases and a solid. In order, they were:

Water (hydrogen and oxygen)

Acetone (hydrogen, oxygen, and carbon), which is flammable.

Carbon tetrachloride (carbon and chlorine), which douses fire.

Nitric acid (hydrogen, oxygen, nitrogen), which underwent a chemical reaction with the copper plate.

Solution of cobalt 60 (water plus a nitrate—nitrogen and oxygen—of radioactive cobalt), which causes the Geiger counter to click away like a lawn mower.

To make the five different liquids, a total of six elements were used. But with only two of these—hydrogen and carbon—thousands of different compounds can be formed. Paraffin, gasoline, and plastics, for example.

The differences between all such “hydrocarbons” depend only upon the relative number of carbon and hydrogen atoms in each molecule and the particular way in which the atoms are joined together.

The chemical formulas for different hydrocarbons reveal at once the number of hydrogen and carbon atoms in each molecule. CH_4 is methane. Some others are:

C_2H_2 acetylene

C_2H_4 ethylene

C_8H_{18} octane

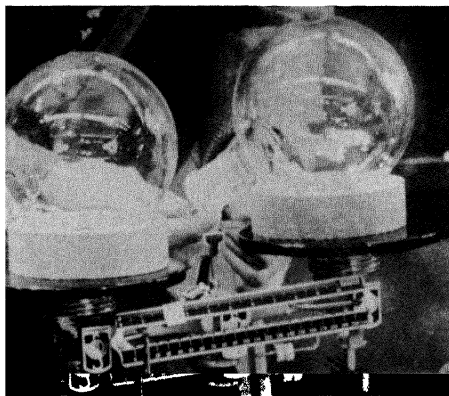
$\text{C}_{10}\text{H}_{18}$ decahydronaphthalene

Of course it takes many quadrillions of molecules to make a visible or weighable amount of a hydrocarbon or any other compound; the size of a carbon atom, for instance, is on the order of a hundred-millionth of an inch.

The atom's nucleus is still much smaller than this, the diameter of an atom being about ten thousand times that of its nucleus.

If we imagine a carbon atom enlarged to the size of a football stadium, then the electrons would be like flies flying around the outside of the stadium and the nucleus would be the size of a golf ball in the center and thousands of times as heavy as all the flies put together.

This means, of course, that more than 99.9 per cent of everything in the universe is concentrated in the tiny nuclei of atoms. And it means that an atom is mostly empty space.



The weight of an atom is something we can show with very little trouble if we take a gas, because equal volumes of all simple gases and vapors contain the same number of molecules.

If we put a flask on each pan of a balance, with the necks

pointed down, we are balancing one liter of air against another.

Now if we pour hydrogen up into one of the flasks, it will fill the flask and shove the air out. This tips the balance, showing us how much heavier a liter of air is than a liter of hydrogen.

We could translate this into terms of single molecules—by means of some very long division—because we know the number of molecules in each flask is the same, namely about

26,870,000,000,000,000,000.

The number of atoms is double this figure since free nitrogen and oxygen (the chief components of air), as well as hydrogen, have 2 atoms per molecule.

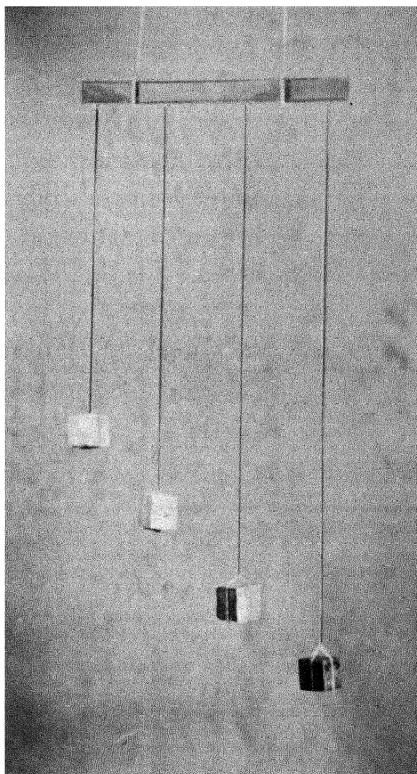
Equal volumes of some metals also reflect the relative weight of their atoms. Cubes of magnesium, iron, lead, and uranium, for instance, will hang at different heights when attached to similar springs.

This, however, only *suggests* the relative weights of atoms. It is more difficult to weigh the atoms of a solid because we also have to contend with how tightly or loosely the atoms are packed in different materials. The actual

weights—as compared with a weight of 16 given for oxygen—are listed in the Periodic Table. They are approximately 238 for uranium, 207 for lead, 56 for iron, and 24 for magnesium. A uranium atom, therefore, weighs about as much as 238 hydrogen atoms or 15 oxygen atoms, and this is one of the many pieces of information so economically contained in the Periodic Table of the Elements.

The Table also supplies the basic facts we need to explain how atoms are put together.

We shall start with the first element, hydrogen, because its nucleus is a proton. And protons are the basic structural particles in the nuclei of all other elements.



Every material object in the universe is made from one or more of the hundred-odd elements. So we can rightly call the elements the building blocks of matter.

But all elements are made from the same fundamental particles—protons, neutrons, and electrons.

What makes one element different from another is nothing more than the specific number of protons, neutrons, and electrons involved.



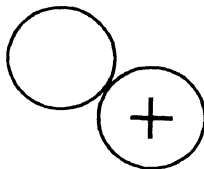
A proton is a particle with a single charge of positive electricity. It is at the same time the nucleus of a hydrogen atom, and it accounts for more than 99.9 per cent of this atom's weight.

1	Hydrogen
H	
1.008	

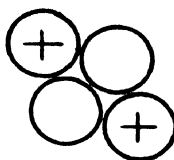
So the hydrogen nucleus—a proton—has a charge of one, as shown in the Periodic Table by hydrogen's atomic number. We can represent it by a plus sign inside a circle.

If we add to this proton a second particle of the same weight but without any electric charge, the combination will have a weight of about two. Its charge will still be 1, since there is only 1 proton.

The neutral particle, a neutron, appears in the nucleus of every element except common hydrogen.



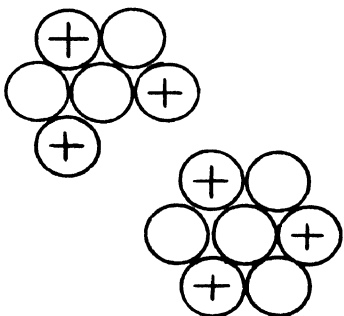
If we now add a second proton and a second neutron, we have a compound particle with a charge of two and a weight of four.



As we might easily guess by referring to the Periodic Table, this is the nucleus of the helium atom. It agrees with helium's atomic number (2) and with its mass number (4). A "mass number" is simply the total number of protons plus neutrons, or an element's atomic weight rounded off to the nearest whole number.

2	Helium
He	
4.003	

If we add a third proton and a third neutron to a helium nucleus, we have the nucleus of a species of lithium called lithium-6. This nucleus has a positive charge of three and a mass number of 6.



The more common form of this silver-white metal is lithium-7, which has one more neutron.

Ninety-two per cent of natural lithium is lithium-7, with an atomic weight of 7.02 and a mass number of 7.

The atomic weight of natural lithium (containing both lithium-6 and lithium-7) is 6.940.

By adding more protons and more neutrons to this nucleus, we can build up the nuclei of heavier and heavier elements.

3	Lithium
Li	
6.940	

But an atom isn't an atom unless the nucleus is surrounded by the appropriate number of electrons—one electron for each proton.

An electron is a single charge of negative electricity. It is much lighter than a proton, but it exactly balances the proton's positive charge.

Thus a hydrogen atom consists of 1 proton with a single electron circling it. The electron is represented here as a small minus sign. The drawing is not very accurate since the electron would orbit in a circle a half mile in diameter if the proton were really as large as we've drawn it.

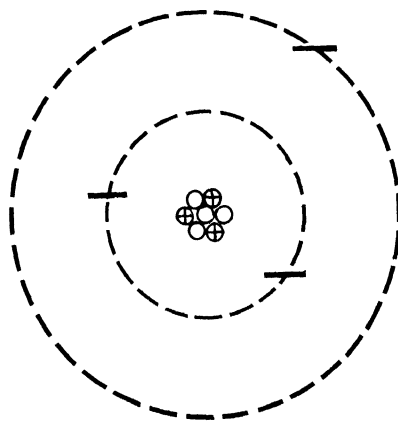
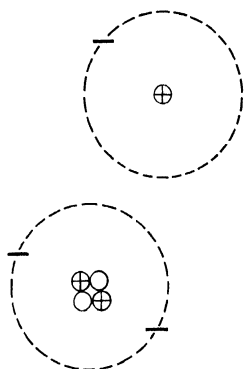
A helium atom requires 2 electrons to balance the positive charge of two in its 2 protons.

A peculiar fact about atoms is that only so many electrons will fit into any one particular orbit or "shell."

The orbit nearest the nucleus will hold only two electrons. After that, the front row is sold out, so to speak, and the next orbit—which can hold eight electrons—starts filling up.

The lithium atom, therefore, has 2 electrons in the first orbit, like helium, plus a third electron circling alone in an outer orbit.

Actually, the orbits are three-dimensional with the electrons apparently traveling around the surface of an imaginary sphere or shell, and the actual paths are much more fuzzy than our simple description here implies.



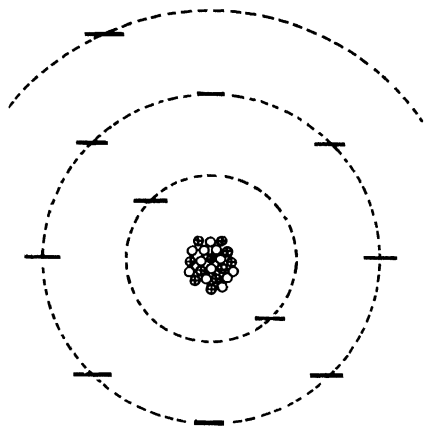
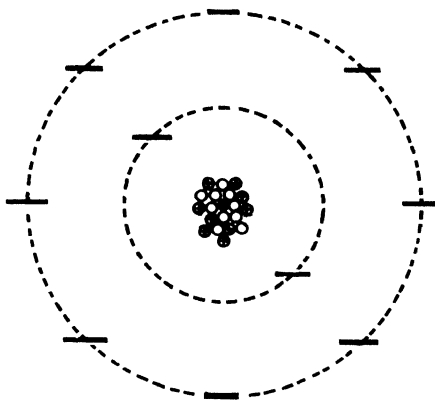
When we consider the next seven elements, we find that the second shell is progressively filled until we get to neon, with 10 protons in the nucleus and 10 electrons.

In a neon atom, the second shell with its eight available places is completely filled. The inner, two-electron shell is also filled, of course. So neon, like helium, is a rather self-contained atom without either unoccupied places or extra “dangling” electrons.

After neon comes sodium, with 11 protons and 11 electrons. The eleventh electron orbits all alone in a third shell.

Sodium is like lithium and hydrogen—which are listed directly above it in the Periodic Table—in that it has a single, “dangling” electron in its outermost shell.

Here we begin to see why we say “*Periodic* Table of the Elements.” Certain properties are found to recur periodically when we list the elements in order of their atomic numbers.



The important thing in chemistry is to remember that every time a nucleus is increased by one proton, then you have to have one more electron—usually in an outer orbit—in order to keep the whole thing electrically neutral.

So the number of protons in the nucleus determines how many electrons there are circulating around the nucleus. There must always be the same number in a neutral atom.

And it is the number of these electrons and the way they are arranged in a limited number of possible orbits that determine the chemical properties of any particular element.

A “chemical property” simply means the way in which the element will combine with—or refuse to combine with—other elements.

This includes what elements it can combine with, how readily it will join up, and how stable the combination will be.

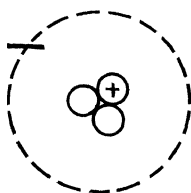
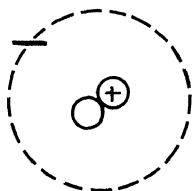
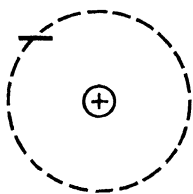
These chemical properties depend upon the number and arrangement of an atom's electrons and have nothing to do with the protons or neutrons in the nucleus itself.

The number of neutrons in an atom's nucleus does, however, account for different isotopes of one and the same element.

The word *isotope* means occupying the same or equal (*iso*) place (*topos*).

Isotopes of an element appear at the same place in the Periodic Table and have the same number of protons and electrons. A good example is natural lithium, which contains 2 isotopes, lithium-6 and lithium-7.

The main differences between isotopes are those of atomic weight and—often—of radioactivity.



If we take a hydrogen atom (a proton with an electron circling it) and add a neutron to the nucleus, we don't have to add any more electrons. The neutron has no charge and doesn't need an electron to balance it. And since we aren't changing the number of electrons, we don't change the chemical properties of the atom either.

What we do change is the weight of the atom. In the case of hydrogen, there are 3 isotopes: ordinary hydrogen with only a proton in the nucleus; deuterium, which also has a neutron and which has, therefore, a weight of about two; and radioactive tritium, which has a weight of about three and has 1 proton and 2 neutrons in its nucleus.

Natural uranium is mostly U-238 with 92 protons and 146 neutrons in its nucleus. But it also contains another well-known isotope, U-235, the fissionable source of atomic energy. U-235 also has 92 protons but only 143 neutrons.

Its weight, of course, is three units less than that of U-238.

This business of isotopes means that we weren't completely accurate when we defined an element as a substance all of whose atoms are identical.

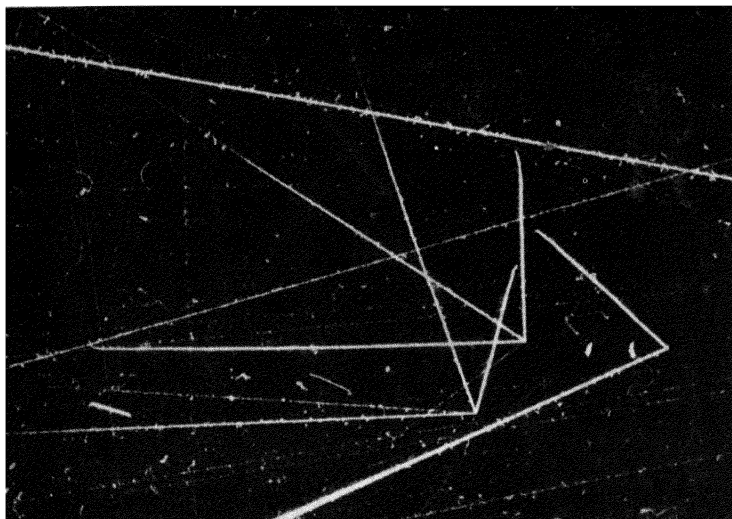
The nucleus of an atom is three-dimensional, not flat as we have pictured it. Some nuclei are roughly spherical, like a basketball. We believe that others, such as uranium, are more like footballs.

ATOMS INTO MOLECULES

How do we know that atoms really exist?

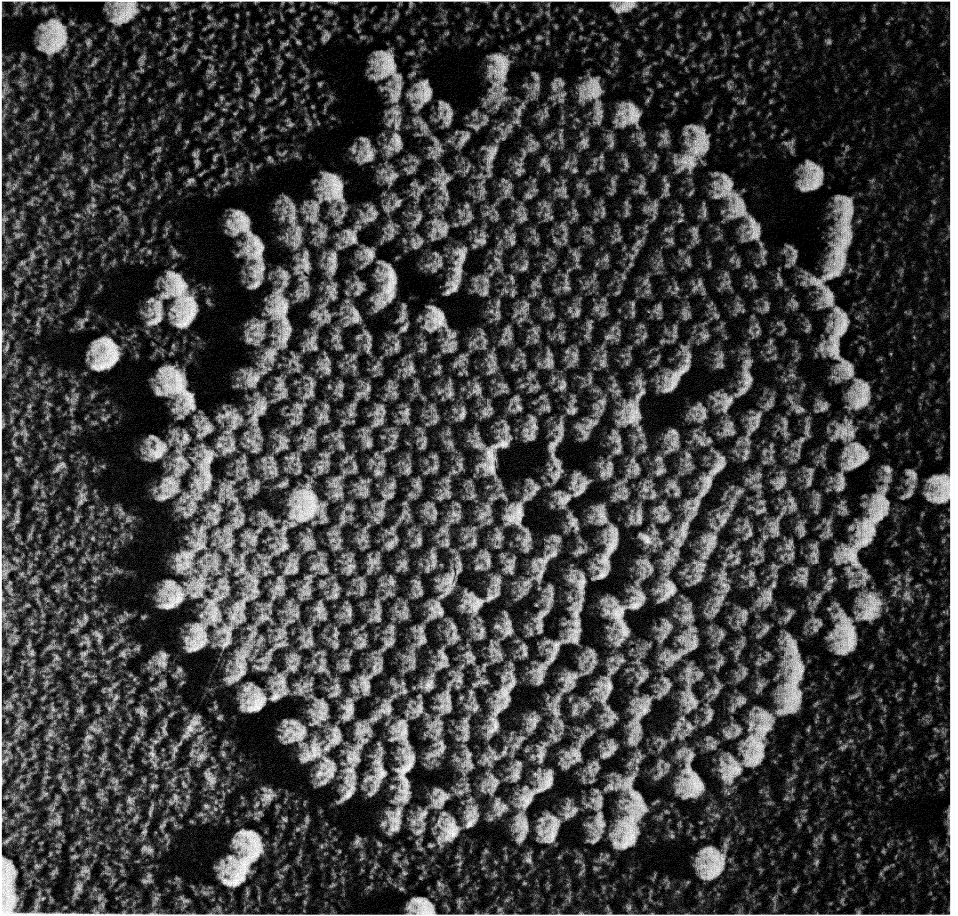
We've talked about them in great detail, but they are too small to be seen with a light microscope or even with an electron microscope. However, we can observe readily some very direct evidence of their existence.

When a charged particle such as the nucleus of a hydrogen or helium atom travels through a moist gas, it leaves behind a vapor trail very much like that of a miniature airplane.



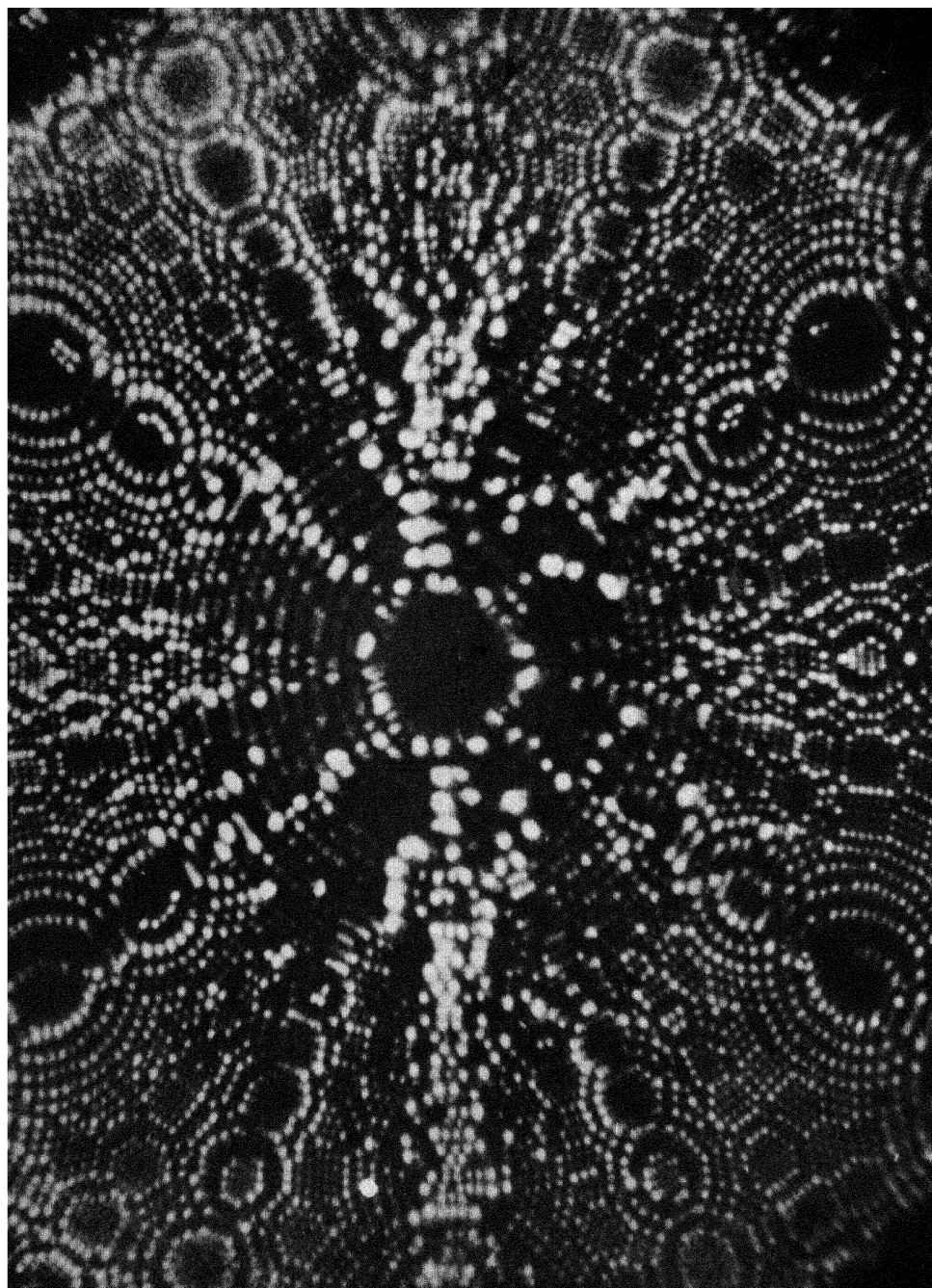
This trail we can see or photograph even without a microscope.

Another way to get close to seeing an atom is to look at a molecule, which is a chemical combination of atoms. Occasionally we find molecules so big that they will register in an electron microscope even though they are too tiny to reflect the longer waves of visible light.



A virus is just such a huge molecule, one of the largest single molecules known. One unfortunately common type is the poliomyelitis virus, a spherical molecule containing many thousands of atoms. It can be seen clearly with an electron microscope, magnified here 180,000 diameters.

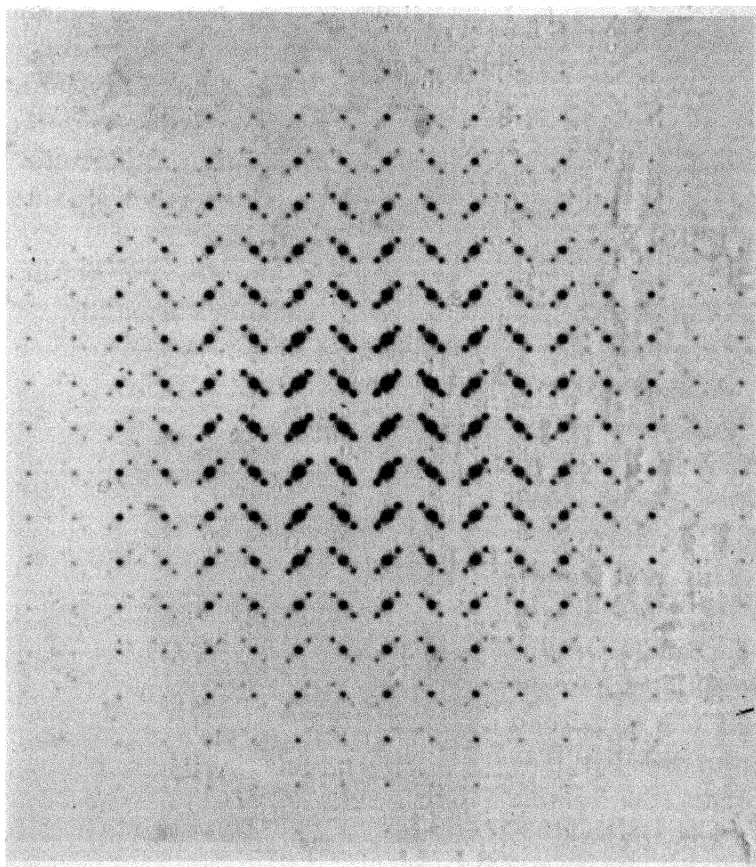
In 1957, Dr. Erwin W. Mueller of the University of Pennsylvania produced the first actual photograph of individual atoms. On the following page are tungsten atoms in a regular arrangement forming a crystal lattice on the surface of a fine metal needle. These are seen with Dr. Mueller's field ion microscope. Each small dot represents a single atom. The bright dots are clusters of several atoms. Magnification is about 2,000,000 diameters.



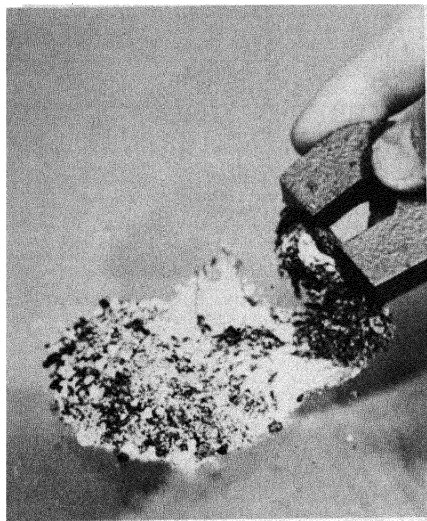
At the Massachusetts Institute of Technology, Dr. Martin Buerger has used X-rays to record the position of single atoms in a metallic crystal of pyrite, or "fool's gold." Pyrite is iron disulfide, a compound of iron and sulfur. Each pyrite unit contains 1 iron atom and 2 sulfur atoms.

An atom, of course, is extremely small. One atom of iron has a diameter of less than a hundred-millionth of an inch.

This photograph may not show just what an atom "looks like," but it does show clearly the position of atoms in the crystal.



Now how do atoms of different elements such as iron and sulfur manage to combine to form molecules?



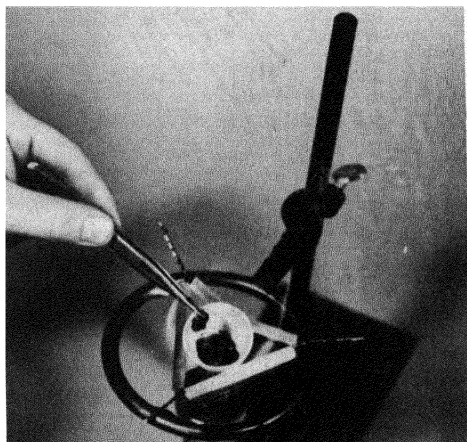
If we thoroughly mix iron filings with sulfur, the result still looks like iron and sulfur. Furthermore, the two are easy to separate again. All we need is a magnet.

So iron and sulfur mixed do not make a compound. They form only a mixture, with individual iron and sulfur atoms merely lying side by side.

If we put the mixture in a crucible, however, and heat it, the iron atoms and sulfur atoms do combine chemically.

Heated together, they each lose their identity and become a compound, iron sulfide, which is different in all its properties from either iron or sulfur.

This iron sulfide, or FeS , has 1 iron atom and 1 sulfur atom in each molecule. It is quite similar to iron disulfide (pyrite), the formula for which is FeS_2 .



It is even easier to make a compound from magnesium because we have to provide only a few magnesium turnings and heat. The necessary oxygen will come from the air.

We can put the magnesium on a balance just

to prove that when burned it gains weight instead of losing it.

What happens is that the magnesium combines with oxygen from the air to form the compound, magnesium oxide, or MgO .

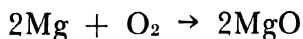
This is heavier than the original magnesium, of course, because we've added to it the weight of the oxygen atoms. And the resulting magnesium oxide doesn't look in the least like either magnesium or oxygen.

The combining of iron and sulfur or of magnesium and oxygen is called a chemical reaction.

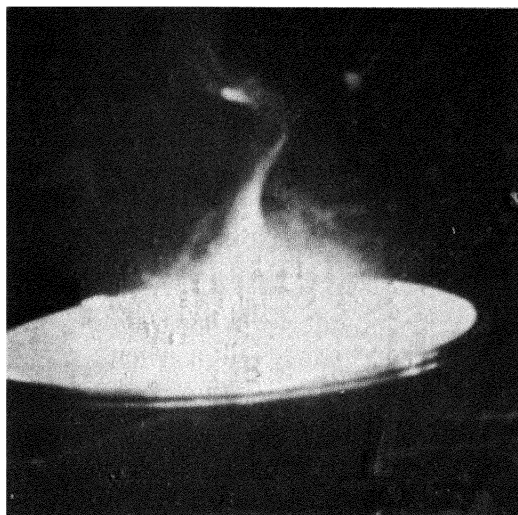
The chemical formulas for both reactions are very simple:



(iron plus sulfur yields
iron sulfide)



(magnesium plus oxygen yields
magnesium oxide)



So far, we've seen how a chemist actually goes about making a compound from two elements, but we haven't seen why the reaction occurs.

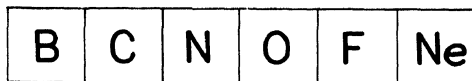
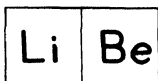
There are many ways in which atoms may combine to form molecules, but it always involves, in one way or another, the rearrangement of the atomic electrons in their orbits.

It is the rearrangement of these electrons, in fact, that defines chemistry.

The first horizontal row of the Periodic Table contains only two elements, hydrogen and helium. Each has a single occupied electron shell.



The second horizontal row contains eight elements, and every one of the eight has two occupied electron shells. We could say that each of the eight has an outer shell with eight "places" and that one or more of these "places" is occupied by an electron.



We've already seen that the first element in this row, lithium, has only 1 electron in this outer shell. Beryllium has 2, and on up to neon with 8.

A similar situation applies to the other five horizontal rows in the Periodic Table. The first element in each row is, in a sense, beginning a new shell; and in this shell it has just a single electron.

H
Li
Na
K
Rb
Cs
Fr

Any vertical column in the Periodic Table shows us at a glance all the elements with the same number of electrons in the outermost shell.

Hydrogen, lithium, sodium, potassium, rubidium, cesium, and francium each has a single electron in its outer shell.

On the other end of the Table are the elements which have all their shells completely filled: helium, neon, argon, krypton, xenon, and radon.

Elements in any one vertical column are all members of the same family. Because they all have the

He
Ne
A
Kr
Xe
Rn

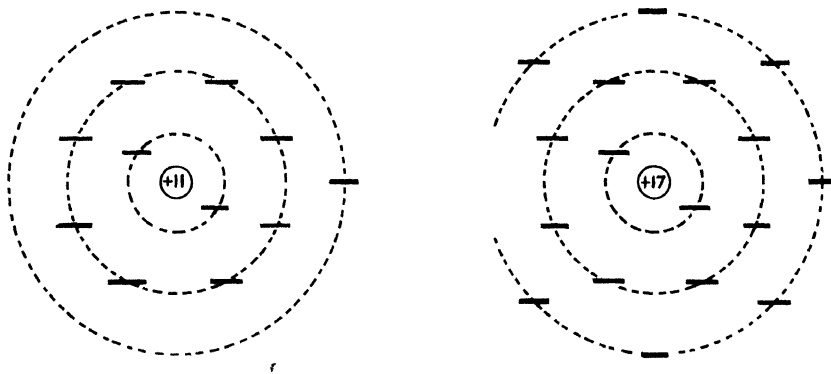
same number of electrons in their outer, or “exposed,” shells, they all have similar chemical properties.

Except for hydrogen itself, the elements in the first column—the chemical “cousins” of hydrogen—are called the alkali metals. Each one has a single electron available to be removed in chemical reactions.

Such a reaction occurs when sodium combines with chlorine to form molecules of common table salt.

A two-dimensional model of a sodium atom would show 11 positively charged protons in the nucleus and 11 electrons to balance it—2 in the first orbit, 8 in the second, and 1 in the third.

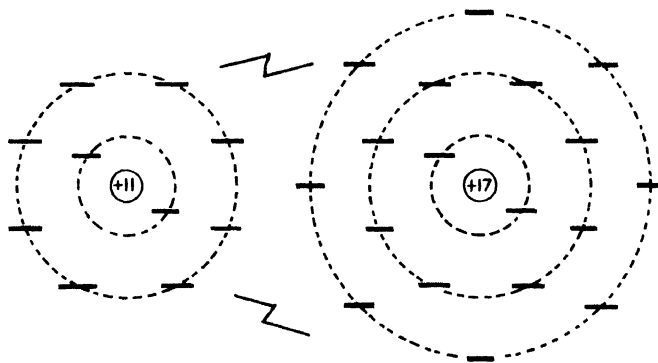
The chlorine atom has 17 electrons—2, 8, and 7. The sodium has an “extra” electron while the chlorine is shy one—the perfect set-up for a cooperative get-together.



The chemical combination occurs when the single electron from the sodium hops across into the gap in the outer ring of the chlorine atom.

The sodium, robbed of an electron, is left with a positive charge. One negative charge (the electron) is added to the chlorine. The two atoms **now** have an opposite charge and are thus held together by a strong electrical bond.

We should say that the two *ions* have an opposite charge, since atoms which have lost their neutrality by gaining or losing electrons are called ions.

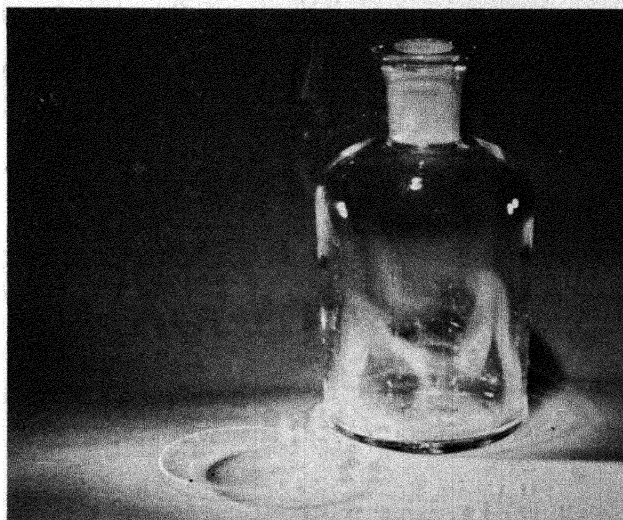
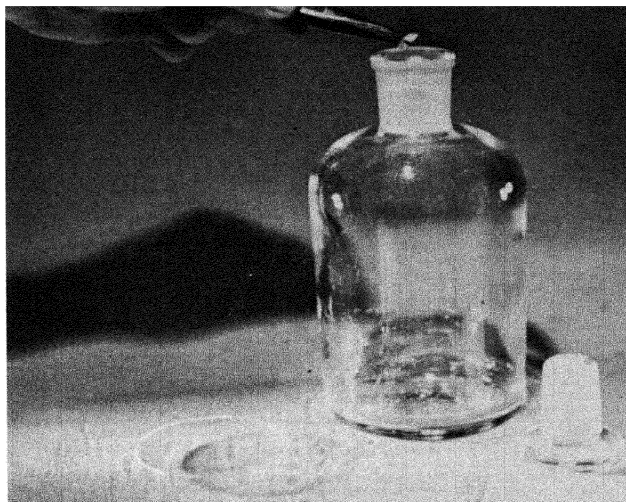


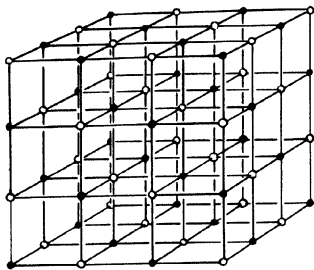
To actually perform this experiment, we can take a flask of chlorine, which is a yellowish, poisonous gas, and drop into it a piece of the soft, poisonous metal, sodium.

After a short while, enough of the sodium and chlorine will have combined to produce an edible hunk of table salt.

Molecules of salt or anything else, of course, must themselves come together in some manner to form visible amounts of material.

In the case of salt, the sodium and chlorine grow together as crystals. The sodium ions and the chlorine ions are stacked alternately in a cubic lattice which might look something like a min-

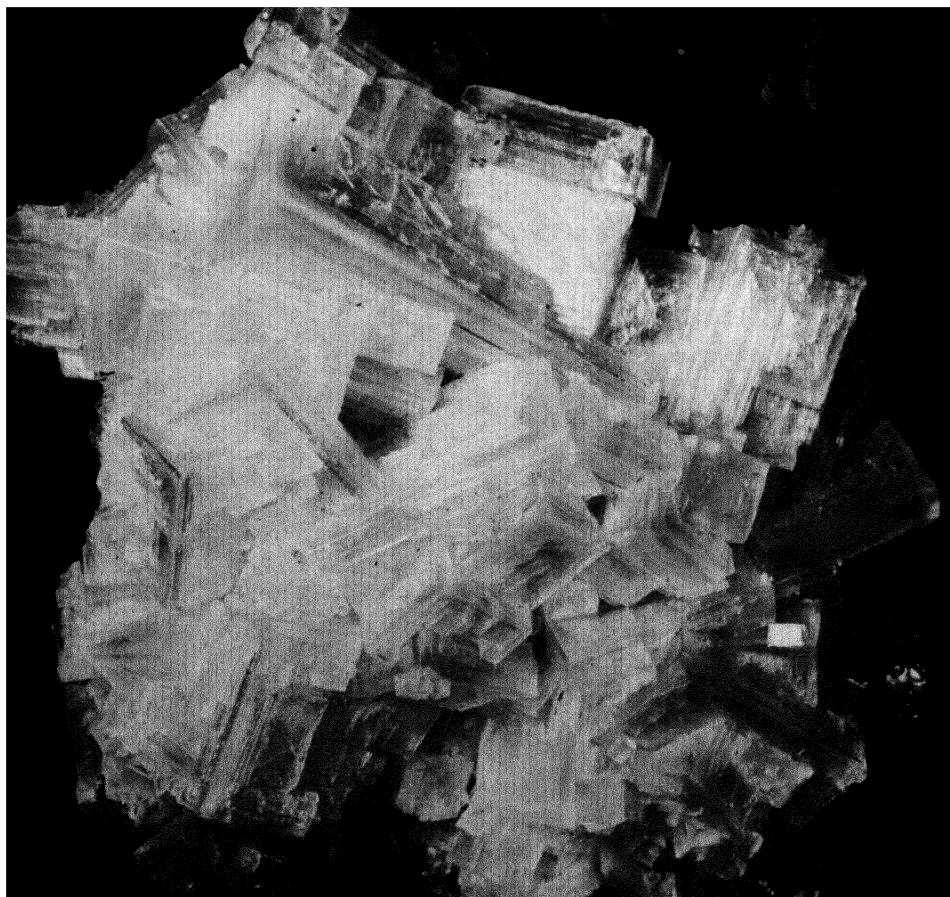




ature, microscopic Jungle-Gym.

Salt crystals usually are imperfect and broken, but the cubic structure is obvious in crystals which have grown undisturbed. The complex crystal below contains a rather large number of atoms—roughly 10^{25} atoms of sodium plus the same number of chlorine atoms.

This number, when written out, looks rather impressive:
10,000,000,000,000,000,000,000,000.

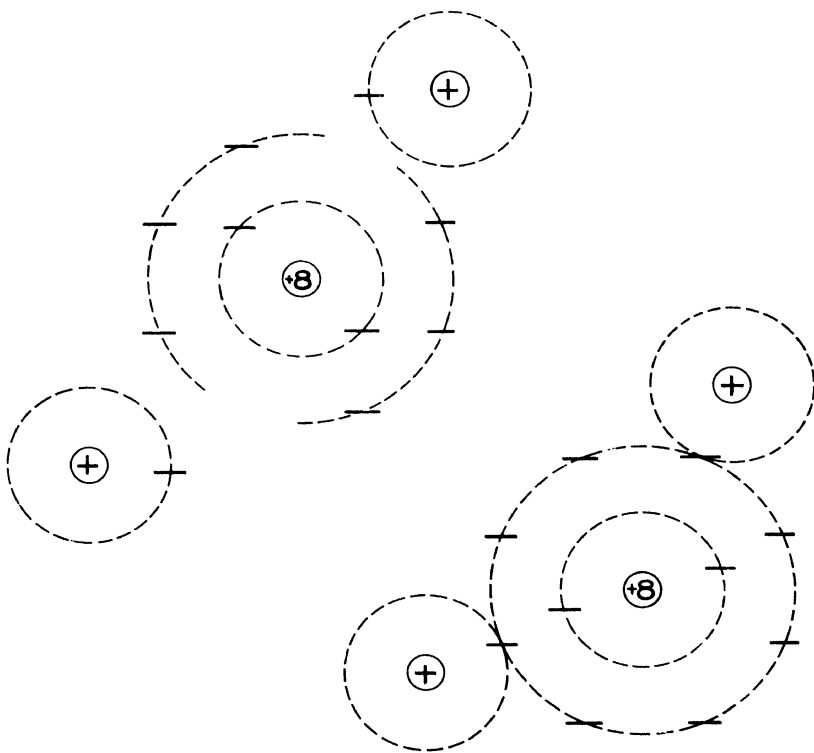


Salt illustrates only one of the ways in which atoms combine to form molecules, and perhaps not even the most important way.

A different example is water.

We start with 2 hydrogen atoms—each with a single electron—and 1 atom of oxygen. The oxygen has 8 protons in its nucleus and, therefore, 8 electrons. Two of the electrons are in the inner shell and 6 in the outer shell, leaving two gaps into which the electrons from the two hydrogen atoms might presumably fit.

But in this case, the electrons are not transferred, as with the sodium and chlorine. Instead, the two hydrogen atoms sort of join the oxygen atom and share their electrons with it.



The joining of atoms into molecules in this particular manner is known by several names: shared electrons; electron pair bonds; co-valences. A great many molecules—those of sugar, for example—are formed in this way.

Such molecules have tiny electric currents in them, constantly changing direction. This causes the molecules to be attracted to each other and to cling together to form visible amounts of water, sugar, or other material.

If there were no such force holding molecules together, then all molecules would drift independently, like air, and every substance we know would be gaseous.

FIRST ELEMENTS

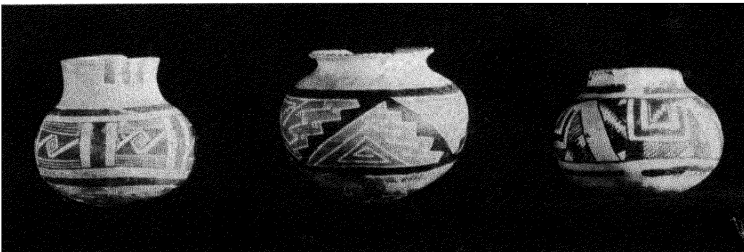
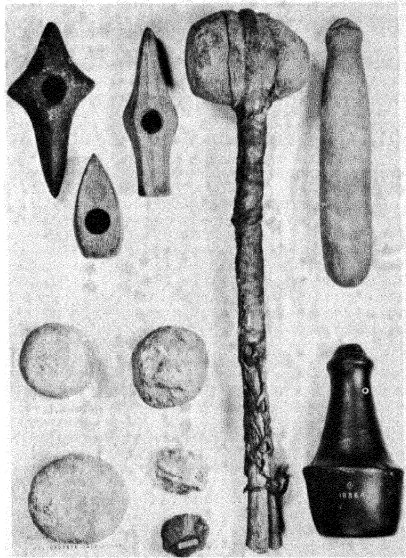
How and when were the chemical elements discovered?

The story of the elements begins in ancient times. When man first discovered fire, the carbon from burned wood was left in the forests. Man's first art work was likely drawn with charcoal on the walls of caves.

In the Stone Age, tools and weapons were carved from stone: spearheads, axheads, knives. Early Indians developed remarkable skills in reshaping natural materials. Their clay bowls were made largely of compounds of aluminum, silicon, and oxygen.

Of course, no one at the time had any idea what an element was or that clay and stone contained a number of separate ingredients.

As time went on and man began to master his environment, it became possible to rearrange and extract elements from the materials that men found in the earth.



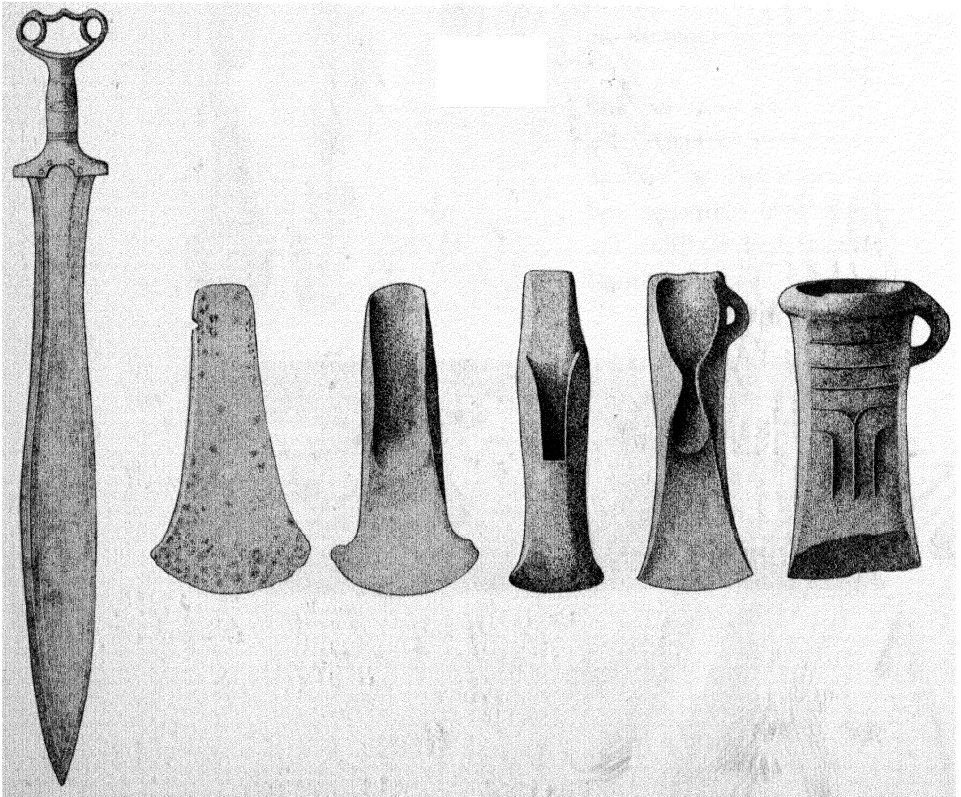
This "rich earth" is now known as ore.

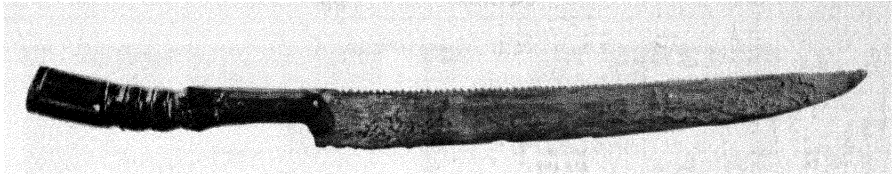
Galena, or lead sulfide, is a common ore, and early man must have refined lead from galena by a process that was essentially accidental. Lead ore mixed with carbon in the heat of a campfire resulted in droplets of pure lead metal.

Another ore known to early man was cinnabar, or mercuric sulfide. This ore, when heated, underwent a chemical reaction that resulted in pure mercury.

As man's curiosity and his ability to handle materials grew, he discovered free copper and learned to extract both copper and tin from their ores. His mixing of copper and tin to make bronze was such a significant step that we speak of this period of pre-history as the Bronze Age.

This period produced excellent tools and weapons and highly ornamental personal jewelry, and it marked the beginning of the science of metallurgy.



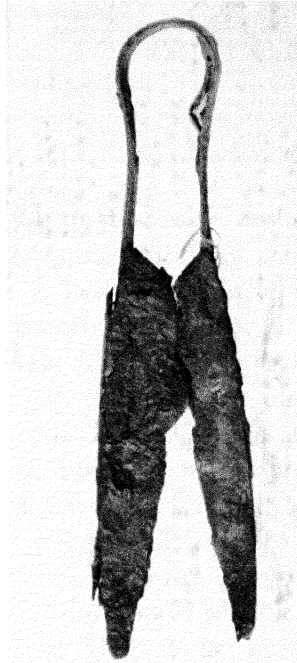


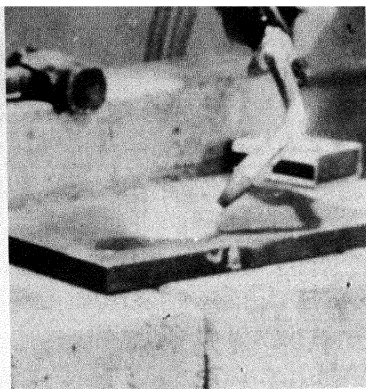
The Iron Age began about a thousand years before Christ with the discovery of iron smelting.

Actually, iron must have been discovered and rediscovered many times before this. It was perhaps first found in the ashes of a big fire built against a red, ore-bearing rock.

Iron was fashioned into hammers and awls, keys and combs, and, of course, into weapons. The rise and fall of civilizations during these times was distinctly related to progress in metallurgy among the artisans of various nations.

The main point is that man had learned how to extract these various elements from their natural environment, from their ores. The method was basically a crude one and involved only the use of heat and, in some cases, carbon. It is a method quite possible in a campfire and, of course, rather easy to demonstrate in a laboratory.





It is only a matter of putting an ore—a lead ore, for example—on a graphite plate and heating it. The result is a relatively pure piece of lead.

Once the metal was extracted, or discovered in the pure state as with gold, ancient man quickly found out that he could shape it to various forms. He learned to pound out metals and even beat them into leaf-thin sheets.

Early man, then, learned to handle quite a number of the individual chemical elements, although he, of course, did not recognize them as elements.

Naturally, he had carbon in the form of charcoal.

He also knew about sulfur and about the elements he found free in nature—gold, silver, and copper.

He learned to extract the pure metal from ores of copper, mercury, lead, and tin.

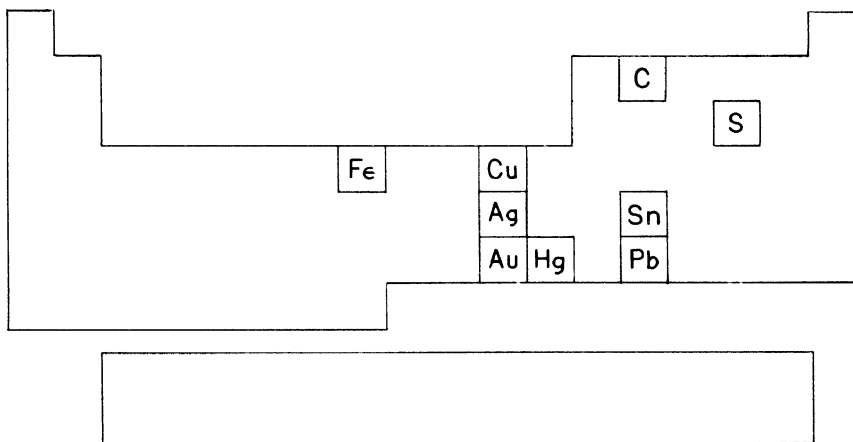
Probably his main achievement was the ability to produce metallic iron from its ores, and this accomplishment among certain peoples determined to some degree where the centers of civilization existed during the early days of metallurgy.

By the time of Christ, these nine chemical elements were known to man and had been isolated and put to use in deliberate ways.

Located on our modern Periodic Table, several of them prove to be closely related chemically. Copper, silver, and gold all have similar properties. So also do tin and lead.

The nine, with their chemical symbols, are:

C (carbon)
S (sulfur)
Fe (iron)
Cu (copper)
Ag (silver)
Sn (tin)
Au (gold)
Hg (mercury)
Pb (lead)



FROM ALCHEMY TO CHEMISTRY

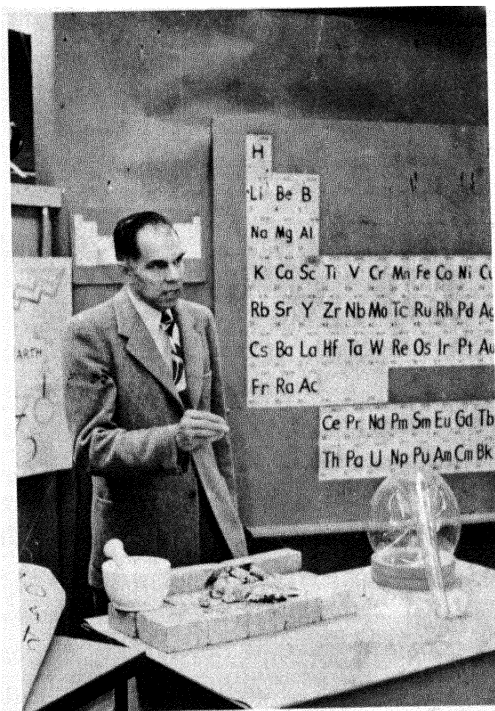
Nothing more was done about actually isolating chemical elements until the period we crudely define as the Middle Ages. And at that time the alchemists came into being.

They worked with crude equipment, including the retort and the mortar and pestle which still have symbolic meaning for us today.

The alchemists did a number of experiments. These ranged from

such intimations of magic as the search for the elixir of life to experiments that are the forerunners of modern chemistry.

Alchemists spoke often of the "philosopher's stone" by which they hoped to transmute baser metals into gold. We're not quite sure what this illusory substance was. It probably



was not any one thing, or even a stone. Some historians have suggested that it was mercuric sulfide, but different people apparently thought of it in different ways.

Aside from such trivial endeavors, the alchemists performed a number of significant chemical experiments. They extracted metals from ores, for instance, although this was not unusual in view of earlier developments in metallurgy.

More important was their production of a number of the acids which much later formed a basis for industrial chemistry.

One of their experiments consisted in heating a substance like iron sulfate and expelling what they called oil of vitriol. This we know as sulfuric acid.

The alchemists also were able to produce muriatic acid (another name for hydrochloric acid) and nitric acid. And they made chemicals like potash and sodium carbonate, which later proved to be commercially important.

However quaint some of their aims and methods may have been, the alchemists were noteworthy because they were interested in both theory and practical experiment. The knowledge they gleaned from experimenting they tried to systematize with a chart like that on page 18. They thought that the elementary substances of nature were fire, earth, water, and air, and they tried to establish logical relationships among these four "elements." And in a way their rather fanciful chart was a forerunner of our modern Periodic Table.

A great deal of credit is certainly due the alchemists. They made a lot of discoveries and they managed to pin down three important chemical elements during the twelfth, thirteenth, and fourteenth centuries.

These elements are arsenic (As), antimony (Sb), and bismuth (Bi). They are all members of the same chemical "family" and are placed—on our modern Periodic Table—in

33	Arsenic
As	
74.91	
51	Antimony
Sb	
121.76	
83	Bismuth
Bi	
208.99	

the same vertical column.

This similarity among the three elements indicates that the crude chemical methods of the alchemists were probably geared to one particular type of experiment in which chemical properties of a certain type were important.

After this trio, no new elements were discovered for several centuries—with the possible exception of platinum, which had been isolated in Mexico some time before the middle of the sixteenth century. The name comes from the Spanish and means “little silver.”

In the eighteenth century, the only recorded use for platinum was in the adulteration of gold coins. For a few years in the early nineteenth century, Russia minted platinum coins.

Of the 13 elements known by the middle of the seventeenth century, none of the discoveries is mentioned in recorded history. This also applies to zinc, which was recognized as a distinct substance in the late 1600's or possibly earlier.

But about this time, science had begun to assume a fairly modern form. People began to study nature and chemistry and the elements for their own sake, for the knowledge that could be had. New discoveries were recorded and published.

It is true that the ancient Greeks were concerned with knowledge for its own sake. They even had a well-developed atomic theory which was in many ways similar to atomic theory today. But the Greeks were not fond of doing experimental work, so their theory just stayed on the books and was never developed.

The first element that was discovered by one man and is really recognizable as his work was phosphorus, which means "bearer of light."

It was discovered by an alchemist and merchant named Hennig Brand while he was searching for the "philosopher's stone" in Hamburg, Germany, in 1669.

Brand made phosphorous from dried urine but kept the process a secret. He discovered that the new substance had the remarkable property of glowing in the dark when exposed to air. He had a lot of fun playing with it and showing it to his friends, and he likely made some money out of it on the side, too. It was later recognized as an element and named.

Cobalt was discovered in 1737 and nickel about 14 years later. Cobalt and nickel ores previously had been mistaken for copper ore. Since they refused to yield copper, these ores obviously were possessed by evil spirits and thus had gained the names "Kobold" (goblin) and "Kupfernickel" (Old Nick's copper). And the names persisted.

Next to be discovered was hydrogen.

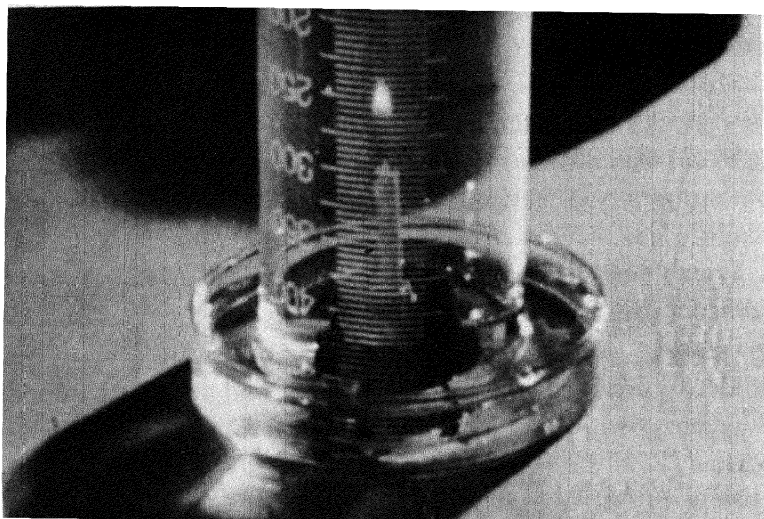
Hydrogen is easily made by putting a metal into an acid solution—hydrochloric acid, for example. The bubbles produced in this reaction are bubbles of hydrogen.

The fact that metal in an acid gives off bubbles had been known for a long time, but nobody had clearly recognized that the gas involved was different from any other gas.

It was Henry Cavendish who studied the properties of the gas produced in this way and, in 1766, gave an accurate description of it. When it was later discovered that burning the gas produces water, it was named "hydrogen" or "water-maker."

In the 1770's a lot of people began to play with the nature of ordinary air, trying to find what it was made of.

Daniel Rutherford discovered that burning or breathing uses up a part of any given amount of air. For example, if you light a candle and trap a certain amount of air over it,



the candle will burn for a while and then go out. The burning uses up part of the air, and the candle refuses to burn in what is left. If you replace the candle with a mouse, the mouse will use up part of the air and then die.

What Rutherford did was to investigate the gas left after the candle went out or the mouse stopped breathing. He found that this gas was different from ordinary air. Nothing would burn in it; animals could not live in it.

So Rutherford is credited with the discovery of nitrogen.

A number of others were doing similar work at the same time, notably Cavendish, Joseph Priestley and Karl Scheele. But Rutherford was first to describe nitrogen accurately.

THE PERIODIC TABLE

During the next twenty-five years, eleven more elements were discovered: chlorine (Cl), manganese (Mn), molybdenum (Mo), tellurium (Te), tungsten (W), zirconium (Zr), uranium (U), titanium (Ti), yttrium (Y), chromium (Cr), and beryllium (Be).

Also, the electric battery had been discovered by the Italian physicist Alessandro Volta.

Early in the nineteenth century, the English chemist Humphry Davy was experimenting with a great big battery and a compound which we know as potassium hydroxide. This compound was well known, but no one fully understood what it contained.

Davy discovered he could make a new metallic element by melting it and passing an electric current through it.

To repeat this experiment today we would simply melt potassium hydroxide in a metal crucible attached to one terminal of a power supply. When a platinum wire from the other terminal is dipped into the molten compound, a small quantity of potassium metal forms around the end of the wire.

A few days after discovering potassium, Davy tried the same experiment using sodium hydroxide instead of potassium hydroxide.

So Davy is credited with the discovery of both of these metallic elements, potassium and sodium.

Between 1800 and 1869, the pace of chemistry was stepped up and the number of elements known to man was nearly doubled. Scientists of all nationalities during this time were involved in the discovery of a rash of new elements.

These elements, in order of their discovery, were:

As late as 1817, when some fifty elements were known, no one had attempted to classify them or arrange them in any particular order. In fact, chemical elements as such had just been clearly distinguished from chemical compounds.

There was a distinct need for some kind of classification of the known elements, and scientists soon began to work with the problem of making order out of the situation.

They worked much in the following way: they began to play with the idea that perhaps different elements were related in terms of their atomic weights.

The first man to discover a promising relationship among several elements was the German chemist Johann Wolfgang Döbereiner. In 1829, he ran into the idea of triads.

He found that if you listed elements with similar properties—like lithium, sodium, and potassium — one above the other, the atomic weight of the middle one was halfway between that of the one above and the one below. Furthermore, the chemical properties of the middle element were a mean between those of the other two.

Other examples were calcium - strontium - barium and chlorine-bromine-iodine.

In the following twenty-five years, other chemists expanded Döbereiner's triads and came out with groups of four or five related elements.

3	Lithium
Li	
6.940	
11	Sodium
Na	
22.991	
19	Potassium
K	
39.100	

20	Calcium
Ca	
40.08	
38	Strontium
Sr	
87.63	
56	Barium
Ba	
137.36	

Li		N	O	F
Na	Mg	P	S	Cl
K	Ca	As	Se	Br
	Sr	Sb	Te	I
	Ba	Bi		

These were important steps in building a system.

In 1862, the French chemist A. E. B. de Chancourtois plotted the elements along a spiral, according to their atomic weights. This brought elements with similar properties together, and each pair of these corresponding elements differed by 16 (the atomic weight of oxygen) in their atomic weights. De Chancourtois suggested that the properties of the elements are related just as whole numbers are related.

Two years later, J. A. R. Newlands, in England, listed the first seven elements as hydrogen, lithium, beryllium, boron, carbon, nitrogen, and oxygen, like the notes in a musical scale. Each of the seven stood at the top of a group of elements with similar properties. Newlands called his arrangement of elements into seven groups the "law of octaves."

The real breakthrough came in 1869 when the German chemist Lothar Meyer and the great Russian chemist Dmitri Mendeleev uncovered the principle of the Periodic Table.

First, they listed all the known elements in the order of their increasing atomic weights.

Hydrogen didn't fit in very well, and they started lining up the elements with lithium and beryllium. They found that if they finished a row and then started a second row below it, they got elements with similar chemical properties falling one below the other clear across the table.

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

As they extended the table, they found several groups of elements that did not fit their seven categories, and these elements were incorporated into the table later.

Mendeleev in particular noticed that he had to skip sev-

Li	Be	B	C	N	O	F			
Na	Mg	Al	Si	P	S	Cl			
K	Ca		Ti	V	Cr	Mn	Fe	Co	Ni
Cu	Zn			As	Se	Br			
Rb	Sr	Y	Zr	Nb	Mo				

eral places in order to maintain vertical columns in which all the elements listed had similar properties. This particular form of the Periodic Table is different from the later, more extended form shown on pages 81 and 85.

It was Mendeleev's big contribution that he recognized the gaps in his Periodic Table and stated that they should be filled by elements as yet undiscovered.

He went further than that. He had the courage to stick his neck way out and predict what the undiscovered elements would look like, how much they would weigh, and how they would react chemically.

Since he expected the three elements to have properties similar to those of boron, aluminum, and silicon, respectively, he tentatively named them eka-boron, eka-aluminum, and eka-silicon.

In the case of eka-silicon, for example, Mendeleev predicted that it would be a dirty-gray solid, would have an atomic weight of 72 and a density of 5.5, and would form a liquid chloride.

Specific elements never had been predicted before. If any one of this trio should be discovered, it would prove for all time the value and power of Mendeleev's particular system of arranging the chemical elements.

Throughout the rest of the nineteenth century, this arrangement was based upon

	B	C	
	Al	Si	
	eka-boron	Ti	
	eka-aluminum	eka-silicon	
	Y	Zr	

the relative weights of successive elements.

As it turned out, the correct way to list the elements depends upon their atomic numbers rather than upon their weights.

The atomic number of an element represents the number of positive electrical charges on the element's nucleus. In general, elements with a larger charge also have a greater atomic weight than those preceding them in the Periodic Table; but this is not always so, as in the case of cobalt and nickel.

The fact that the positive charge of the atom resides in a dense, tiny nucleus was discovered in 1911 by the English physicist Lord Ernest Rutherford.

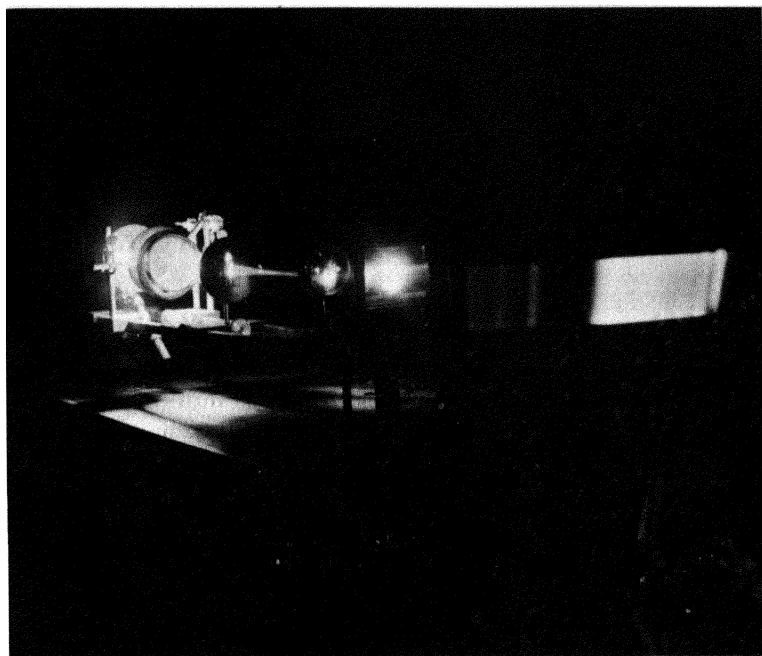
Within the next two years, the Danish physicist Niels Bohr worked out the details of atomic structure and the orbits of various groups of electrons around this nucleus.

In 1913 and 1914, Henry Moseley, in England, clinched the concept of atomic number as referring to the positive charge on the nucleus.

With this new concept in hand, it was possible to review the Periodic Table and explain some of the apparent discrepancies.

Long before this, however—even while Mendeleev was still working out his Periodic Table—the discovery of other elements was made possible by the development of a marvelous new instrument, the spectroscope.

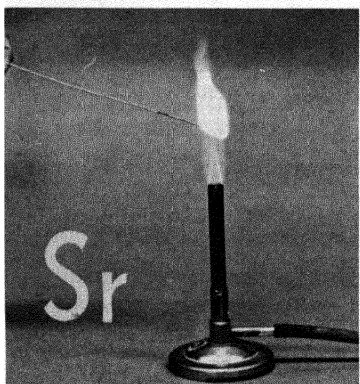
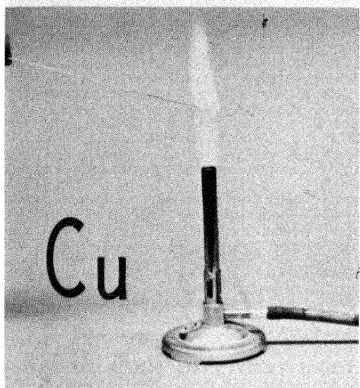
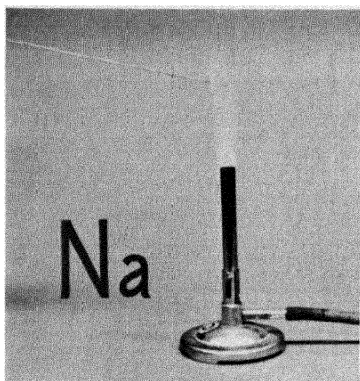
"FINGERPRINTING" BY SPECTROSCOPE



The spectroscope is one of the most powerful tools of laboratory science. Its purpose is to dissect light. This it does by spreading out light from a particular source as raindrops spread out the colors of sunlight in a rainbow. The instrument makes use of a prism or a diffraction grating, of course, instead of raindrops. The resulting spectrum reveals the specific colors—or wavelengths—present in any given beam of light.

And with the spectroscope, scientists have been able to establish a kind of atomic fingerprinting system which labels matter according to the composition of the light it emits.

It was the spectroscope that made possible the discovery of one of the unknown elements predicted by Mendeleev.



We can even identify elements in a crude way by the obvious color of the light they emit.

Sodium sprinkled into a flame gives off a bright yellow light.

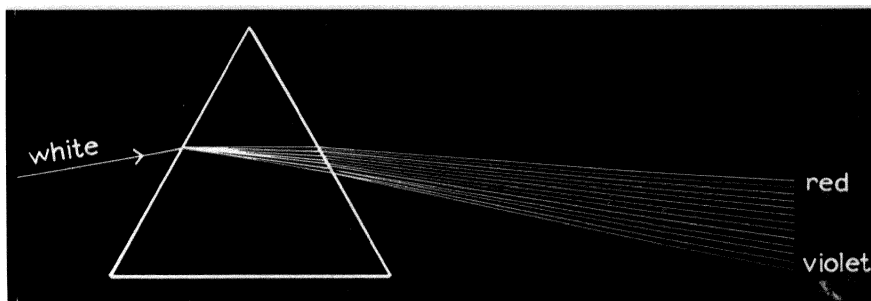
A copper compound thrown into a flame emits a bright green. The light from a strontium compound is crimson.

The light emitted by each of these elements is characteristic of the element and therefore can be used in its identification.

To see this light in far greater detail, we can pass it through a prism, or through a diffraction grating, which acts like a long line of equally spaced, microscopic prisms.

What a prism or grating does is to bend light according to its color. When white light—which contains in it all the colors—is beamed through a prism, the red is bent least.

Orange light is bent more than red; yellow more than orange; and so on through green, blue, and violet.

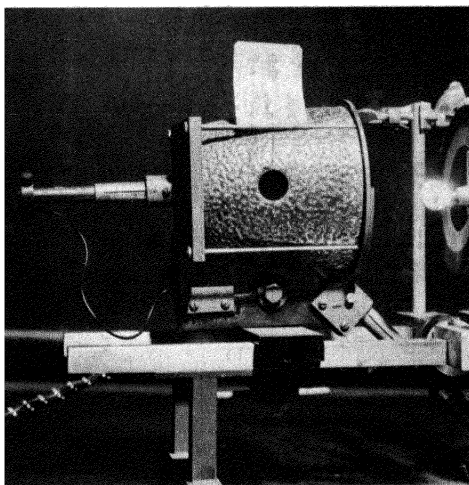


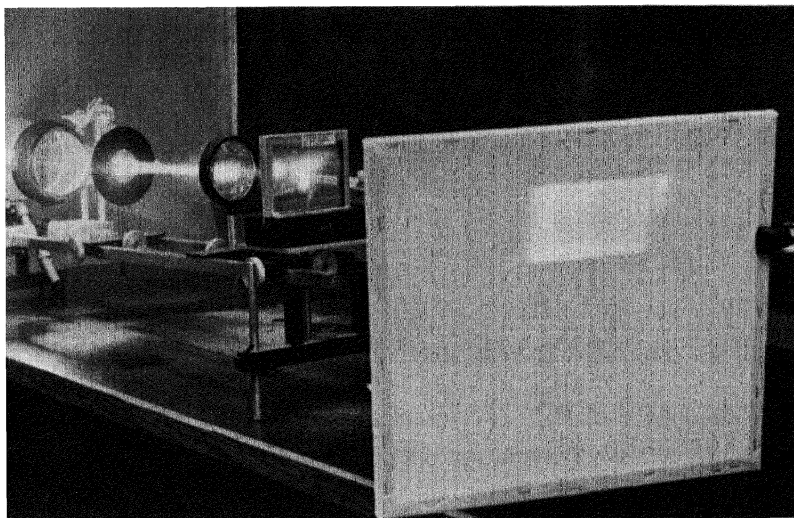
If the source of light is a carbon arc, the light approximates white light because the characteristic light of carbon goes clear across the spectrum—as the range of colors is called—from red through orange, yellow, green, blue, and indigo to violet.

A carbon arc consists basically of two sticks of carbon perhaps half an inch in diameter with a gap between them. In the photograph, one carbon is inserted horizontally on the left and the other slants upward from below. The ends almost meet inside a metal housing. The gap between the two carbons is directly behind the black circle, which is a dark red viewing window.

Light from the arc is projected out to the right.

The beam of light from the arc passes





through a lens, a vertical slit, another lens, and then through the rectangular diffraction grating. The spectrum is projected onto the screen at the right.

When a strong electric current is sent through the carbons, a steady spark—the “arc”—jumps across the gap and the heat and electric current excite the carbon atoms. The atoms thus emit the light typical of carbon.

If one of the carbons is first dipped into a solution of some other element, the light of this other element is superimposed upon the light from the carbon.

For example, if one of the carbons is first dipped into a sodium solution, thus picking up some sodium, the resulting spectrum when the arc is struck will be that of carbon plus sodium. The sodium light stands out because the sodium spectrum is very strong in the yellow region while the light from carbon is relatively continuous, spreading out clear across the spectrum.

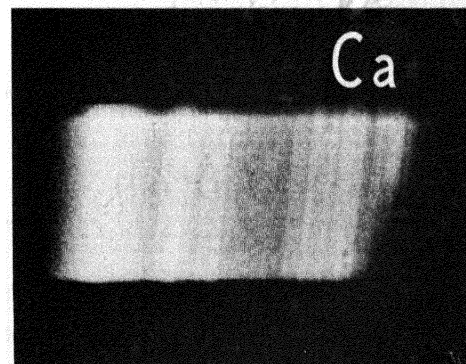
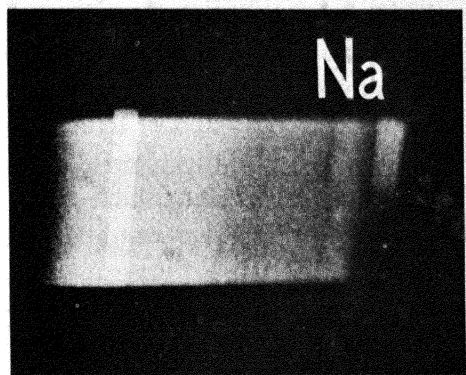
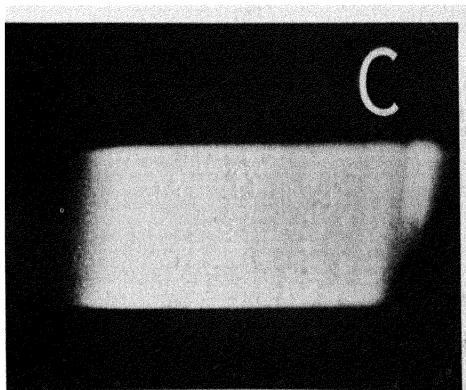
Even a crude spectrum will show these differences clearly. The carbon spectrum itself is a rather steady band, somewhat

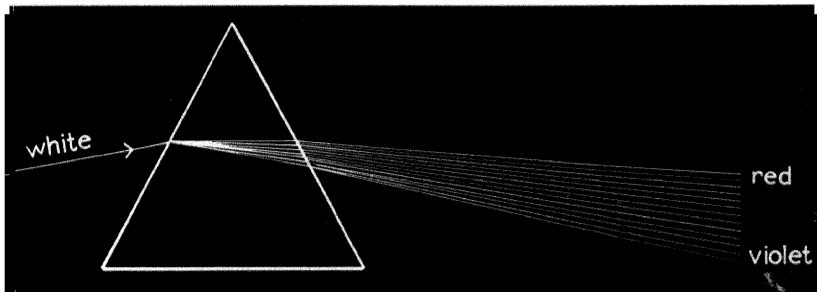
similar to white light from the sun.

When we use a carbon soaked in a solution of sodium, the spectrum which shows through the diffraction grating has a strong yellow component.

This is the color responsible for the yellow of a sodium vapor lamp. Objects colored white or yellow reflect the light of sodium very well. A red dress, however, will look dull brown because there is no red in the sodium light to be reflected by the red material.

If we use a carbon soaked in a solution of calcium, the resulting spectrum shows the colors characteristic of calcium.



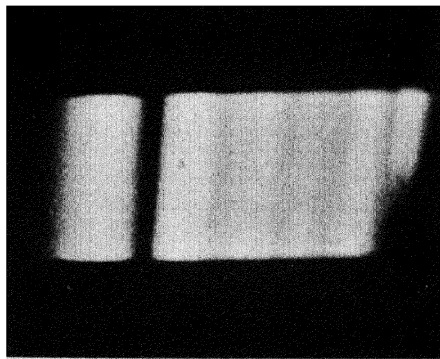


Sodium light passed through a grating or prism results in a spectrum almost blank except for the narrow yellow lines.

We've been talking so far about the "emission spectrum" of an element, the kind of light a glowing element emits.

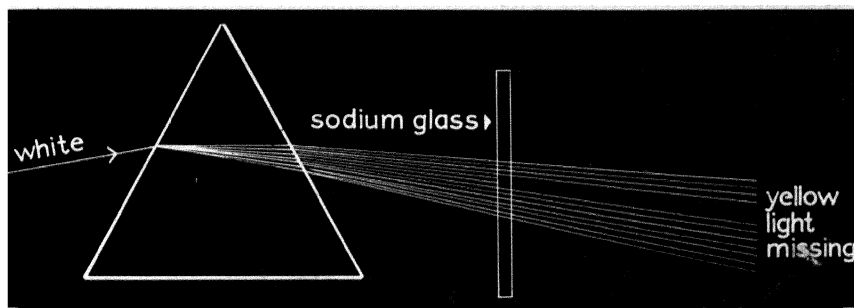
An element also can absorb precisely the same light as it is capable of emitting.

Sodium, for instance, can absorb the same wavelengths of yellow light that a sodium vapor lamp gives off. Glass with sodium in it, when held in front of a carbon arc, will absorb and thus screen out these same wavelengths. If we look at the resulting light through a prism or diffraction grating, the



sodium yellow light will be missing; the characteristic pattern of the sodium emission spectrum will now appear in reverse like a negative; that is, as dark "lines" instead of bright "lines."

So if we hold a



piece of glass containing sodium between a pure carbon arc and our grating, we come out with a carbon spectrum *minus* the sodium spectrum. This is the "absorption spectrum" or "dark line spectrum" of the element sodium.

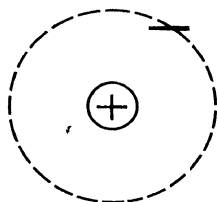
Thus there are two ways of identifying an element with a spectroscope: by the peculiar pattern of colors, or wavelengths, which the element emits when excited; or by the negative of this pattern, the colors which the element has absorbed and thus fails to transmit.

Obviously, we have in the spectroscope a powerful technique for identifying unknown substances as well as known elements and compounds.

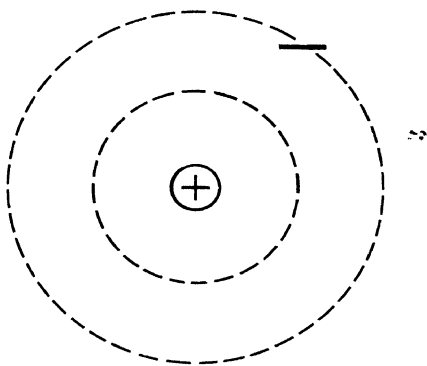
Furthermore, the method is so sensitive that it can be used to identify as little as one ten-billionth of an ounce of such an element as sodium. The spectroscope is also independent of distance, and sunlight or starlight can therefore reveal the presence of specific elements in the sun or in the stars.

The mechanism by which atoms absorb or emit light has to do with the positions of the outer atomic electrons in the atoms.

The simplest case is that of the hydrogen atom with its nucleus—a proton—and a single electron in orbit around the nucleus. When the atom absorbs light, the electron jumps out into another orbit with a greater radius. When the atom emits light and releases energy, the electron jumps inward to a smaller orbit.



The diagram would be similar in the case of the outer electron of a lithium atom or of a sodium atom. When the electron jumps into a tighter orbit, the atom emits light characteristic of it.



We can now return to Mendeleev's famous prediction that three new elements should be discovered to fill the three blanks in his Periodic Table.

A few years after this prediction, in 1875, the French chemist Lecoq de Boisbaudran was working with zinc ores. He was well aware of Mendeleev's "prophecy" and knew what kind of a new element to watch for. Using a spectroscope, de Boisbaudran identified the predicted element eka-aluminum in a sample of zinc blende. He named it gallium after Gallia, the old Roman name for his native country.

Gallium, incidentally, follows zinc in the modern Periodic Table. The fact that it is found in zinc ores is an example of similar chemical properties sometimes found between two neighboring elements in the same horizontal row; such similarities are usually and more normally found among elements in the same vertical column of the Periodic Table.

Gallium is a solid, but it has a melting point just above room temperature; gallium will melt if you hold its container in your hand for a few minutes.

Eka-boron was discovered in 1879 by Lars F. Nilson of Sweden. He named it scandium, for Scandinavia.

A German chemist, Clemens Alexander Winkler, discovered Mendeleev's predicted element eka-silicon in 1886 and named it germanium, for Germany.

Following is a sample of some of the chemical properties predicted by Mendeleev for eka-silicon, compared with the properties actually found in the newly discovered germanium

	eka-silicon	germanium
atomic weight	72	72.6
density	5.5	5.47
atomic volume	13	13.2
color	dirty-gray	grayish-white
properties of its chloride	liquid, boils below 100°	liquid, boils at 86.5°

In all three elements there was an amazing correspondence between predicted and actual properties. And this proved the genius of Mendeleev and his Periodic Table of the Elements.



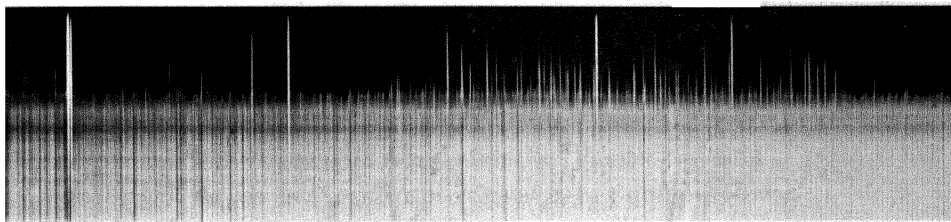
Mendeleev had the great satisfaction of seeing these three elements discovered during his own lifetime.

Half a century after his death, the element with the atomic number 101 was discovered at the University of California's Radiation Laboratory and named mendeleevium in honor of this great Russian chemist.

Even before the discovery of gallium, scandium, and germanium, the spectroscope had been used in a spectacular way to identify another unknown element.

During an eclipse of the sun in 1868, a spectroscope was used to spread out and examine for the first time the light from the chromosphere—the sun's glowing atmosphere.

Here is part of the spectrum which the French astronomer Pierre Jules César Janssen observed.



Janssen discovered a brilliant yellow line in this solar spectrum. Two other yellow lines were easily recognized as those of sodium. This new line obviously came from some element, but it was a line that had never been seen before.

It was soon agreed that this particular yellow light came from an unknown element in the sun. The element was named helium, after the Greek word for sun, helios.

It was twenty-seven years before helium was known to exist on earth as well as on the sun.

Meanwhile, there was no way to fit helium into Mendeleev's Periodic Table as the table existed in the 1870's and 1880's.

The discovery of earthly helium and an addition to the Periodic Table to include it both resulted from a new line of experiments which had nothing to do with the discovery of helium in the sun.

In the 1880's, Lord Rayleigh was professor of physics at the Cavendish Laboratory at the University of Cambridge in England. He spent quite a number of years studying the density of gases, and he was particularly interested in nitrogen.

He found that nitrogen made from ammonia had a density which was about one half of one per cent less than that of nitrogen he was able to extract from air.

Why?

This was actually a pretty small difference, and a lot of people wouldn't have bothered pursuing the matter any further. But the chemist William Ramsay thought the difference was well worth investigating, and he got permission from Rayleigh to continue the work.

He took a quantity of air and removed the oxygen from it. Then he removed the nitrogen. And he still had a small residue left.

This residue of gas he put into a glass tube and excited it electrically. He studied the glowing gas with a spectroscope and discovered that he had some new lines which didn't belong to any known element. The peculiar thing was that there seemed to be no place in the Periodic Table for such a new gaseous element to fit in.

In 1894, however, Ramsay got the idea that there might be an entire vertical column of elements missing from the Periodic Table. No element in this column had been known to Mendeleev, and the great Russian chemist therefore had no clue which might suggest an "extra" column of elements.

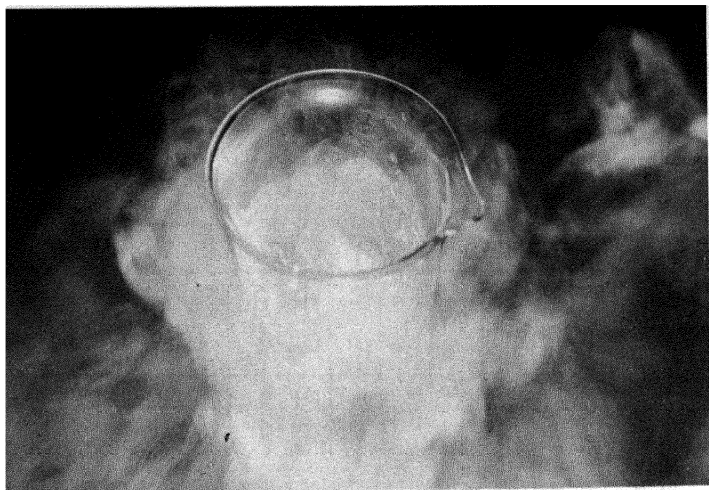
Ramsay thought that perhaps his newly discovered gas was

		[H ?				[He ?												[A ?				
Li	Be											B	C	N	O	F	?					
Na	Mg											Al	Si	P	S	Cl	?					
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	?					
Rb	Sr	Y	Zr	Nb	Mo		Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	?					
Cs	Ba			Ta	W		Os	Ir	Pt	Au	Hg	Tl	Pb	Bi								
		↓		Th		U																
		La	Ce	Pr	Nd		Sm		Gd	Tb	Dy	Ho	Er	Tm	Yb							

the first member of a column of new elements, illustrated here within the frame of a modern Periodic Table but including only the elements known at that time.

He named the gas "argon," which means "lazy" in Greek. The reason for the name was that argon seemed to have no chemical properties at all. It had no color, odor, or taste, and it would not react chemically with other elements.

The following year Ramsay discovered another new gas which appeared when the rare mineral cleveite was heated. He studied the spectrum of this gas and found that it was identical with the spectrum of helium in the sun. He continued his investigations of air with the help of a young student, M. W. Travers, and they soon managed to separate helium from air. Then they began looking for more gases like argon.



What they did was to distill liquid air.

The various components of liquid air can be separated by boiling because each component has a different boiling point. For example, the nitrogen in liquid air boils away before the oxygen. This is called fractional distillation, and it is the basic method of separating the components of crude oil.

Ramsay and Travers used a discharge tube to excite the gas they obtained, and they analyzed the resulting light with a spectroscope to help find if they had, indeed, isolated any new elements. In a short space of time they did find three new elements. These they named neon (meaning “new”), krypton (“hidden”), and xenon (“stranger”).

All three were just as inert as helium and the “lazy” argon. They are now called the “rare gases” or “inert gases.”

The last of these gases was discovered in 1900 by Ernst Dorn. It had been known that the radioactive decay of radium produced a gas, and it was Dorn who recognized this gas as being the last and heaviest member of this whole new family of gases. It was known to be radioactive and was named radon.

The inert gases have a number of uses, and many of these depend upon the very fact that they are inert. The first commercial production of helium caught the British by surprise when a German Zeppelin was hit by incendiary bullets but did not catch fire. The Zeppelin, of course, had been filled with helium instead of highly flammable hydrogen.

Although it is the next lightest element after hydrogen, helium's properties are very different.

Divers now breathe a mixture of oxygen and helium instead of the oxygen-nitrogen mixture of ordinary air. Nitrogen dissolves in a diver's blood and causes the "bends"; helium will not dissolve to such an extent.

As early as 1903, Ramsay demonstrated that helium was a product of the radioactive disintegration of radium. His experiment was an early example of the transmutation of one element into another. Today, helium is used as a source of bombarding particles in modern atom smashers. The helium nucleus—stripped of the atom's two electrons—is a common product of radioactivity and is known also as an "alpha particle."

Argon is used in welding because it is inert and thus prevents metals from being oxidized—burned—when welded. Argon is used in Geiger counters and in incandescent lamps.

Argon, neon, and some of the other inert gases are responsible for our "neon" signs. Various colors are obtained by coating the inside of a glass tube with fluorescent material and filling the tube with one or another mixture of inert gases.

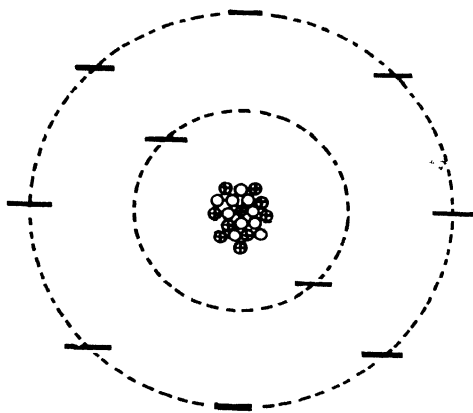
He
Ne
A
Kr
Xe
Rn

The most unusual thing about these gases is that they have essentially no chemical properties at all. They do not combine or react with anything else.

This odd property of not having any appreciable chemical properties is due to the electron structure of the atom.

In an atom of sodium there are two inner electrons, a complete group of eight electrons, and a single electron still farther from the nucleus which circles all by itself. Sodium is a very reactive and potentially dangerous element, and all of its reactive properties are due to the presence of this one outer electron.

The element before sodium does not have this active outer electron, and so we have this symmetrical picture of the



electron situation in an atom of neon. The ring of eight electrons, as well as the inner group of two, is complete in itself. You might say that neon is simply not on the make. It has no tendency to borrow another electron from some other atom, nor has it an extra electron to spare or share. Thus neon has almost no properties

at all.

If we have an atom with one less electron than neon has, we find a gap in the outer ring of electrons. This is the case with fluorine, which acts as if it dearly wished to find an extra electron to fill the gap and complete its outer ring. So fluorine, like sodium, is one of the most reactive of all the elements.

The inert, or rare, gases are indeed rare on earth, which

likely accounts for the fact that they were discovered rather late. They were all fitted into the Periodic Table in a vertical column that Mendeleev could not have predicted.

Meanwhile, gallium (Ga), scandium (Sc), germanium (Ge), and fluorine (F) had been discovered, as well as eight new elements in the “rare earth” group—ytterbium (Yb) in 1878, samarium (Sm), holmium (Ho), and thulium (Tm) in 1879, praseodymium (Pr) and neodymium (Nd) in 1885, and gadolinium (Gd) and dysprosium (Dy) in 1886.

Polonium (Po) and radium (Ra) were discovered in 1898, and then followed actinium (Ac) in 1899, europium (Eu) in 1901, lutetium (Lu) in 1907, and protactinium (Pa) in 1917.

Lanthanum and the fourteen “rare earth” elements called “lanthanides” (because they are like lanthanum) belong in the Periodic Table between barium and hafnium. They are listed in a separate row at the bottom for convenience. The “rare earth” elements were so called because they resembled

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo		Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po		Rn
	Ra	↓															
		La	Ce	Pr	Nd		Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
		Ac	Th	Pa	U												

substances like lime and magnesia which were once known as "earths."

Actinium and the "actinide" elements (thorium, etc.) also are listed separately at the bottom. Each actinide element is somewhat similar to actinium and to the lanthanide element with which it is paired. The discovery of hafnium (Hf) in 1923 and rhenium (Re) in 1925 completed the Periodic Table through uranium except for four blank spaces.

Every element is listed in the Periodic Table according to its atomic number. This number represents the magnitude of the positive electrical charge on the atom's nucleus or—which is the same thing—the number of negative charges (electrons) revolving around the nucleus.

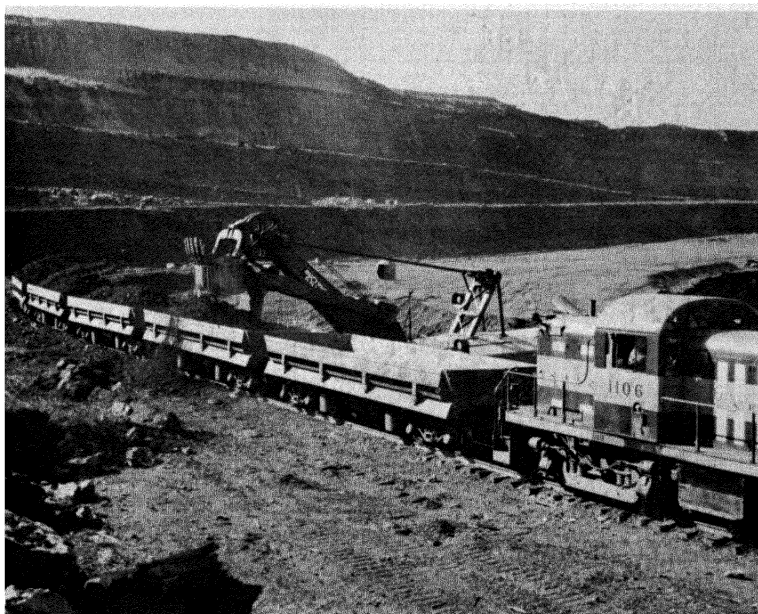
In the first horizontal row, or period, of the Periodic Table are two elements, hydrogen and helium. There are eight elements in the second period and eight more in the third, then eighteen in the fourth period, and eighteen more in the fifth.

In the sixth period are eighteen elements plus the fourteen lanthanide elements, making thirty-two in all. Our modern Periodic Table shows a seventh period which theoretically has spaces for eighteen elements plus fourteen actinide elements, although this is not the way the table looked back in 1925.

If we consider the vertical columns instead of the horizontal rows, we find elements with similar chemical properties listed one below the other. Group I, for example, includes hydrogen and the "alkali metals" like lithium and sodium. Group III includes scandium, yttrium and all the lanthanide and actinide elements.

The group on the far right is that of the "noble gases"; next to it are the "halogens" ("salt-generating" elements) like fluorine and chlorine, each of which has a "gap" in its outer ring of electrons.

PUTTING ELEMENTS TO WORK



The eighty-eight elements known to man by 1925 are the natural elements, those found in the earth and in the water and in the air.

In their natural, unrefined form they usually look like nothing more or less than rock, or the pulverized rock that we call soil and sand.

And it is from the earth, from carloads of rock, that man has learned to draw out the raw materials of ships and sky-scrapers, radios and artificial satellites.

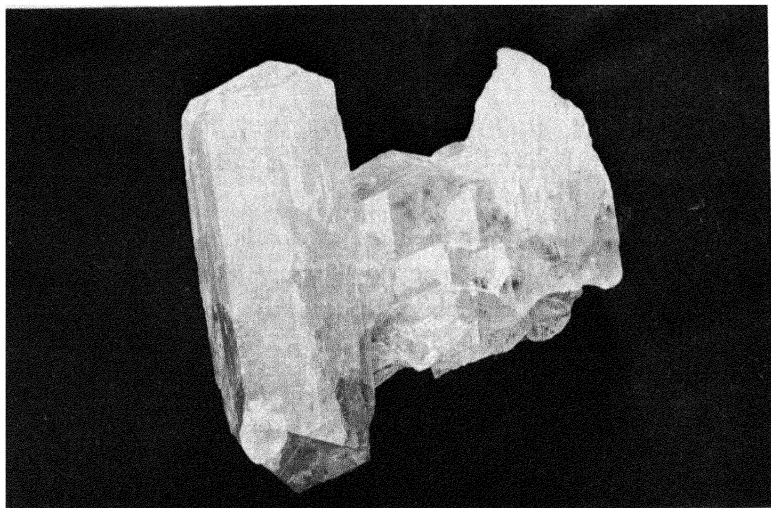
The problem of getting at the elements we want and putting them to work is the problem of industry, ancient and modern. We can consider three broad categories of elements, according to how we attack this problem in practice: those elements existing free in nature; those we have to extract from

their ores; and those that exist in natural combinations which can be used as they are or rearranged without being converted back into an elementary form.

Among the elements found in nature alone and uncombined are carbon, sulfur, oxygen, nitrogen, and the entire group of inert gases—helium, neon, argon, krypton, xenon, and radon.

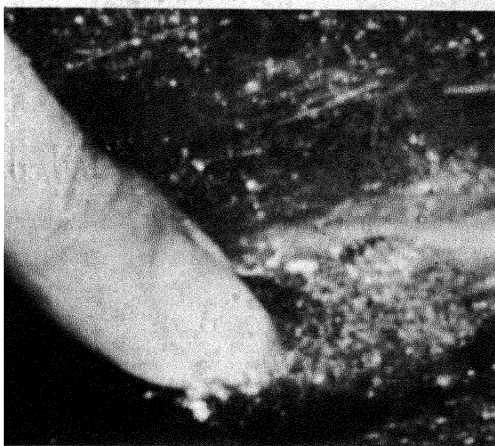
These inert gases, of course, all fall in the same vertical column in the Periodic Table, which indicates that they have similar chemical properties. In the present case, we find that this similarity includes the property of existing free in nature. Another example of similar elements in the same column is copper, silver, and gold, all of which exist free in nature on occasion. Sulfur (below) may occur naturally in the form of large crystals.

Many of these “free” elements occur in mixtures and can be separated with relatively little trouble. A mixture of iron and sulfur, as we know, can be separated with a simple magnet.



One of the most vivid chapters in America's history is due to the simple fact that gold may occur in a mixture with sand and can be separated with a very simple tool, namely a pie tin.

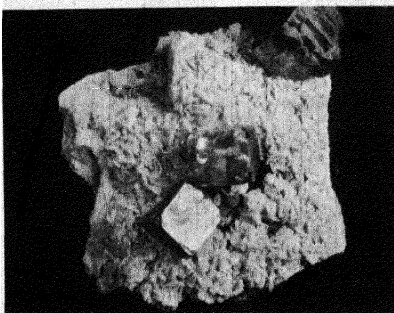
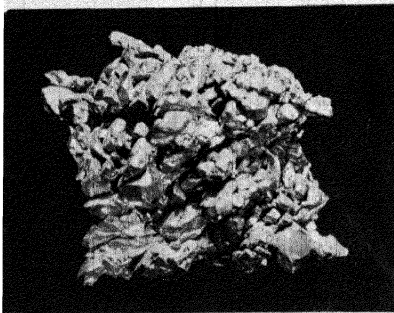
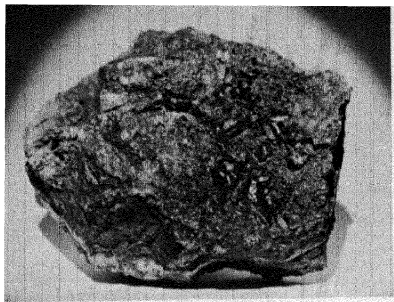
Because gold happens to be heavier than sand, it will sink to the bottom when a kind of stream-bed soup — from the right stream, of course—is agitated in a flat pan. When sand and water are



skillfully poured off, relatively pure gold flakes remain.

The most familiar mixture of chemically free elements is air, which is about four-fifths nitrogen and one-fifth oxygen, with a measurable trace of inert gases. We've already seen how these elements can be separated by distilling liquid air because each has a different boiling point. The nitrogen in liquid air will always boil off first.

Most elements occur naturally in the earth in chemical compounds—ores.



Ores often contain oxygen (as bauxite does) or sulfur (as do cinnabar, argentite, and galena).

A cinnabar molecule (HgS) contains 1 atom of mercury and 1 of sulfur. Argentite (Ag_2S) consists of 2 atoms of silver for every sulfur atom. Galena (PbS) is lead sulfide.

An ore from which the metal is easily extracted is mercuric oxide (HgO). We only have to heat it in a simple still, as in the diagrams on page 27. Heat separates the mercury atom from the oxygen atom in each molecule of mercuric oxide; single mercury atoms form the resulting mercury

while the oxygen atoms combine in pairs and escape as molecules of gaseous oxygen.

It is a little more difficult to reduce lead from the ore, lead oxide (PbO), because the oxygen atom seems reluctant to tear itself away from the lead atom. The difficulty is overcome by heating lead oxide on a graphite plate as on page 54. The graphite, which is a form of pure carbon, introduces carbon into the resulting chemical reaction.

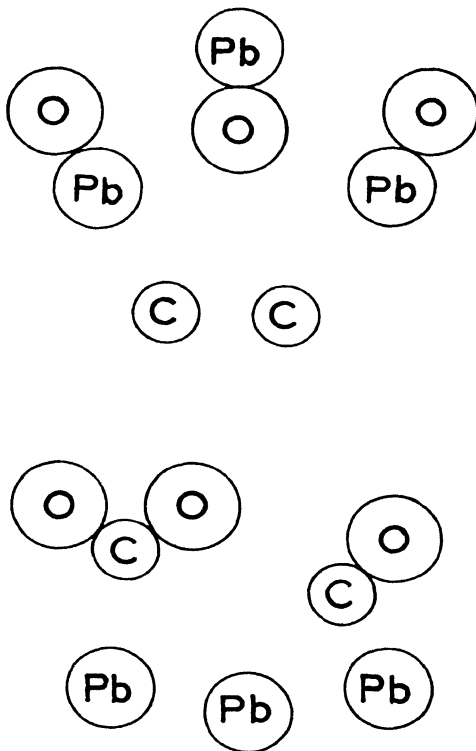
Carbon has a greater attraction for oxygen than lead has, and the carbon thus succeeds in "stealing" the oxygen atoms away from the lead.

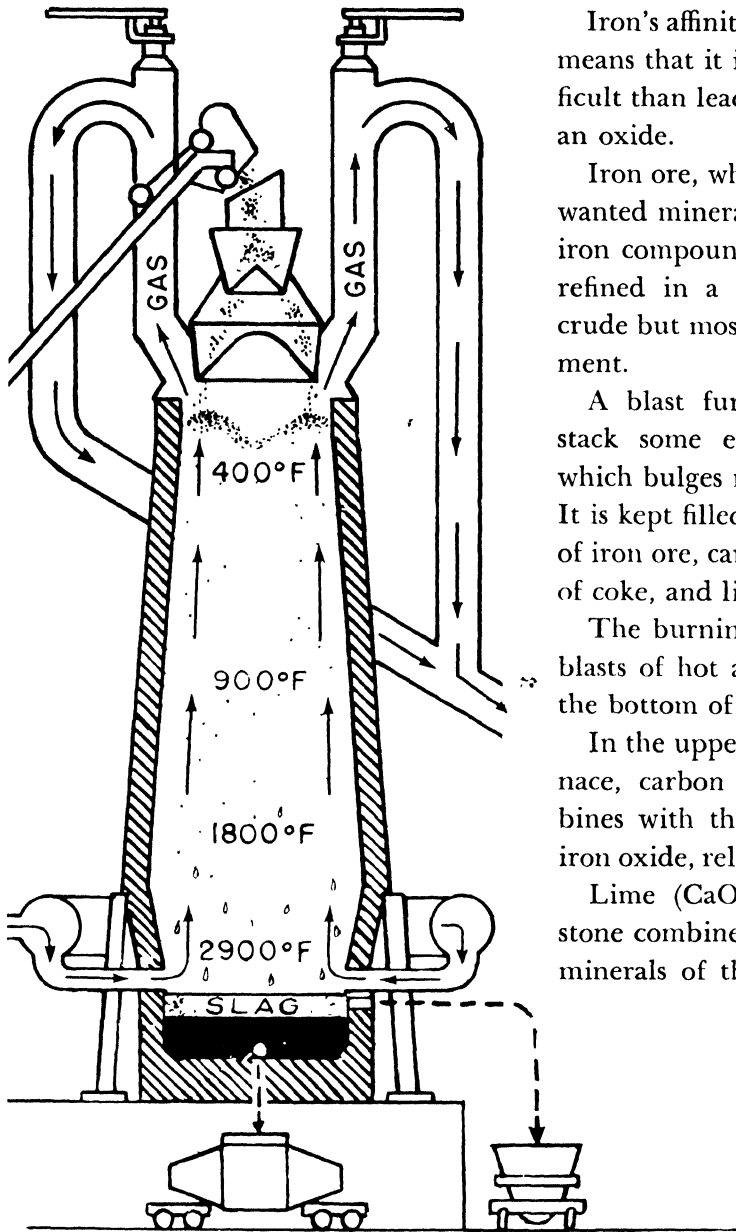
When lead oxide is heated on a graphite plate, each carbon atom combines with one oxygen atom, forming carbon monoxide, or with two oxygens, forming carbon dioxide. Both of these gases escape, leaving the pure lead behind.

The process is more complex in the case of iron, the cheapest and most abundant of the heavy metals.

The value of iron to mankind stems from a number of remarkable properties. It can be made hard by sudden cooling, for example, or made pliable by slow cooling. Also, the brittleness of hard steel can be removed by tempering, which involves very careful reheating.

Elementary iron is difficult to isolate because it combines most readily with oxygen. Iron implements therefore rust easily, forming an iron oxide. And it is not surprising to find that the most plentiful iron ores are oxides such as hematite (Fe_2O_3), with two atoms of iron and three of oxygen in each molecule, or magnetite (Fe_3O_4).





Iron's affinity for oxygen also means that it is even more difficult than lead to reduce from an oxide.

Iron ore, which contains unwanted minerals as well as the iron compounds, is commonly refined in a blast furnace, a crude but most efficient instrument.

A blast furnace is a great stack some eighty feet high which bulges near the bottom. It is kept filled with a mixture of iron ore, carbon in the form of coke, and limestone.

The burning coke is fed by blasts of hot air coming in at the bottom of the stack.

In the upper part of the furnace, carbon monoxide combines with the oxygen of the iron oxide, releasing pure iron.

Lime (CaO) from the limestone combines with the waste minerals of the ore and with

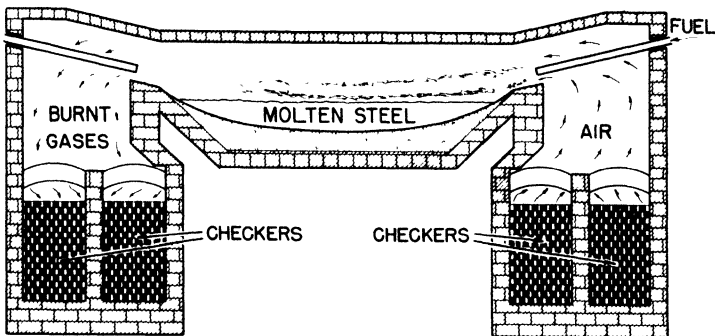
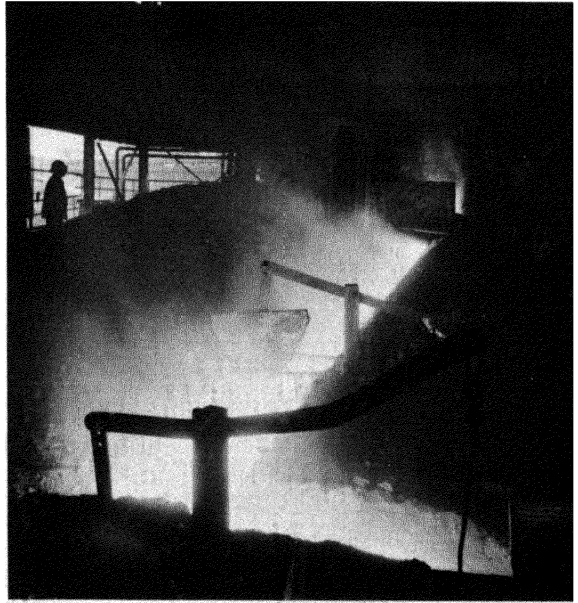
ash from the burned coke, producing a molten slag which can be drawn off at the bottom.

The iron meanwhile has melted, picked up a quantity of carbon, and collected at the very bottom, directly beneath the lighter slag.

The iron is drawn off in the form of molten pig or cast iron which contains as much as four per cent carbon.

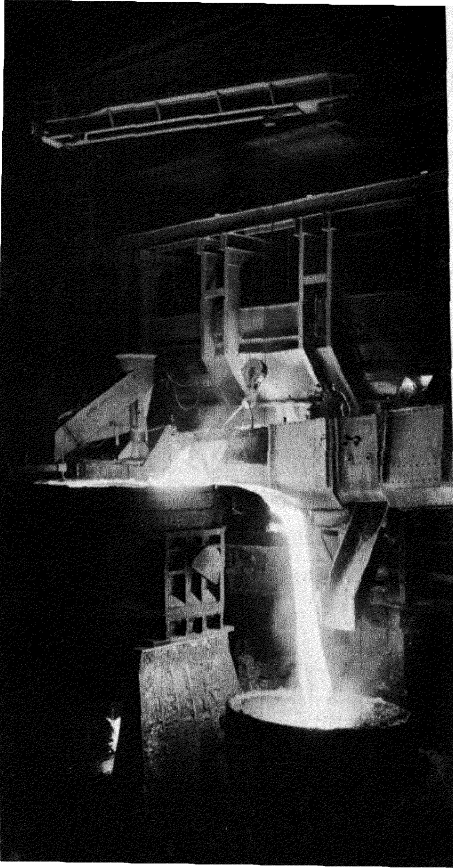
The chemical reaction is similar to that of carbon with lead oxide. Carbon monoxide takes oxygen atoms from the iron oxide, forming carbon dioxide and freeing the iron.

The pig iron from the blast furnace may be further purified by various processes. One example is the open hearth furnace which reduces the amount of carbon from four per cent



to about one-half of one per cent. This is accomplished essentially by a flame played across the molten pig iron in an enclosed, shallow basin.

The product of the open-hearth furnace is now referred to as steel, or open-hearth steel. Here the molten steel flows from a 275-ton capacity furnace.

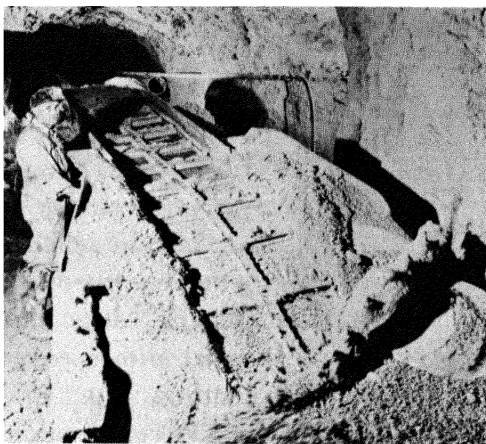


Different quantities of carbon in the steel account for wide differences in the metal's properties.

Mild steel, for instance, has a low carbon content and is quite malleable while cast iron contains so much carbon that it may shatter readily under a hammer blow.

Steel can be further improved by the addition of other elements to make alloys. Nickel and chromium are used to make stainless steel.

Molybdenum, vanadium, tungsten, cobalt, titanium and other elements may be added to improve hardness, strength, or magnetic properties.



Some ores are even more reluctant than iron to give up the metal we're after. An example is the aluminum oxide, bauxite (Al_2O_3).

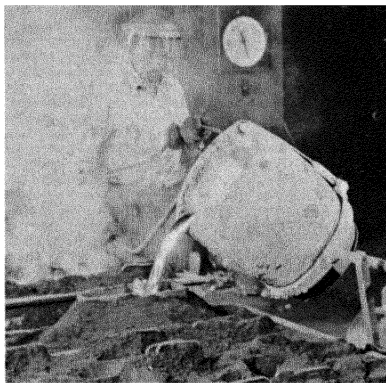
In this case, we can't get the aluminum out by heating. We have to go farther.

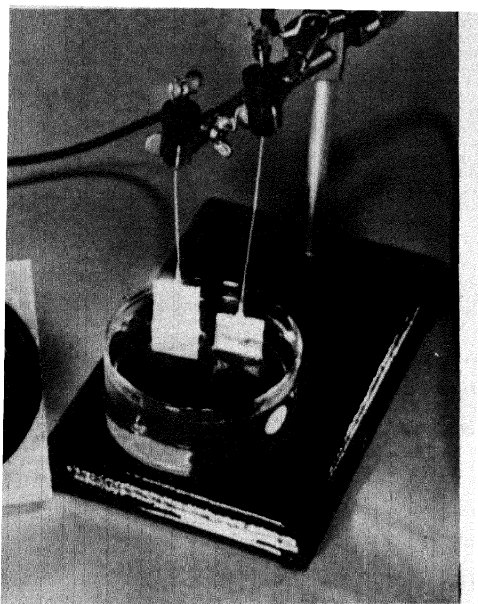
We have to electrolyze the ore; this means passing an electric current through melted ore and adding additional material to the molten bath to improve the electrolysis.

The current passes between a carbon "anode" and a cast-iron plate. Flowing through the molten ore, the current causes oxygen from the oxide to combine with carbon in the anode while aluminum metal collects at the iron plate.

The resulting aluminum may be used directly or alloyed with other elements like magnesium or copper or manganese.

The electrolysis of bauxite is a complicated process, but the basic principle can be illustrated with copper, which can be deposited more easily than aluminum.





We need a solution of copper sulfate and two electrodes connected to an electrical power source.

One electrode, the anode or positive pole, is a strip of copper or some other metal with a copper coating.

The electric current causes copper from the anode on the right to be drawn into the solution while copper from the solution is deposited on the "cathode," which in this case is the platinum strip on the left. Thus the crude copper is transferred to the platinum cathode just as aluminum is deposited on a cast-iron cathode.

The same process is used in electroplating, where a film of one metal is deposited on another.

Elements of every kind are distributed unevenly across the surface of the earth. The particular location of certain ores has had a notable influence upon the history of mankind, and it has naturally determined the location of industrial centers throughout the world.

The refining of ores like aluminum can be accomplished with electric power near the ore source.

In the case of iron, stone furnaces were once built at the site of an ore deposit and fed with charcoal made from nearby trees.

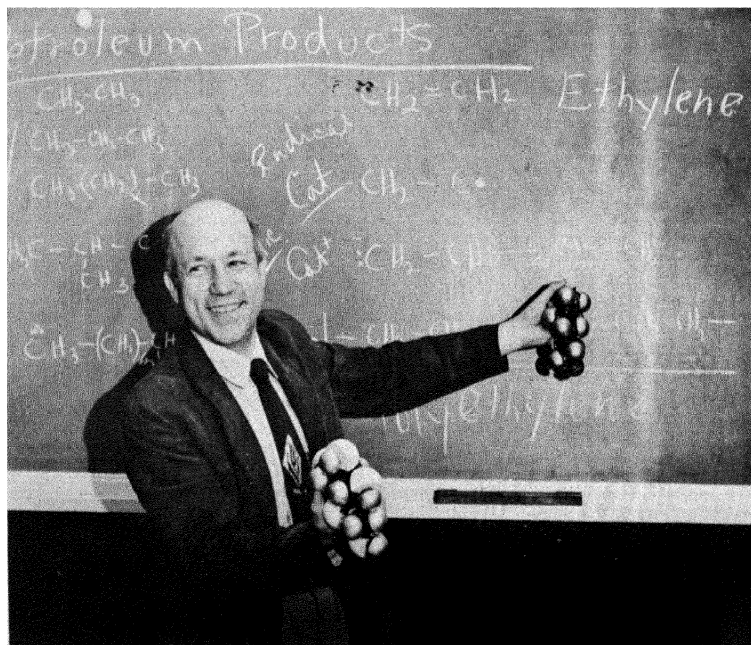
But the mass production of iron and steel today requires vast quantities of coke, and the ore is usually brought in from elsewhere. In North America, for example, iron ore from Michigan is moved eastward to the coal regions. New routes are constantly being opened, and today ore from Labrador is beginning to move through the St. Lawrence River to the furnaces.

Pittsburgh, like Essen and Newcastle, became a great industrial center just because of the rich coal veins in the area. So the chance distribution of elements in the earth millions and billions of years ago is reflected in our maps and history books today.

A third broad category of elements involves those which exist naturally in combination and are merely rearranged for commercial use; they are never handled in their elemental form.

A good example of this method of putting elements to work is the field of organic chemistry—the chemistry of carbon compounds—and particularly in the extraction of various materials from petroleum.

The chemistry of petroleum will be described here by Melvin Calvin, Professor of Chemistry and member of the Radiation Laboratory at the University of California in Berkeley and one of the world's outstanding organic chemists. Dr. Calvin is Director of the Bio-organic Chemistry Group at the Radiation Laboratory and is noted for his productive research in the chemical processes of photosynthesis.



ORGANIC COMPOUNDS

Our problem here is to try to describe how we make use of elements that occur naturally in compounds without transforming them back to their elemental state.

The principal one we're concerned with is the element carbon, which occurs in nature primarily in the form of coal or petroleum.

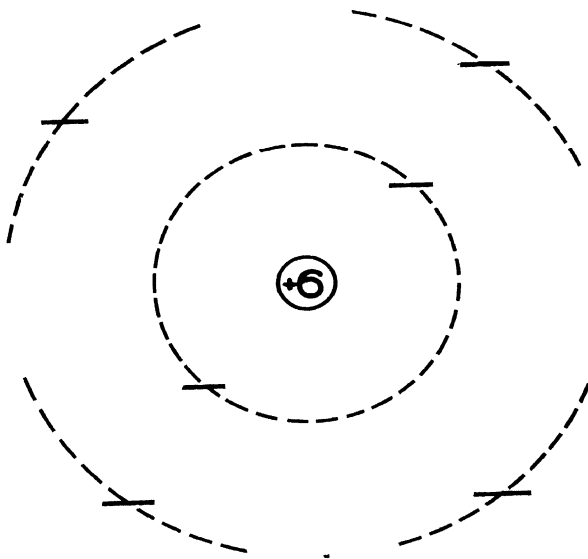
Coal, of course, consists largely of carbon.

Petroleum is a mixture of compounds made mainly of carbon and hydrogen. Crude petroleum, as it comes from the earth, is a black and rather smelly mixture.

Although the crude petroleum may be used directly—in the form of diesel fuel, for example—it is more frequently transformed in various ways.

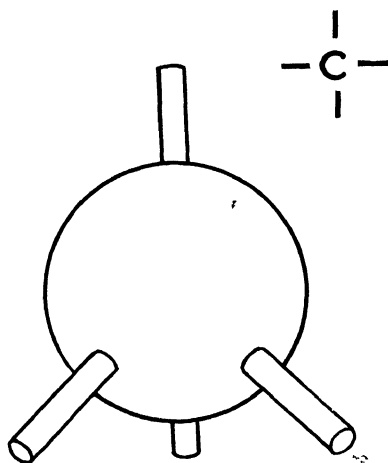
In order to get some idea of how this can be done, it is necessary to examine what the nature of the molecules which go to make up our crude mixture might be.

Carbon being the principal element of the mixture, it is worthwhile to have a look at a model of a carbon atom. It has a nucleus with six positive charges; it has 2 inner electrons and 4 more electrons in its outer shell.



These 4 outer electrons are in a shell or ring which could contain up to 8 electrons, as in the case of neon. They are the electrons which make it possible for a carbon atom to combine with as many as four other atoms.

As organic chemists, we generally use a simple model of the carbon atom—a ball and four little rods to represent the



four electrons by which the atom manages to combine with atoms of other elements and with other carbon atoms. These other atoms contribute electrons in a similar manner so that “bonds,” each consisting of two electrons, are formed.

The carbon atom also may be represented two-dimensionally by a “C” and four short lines.

In its chemical activity, carbon is somewhat like a metal because of its four “extra” electrons, and yet like a halogen because of the four “gaps” in its outer ring of electrons. And on the Periodic Table carbon lies midway between the metal, lithium, and the halogen, fluorine.

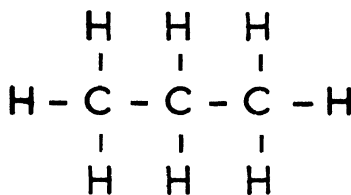
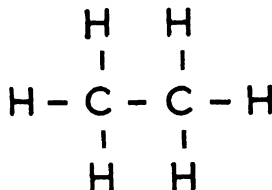
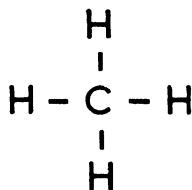
This unique property of carbon means that a carbon atom will combine readily with other atoms in a fantastic variety of ways. Carbon is a constituent of some 500,000 known chemical compounds and a major ingredient of all living matter.

Many thousands of carbon compounds contain only the elements carbon and hydrogen. These compounds, the hydrocarbons, are associated with plant and animal life, which is why the study of them is a part of “organic” chemistry. Since petroleum was formed from ancient living matter, it is no

surprise to find petroleum particularly rich in hydrocarbons.

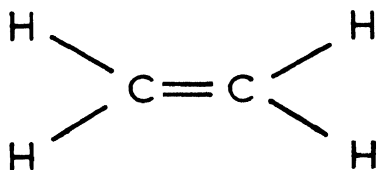
If the four rods of our model carbon atom are combined with hydrogen atoms, we get the molecule, methane (CH_4) or marsh gas, which is the main constituent of natural gas.

If, instead of combining the carbon with four hydrogens, we combine it with three hydrogen atoms and a second carbon atom with three hydrogens attached to it, we will get the next higher molecule in the series, ethane (C_2H_6). This is followed by propane (C_3H_8), and you can see how this process



can be repeated and very long carbon chains constructed.

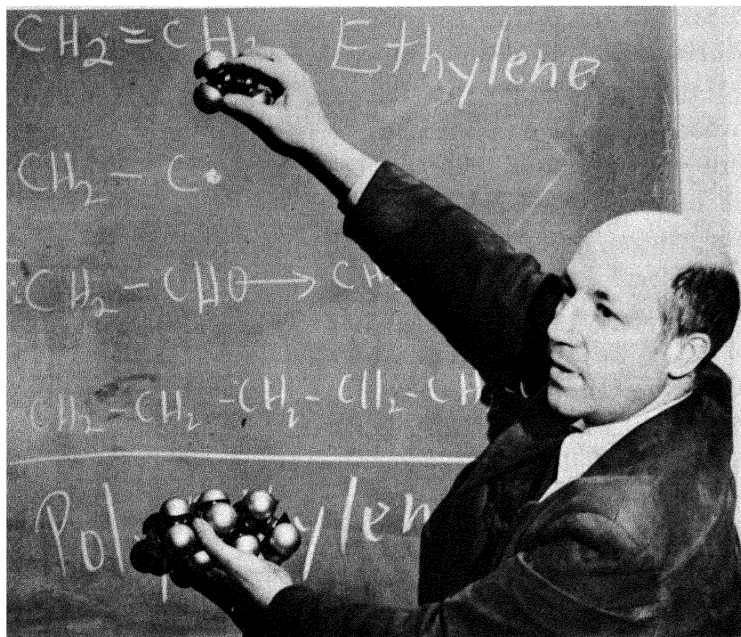
Ethylene (C_2H_4) is characteristic of another series of hydrocarbons, and in this series there is always a double bond between one pair of carbon atoms. A double bond means that each atom contributes 2 electrons to the link so that



4 electrons are involved in the bond between a pair of atoms. This link is stronger than a single bond.

In other groups of hydrocarbons, carbon atoms may be attached by triple bonds, as in acetylene (C_2H_2), or linked in ring formations.



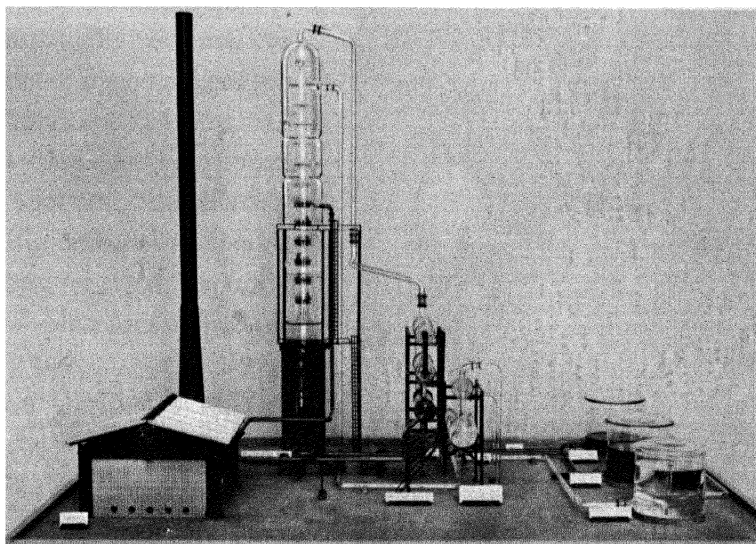


Crude petroleum consists of a mixture of many molecules of this sort, molecules which usually contain from six to several hundred carbon atoms.

In order to make use of these things, it is necessary to separate the various molecules from one another. This is most readily done by a process of distillation.

The crude petroleum is heated at the bottom of a tower-like still. As the vaporized petroleum rises through the still, the heavier molecules such as lubricating oils condense near the bottom. The lighter molecules, such as kerosene, condense higher in the tower, and gasoline, lightest of all, condenses at the top and leaves the column first.

In this way, we are able to separate the petroleum into a whole series of fractions depending upon how easy it is for the various molecules to boil away from the hot petroleum.

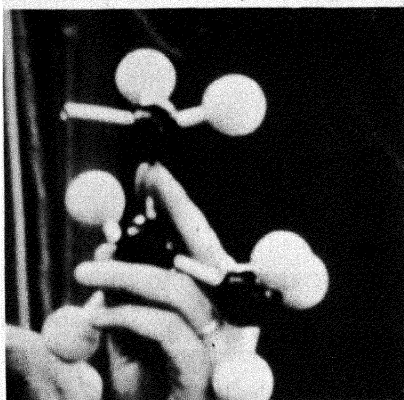
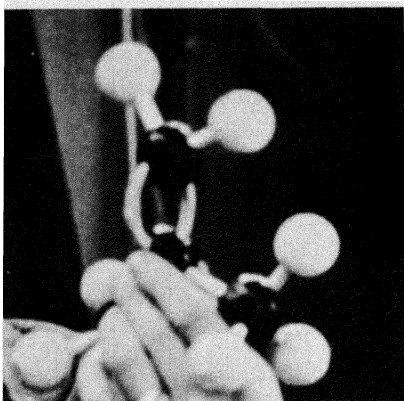


The lighter fractions like kerosene and gasoline are already useful as fuel. The heavier ones, however, are not very useful except for greases and oils, and we must do something about rearranging the atoms in these heavy molecules to make them more useful.

This is done by heating the large, heavy molecules under pressure in an oven to break them apart—crack them down into smaller pieces. Then we put the cracked pieces through a separation process in the still and obtain relatively pure fractions of the various cracked fragments.

Common cracked fragments that we can obtain in this way are fragments containing 4 carbon atoms each, one called iso-butane and the other iso-butene.

The difference between the two is that the iso-butene contains in it a double bond, two carbon atoms tied together by



two bonds instead of one.

This double bond is a very strained affair and tends to come open to add other atoms at every opportunity. When a catalyst is provided, it may open up and then combine with iso-butane to give an eight-carbon molecule, octane.

In this case, it is iso-octane, the principal constituent of what we call high-octane gasoline.

Another way in which the cracked fragments can be reconstructed is again to make use of the fact that this double bond is in a very strained condition and tends to come open easily.

By providing another type of suitable catalyst, one can open up the double bond and allow it to combine with another molecule of the same sort in which the double bond also has been opened. This pair can combine with a third similar molecule, and then with a fourth and a fifth and so on.

Such molecules can be quite large and are called polymers because they are made by the process of "polymerization" of the single unit containing one double bond.

Polymerization simply means the combining of two or more molecules of the same kind into a larger compound with

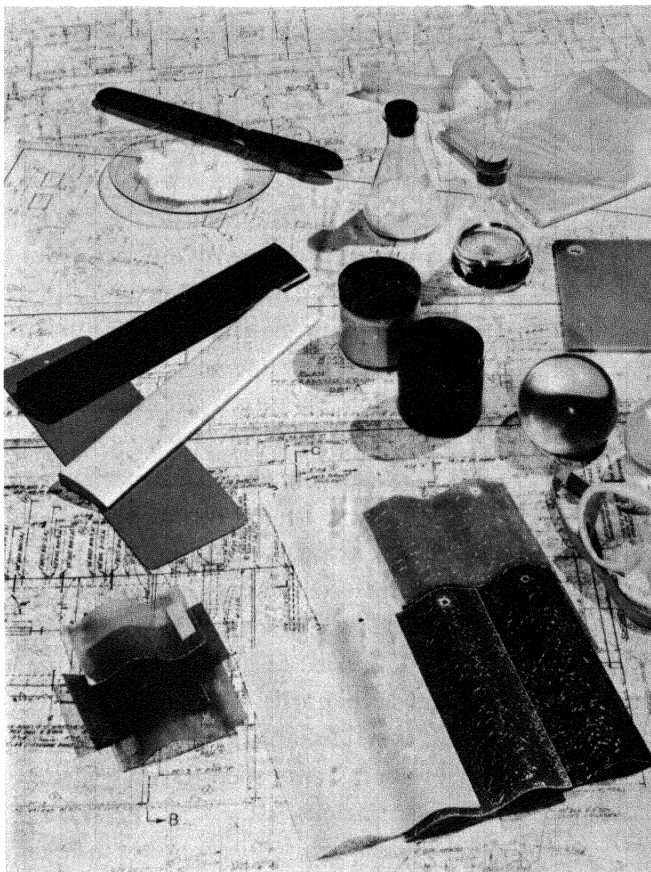
different physical properties.

Familiar products made by this process are polyethylene, made from two-carbon compounds containing one double bond, and polystyrene, made from molecules very closely resembling ethylene but having a large six-carbon ring attached to each.

In addition to many fuels and structural materials, drugs like sulfanilamide are made from the products of coal tar and petroleum.

Thus, by distilling, cracking, and recombining the molecules of crude petroleum, man has learned to make a great variety of useful products from the natural compounds drawn from deep in the earth.

MELVIN CALVIN



THE NUCLEUS

The heating of sugar, the reduction of metals from their ores and the cracking of petroleum molecules are all examples of chemical reactions.

A chemical reaction—a flame, for example—involves the rearrangement of atoms and the exchange or release of the energy involved in this rearrangement.

A nuclear reaction, however, is quite a different matter.

A nuclear reaction involves the rearrangement of subatomic particles within the nucleus of an atom and the exchange or release of energy from within the nucleus. And since any particular element is defined by the number of protons and neutrons in its nucleus, a change in the number of either of these basic particles means the transmutation of the element into a different isotope or into an entirely different element.

Since more than 99.9 per cent of an atom's weight is concentrated in its incredibly small nucleus, this nucleus is a very heavy speck of matter. If the entire volume of an atom were as dense as its nucleus, a golf ball made of such atoms would weigh several billion tons.

The core of a proton or neutron is probably even more dense than this.

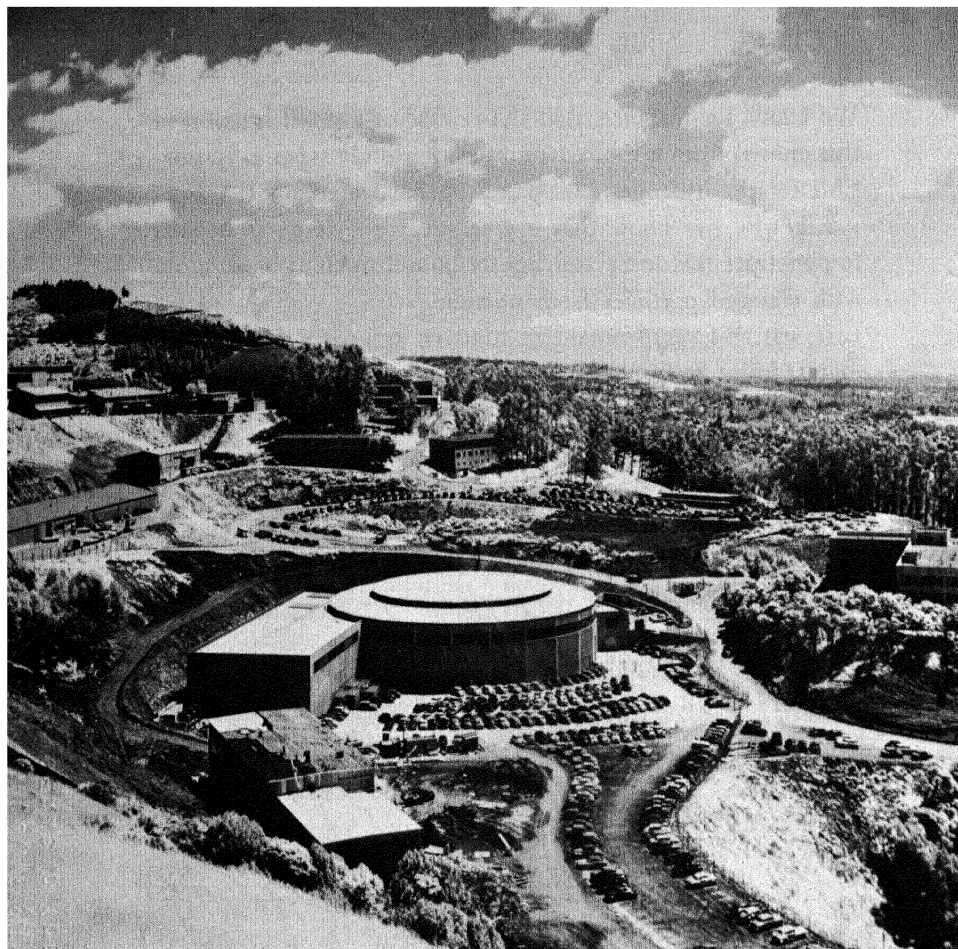
It is perhaps not difficult to understand that a great deal of energy is needed to bind such particles so closely together in a nucleus.

Although we do not yet understand these nuclear forces, we do have an idea of the amount of energy tied up in them.

We know, for instance, that the release of a small fraction of this energy from a limited number of atomic types is responsible for the hydrogen bomb and the nuclear reactor.

Likewise, we know that a great deal of energy is needed to penetrate the dense and tightly bound nucleus of an atom with charged particles from outside.

It was this requirement of high energies, combined with man's desire to investigate the atomic nucleus deliberately, that led to the invention of the cyclotron and other huge particle-accelerators.



High in the Berkeley hills, overlooking San Francisco Bay, is the Radiation Laboratory of the University of California.

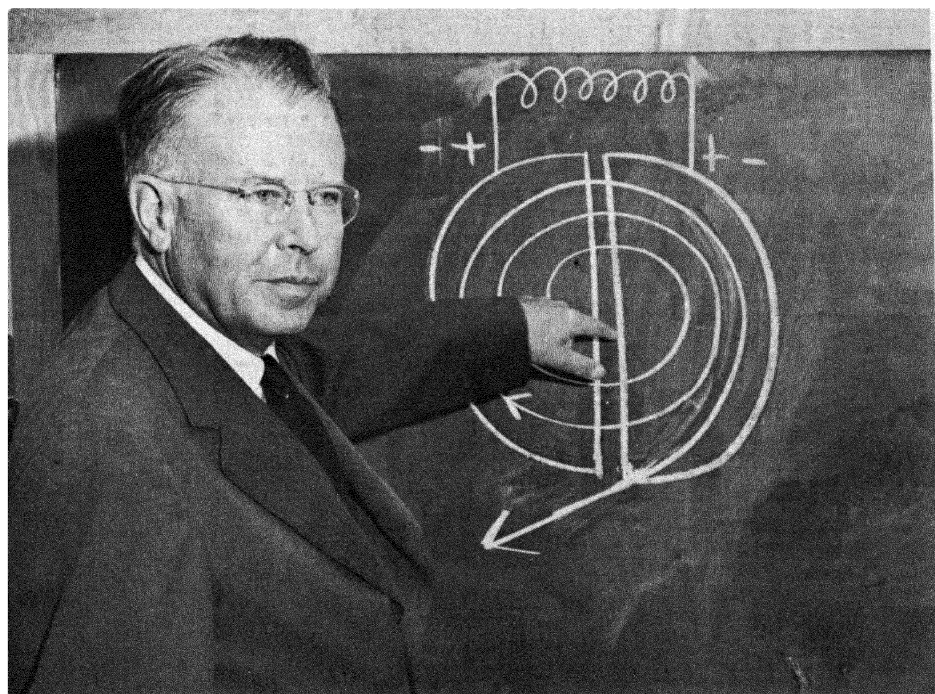
Here are located instruments of modern alchemy with which man has realized the ancient alchemist's dream of transmuting one element into some other element.

With these tools, we have gone even a step further, fashioning entirely new forms of matter by creating elements which do not exist naturally on the earth.

The Berkeley "Rad Lab," where most of the synthetic elements have been produced or identified, is one of the laboratories responsible for this new phase in man's relationship to nature. Its principal tools are the particle accelerators like the bevatron in the foreground and the 184-inch cyclotron above it, all of them involved in fundamental research for the Atomic Energy Commission.

The cyclotron is the atom-smasher which has been in on the birth of practically all of the synthetic elements.

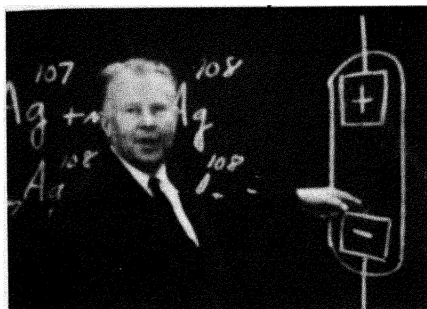
To describe the history and operation of the cyclotron, we called on the late Ernest O. Lawrence, founder of the U.C. Radiation Laboratory in Berkeley and its director for twenty-two years. Dr. Lawrence invented the cyclotron and received the 1939 Nobel Prize in physics for this achievement.



HOW TO MAKE A CYCLOTRON

More than thirty-five years ago Lord Rutherford discovered that nitrogen could be transformed into oxygen by bombarding it with swiftly moving helium nuclei, the alpha rays given off by radium.

Since that time, scientists have been seeking ways to accelerate various atomic particles to very high speed for nuclear investigation.



In the 1920's, the first particle accelerators were simply high-voltage, high-vacuum discharge tubes.

We have a simple vacuum tube containing two electrodes; one has a positive potential and

the other a negative potential, perhaps a million volts.

At the positive electrode, positive particles such as alpha particles are generated. They are allowed to "fall downhill" to the negative target electrode, picking up kinetic energy and producing the nuclear effects and the radiations coming off the negative electrode.

In those days it was found that perhaps such a particle-accelerator would be able to produce only one or two million-volt accelerated particles. And it was necessary to consider means of reaching much higher energy—tens of millions of volts or hundreds of millions of volts.

Now every child knows that in a swing one can reach a high swinging velocity either by one big push—as in a high-

voltage accelerator—or by a succession of small pushes, pumping up in a swing.

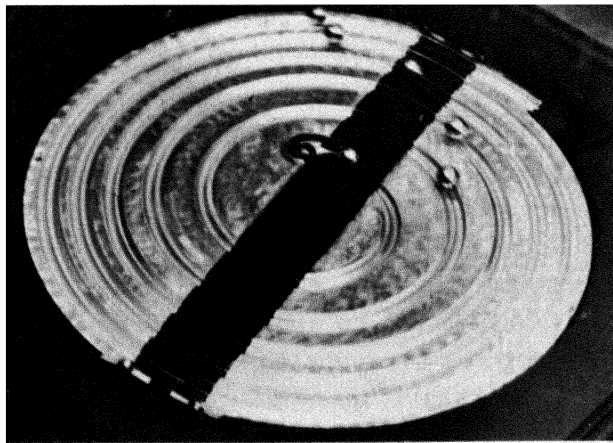
The cyclotron, which was invented in 1929, makes use of this second principle. It gives particles successive pushes in time with their oscillatory motion.

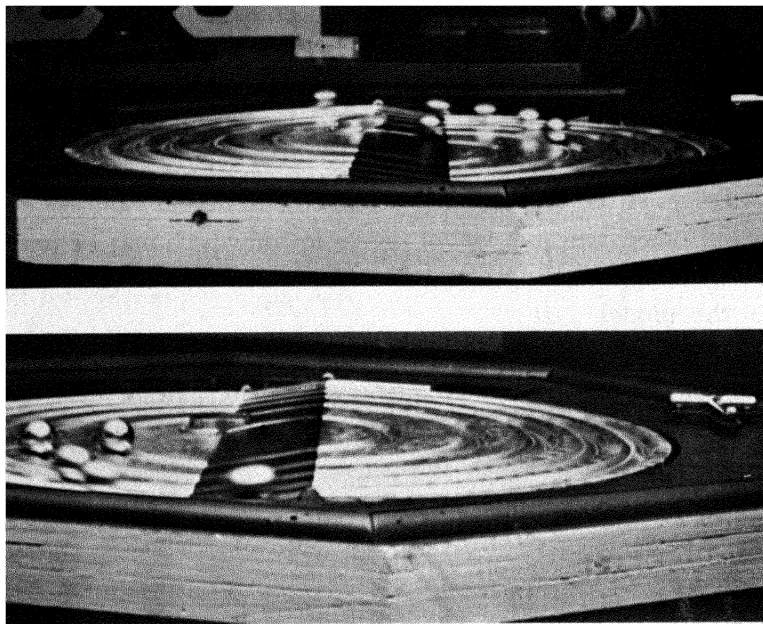
Dr. F. A. B. Ward of the Science Museum in London devised a mechanical model of the cyclotron which may be helpful in understanding how it works. Illustrated here is a copy of this model.

There are two semi-circular electrodes, or “D’s,” in the vacuum chamber of the cyclotron which alternate in electrical potential; they move “up and down” in potential, so to speak.

In the model, which is only an analogy, the two plates actually do move up and down. Acceleration in the model is by gravity instead of by a difference in electrical potential. Also, in the model, the particles run in grooves whereas the particles in a real cyclotron follow circular paths because of a strong magnetic field.

The particles are generated at the center and are accelerated by falling “downhill” into one of the “D’s.”





When they cross back to the second "D," they are again going "downhill" and therefore receive an additional push.

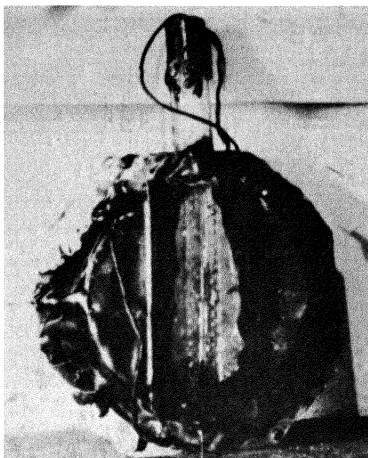
The two "D's" are always alternating—up and down—so that the particles always cross the region between them "downhill."

Each particle therefore repeatedly picks up energy, filing out in ever-widening circles until, at the periphery, it strikes the target. The target, of course, is the point where our nuclear investigations take place.

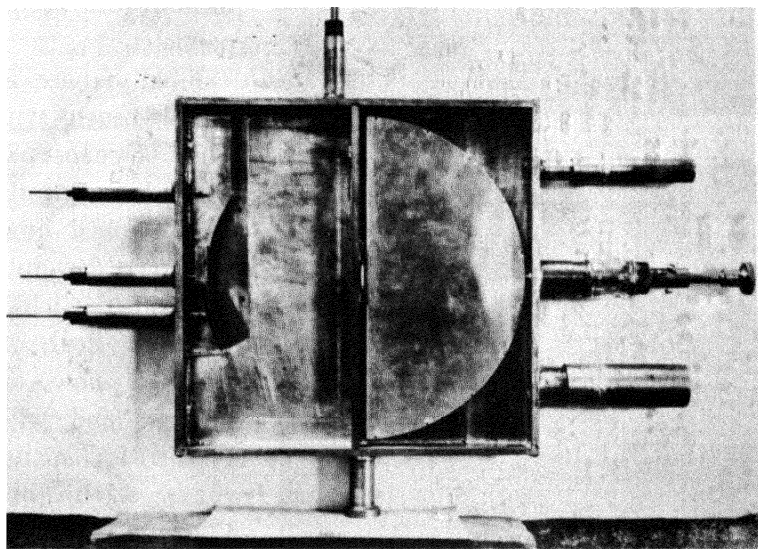
In the actual cyclotron, the particles are caused to go in their circles by a large magnet, and the "up and downhill" is provided by an oscillating electric field—a radio frequency voltage that reverses in step with the circulation of the particles. The particles go around hundreds of times and therefore acquire hundreds of times as much energy as that which corresponds to the maximum accelerating voltage in the

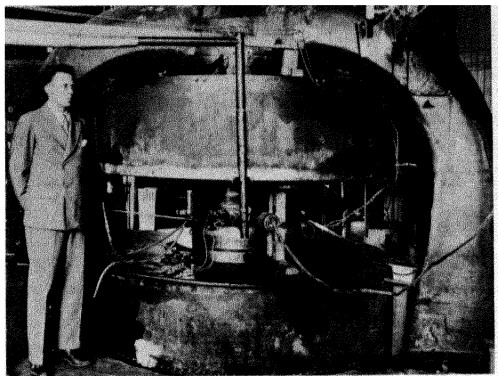
cyclotron chamber at any one time.

Here, without the magnets, is our first attempt to build a cyclotron in 1930. It did work a little bit, and so we naturally kept it as a souvenir. The two electrodes are waxed together, and they would be placed between the poles of the magnets. Next was built a better model about the same size—eight inches—which worked very well.

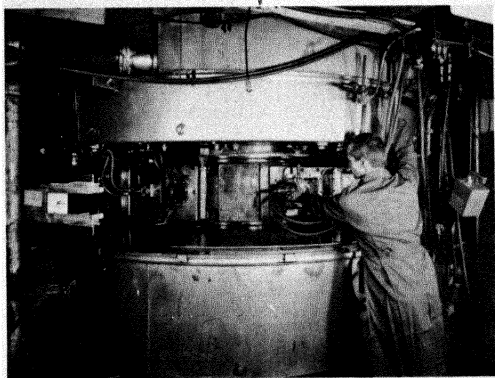


Then followed an 11-inch machine which produced one million-volt protons, and these were used, actually, in some nuclear experiments. Indeed, the first cyclotron-produced nuclear disintegrations were produced by this 11-inch cyclotron.

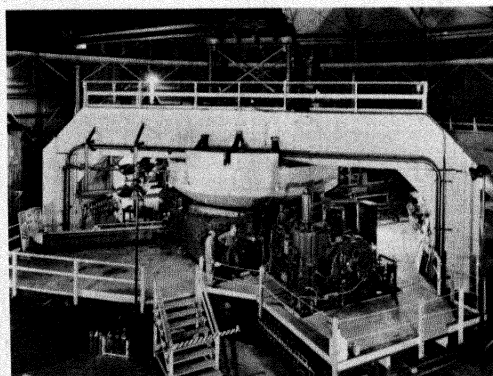




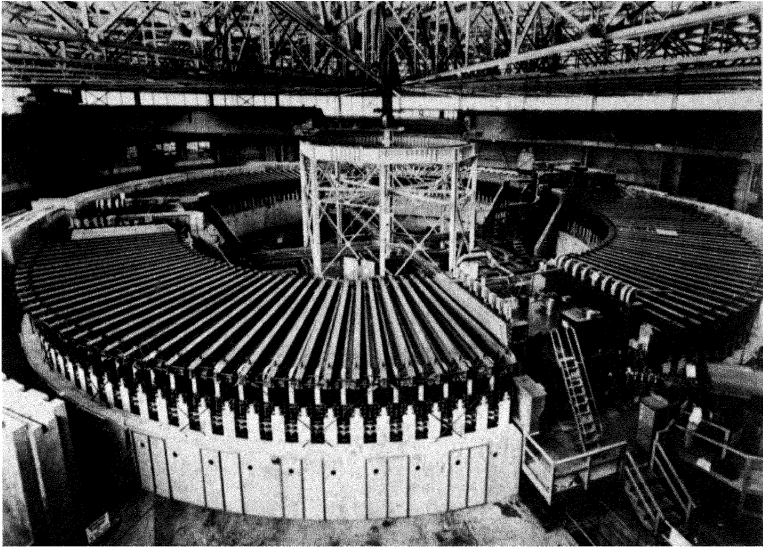
Next came the 27-inch cyclotron, shown above, and then the 37-inch machine. The former produced four or five-million-volt particles, and many new radioactive isotopes were made with it.



These were followed by the 60-inch cyclotron, left, which is still being used today. The 60-inch produces up to 50-million-volt alpha particles.



Finally there is the 184-inch synchro-cyclotron which was rebuilt in 1957 and now produces 720-million-volt particles. A great deal of work is now in progress and will go on for many years with this 4,000-ton machine.



A much larger accelerator in the laboratory is the bevatron. It isn't a cyclotron but it has many characteristics in common with the cyclotron.

The bevatron was invented by Edwin M. McMillan and it derives its name from the expression "bev" because it accelerates atomic particles to 6.2 *billion electron volts*.

The bevatron is housed in a very large circular building and the particles go around in a racetrack-like magnet more than 100 feet in diameter. This is quite a jump from the original 8-inch cyclotron built less than twenty-five years before the bevatron.

We haven't seen the end of this yet. A much larger accelerator is being built at Brookhaven, Long Island, and I'm sure that in the future accelerators will go to the region of 100 billion electron volts.

ERNEST O. LAWRENCE

The cyclotron has been a key instrument in the production and discovery of synthetic, man-made elements.

But the story of the man-made elements really begins in 1925 when the last of the eighty-eight natural elements was discovered. In one sense, the Periodic Table was complete at this time, as shown on page 85.

The heaviest natural element is uranium, number 92. But gaps in the Periodic Table of 1925 reveal four elements not accounted for. These four missing elements are numbers 43, 61, 85, and 87, now known as technetium, promethium, astatine, and francium. Some of them were reportedly discovered before 1937, but these "discoveries" proved to have been erroneous.

All of them are unstable. They must have been present when the earth was formed some five billion years ago. However, they have disappeared meantime due to the radioactivity which is characteristic of atoms with unstable nuclei.

Such unstable elements can be created artificially by the transmutation of lighter and more stable elements.

As to the nature of instability, radioactivity, and transmutation, these are things which can be explained only by examining the structure of the atomic nucleus.

On page 31 we saw that the nucleus of an ordinary hydrogen atom consists of a single particle, a proton, represented by a circle enclosing a plus sign which indicates that it has a single charge of positive electricity.

If a neutron is added to the proton, we have the nucleus of an atom of the hydrogen isotope, deuterium, or heavy hydrogen.

The addition of a second proton and a second neutron gives us the nucleus of a helium atom. A third proton and a third or fourth neutron results in a lithium nucleus, and by adding more and more protons and neutrons we can build up the nuclei of heavier and heavier elements.

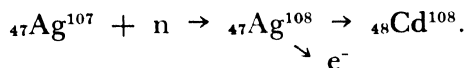
The nucleus of one isotope of silver, for example, contains 47 protons and 60 neutrons.

This silver nucleus, containing a total of 107 neutrons and protons, has an atomic weight of 107. We can represent the silver nucleus by the chemical symbol for silver, Ag, plus the number 47 to indicate its atomic number—which is the same as the number of protons in the nucleus—and the number 107 to indicate its atomic weight: ${}_{47}\text{Ag}^{107}$.

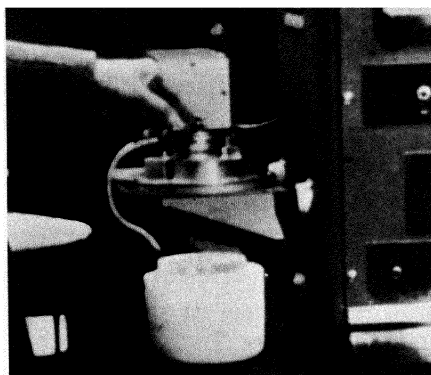
Now if we bombard silver with neutrons, this is what happens: the silver nucleus captures a neutron, which increases its atomic weight so that it becomes: ${}_{47}\text{Ag}^{108}$.

But this heavier isotope of silver is radioactive. Spontaneously, a neutron in the nucleus changes into a proton, gaining the positive charge which makes it a proton and emitting an electron, or single negative charge.

By gaining an additional proton, the nucleus has attained the atomic number 48, which makes it the nucleus of the next heavier element, cadmium. In this case, the cadmium is the isotope, cadmium-108. Using the symbols “n” for neutron and “e” for electron, the transmutation of silver into cadmium by bombardment with neutrons can be written:



This is not a difficult experiment to perform on a modest scale. We can bombard the silver in an ordinary 50-cent piece with neutrons by putting it in a miniature atom-smasher which has a neutron source of radium mixed with beryllium.



After several minutes, the half-dollar becomes radioactive enough to register clearly on a Geiger counter. This radioactivity is due to the spontaneous transmutation of bombarded silver atoms into atoms of cadmium. As in any nuclear bombardment, only a minute percentage of the target atoms are transmuted.

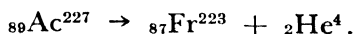
Three of the four elements missing from the Periodic Table of 1925—technetium, promethium, and astatine—were produced synthetically by transmutation in a way similar to that by which the silver was transmuted into cadmium. The fourth “missing element,” francium, was first discovered by a rare observation of natural alpha-decay in the element actinium.

“Alpha-decay,” a type of radioactivity, means the loss or emission of an alpha particle. An alpha particle consists of 2 protons and 2 neutrons and is really just another name for the nucleus of a helium atom.

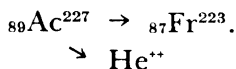
Actinium is one of the radioactive elements. It has the atomic number 89 and a weight of 227. In a rare form of decay discovered in 1939 by Mlle. Marguerite Perey of the Curie Institute in Paris, actinium emits an alpha particle, losing two protons and thus becoming a new element with the atomic

number 87. Mlle. Perey named it francium after her native country.

One way of writing the reaction is as follows:



Another way of writing the same reaction indicates the emission of the alpha particle from actinium more clearly:



Actinium, with 89 protons and a weight of 227, decays into francium, with 87 protons and a weight of 223, losing the alpha particle with its 2 protons and weight of 4.

Francium, then, does occur in nature as a result of radioactive decay, but it is indeed rare. The longest-lived isotope of francium has a half-life of only twenty-one minutes.

A "half-life" is the time required for half of the atoms of any radioactive isotope to undergo radioactive change.

The three other "missing elements"—technetium, promethium, and astatine—were produced as transmutation products by synthetic means, either in the cyclotron or the atomic reactor.

One of our big difficulties is that of identifying a new element once we have produced it. Elements 43, 61, 85, and 87 were identified by what we call the "tracer" technique. This is a procedure used to separate and concentrate an element which is present in a sample in very sparse quantities.

Radium, for example, may be present in unweighable amounts, amounts too small to be seen; it is only detectable at all in such a case because of its radioactivity. However, the radium can be "carried" out of solution by one of its chem-

ical "cousins" such as barium or strontium which appear above radium in the Periodic Table.

These related elements will carry the traces of radium from the solution, and it is by methods such as this that the chemical identification of technetium, promethium, astatine, and francium were made.

The first synthetic element to be produced and identified was technetium, discovered by Emilio Segrè, who is now Professor of Physics at the University of California, and a colleague, Carlo Perrier.

Dr. Segrè's description of the discovery follows.



TECHNETIUM MEANS "ARTIFICIAL"

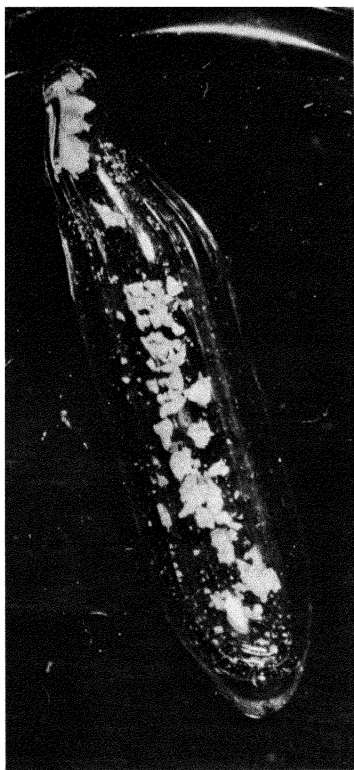
Many years ago, in 1936, after I had been working with radioactivity for several years, I moved to Palermo from Rome, where I had been associated with Enrico Fermi.

In Palermo they had a very poor and small laboratory, nothing like the large laboratories around Berkeley, and I wanted to find some research to do. During the summer of 1936 I had been in Berkeley to visit the 37-inch cyclotron, the little cyclotron we aren't using any more. The cyclotron was used to accelerate deuterons, and I took some pieces of it—which were made of molybdenum and had been exposed to deuterons—back to Palermo with me.

We had very good reason to think that molybdenum bombarded by deuterons would make element 43, which we now call technetium. Molybdenum has 42 protons in its nucleus. A deuteron, the nucleus of a deuterium or heavy hydrogen atom, has 1 proton and 1 neutron. Technetium, of course, has 43 protons in its nucleus.

We thought that the deuteron might be expected to leave 1 proton in the molybdenum and thus make technetium, but the problem was to prove that this hypothesis was true. This was done by the tracer technique. We took the molybdenum target, dissolved it, and added many elements as tracers. Then we tried to see whether there was anything to be found which was a chemical cousin of rhenium and manganese, both of which are in the same vertical column as technetium in the Periodic Table.

Well, after many operations, we succeeded in 1937 in finding a substance which we could prove chemically was not one



of the known elements, although it was very similar to rhenium; ultimately we separated it even from rhenium.

This was the first man-made element, and for this reason we called it technetium, which means "artificial." In this work Carlo Perrier, who was a mineralogist, helped me, and the problem was quite natural for a mineralogist because, at least with respect to effort, it was similar to digging an ore out of a mine.

In 1938 Dr. Seaborg and I, working in Berkeley, found a longer-lived form of technetium with a half-life on the order of 200,000 years, and Miss Chien-Shiung Wu and I found

the same substance among the fission products of uranium.

This sample of a technetium compound is an appreciable fraction of the world's present supply of technetium; and the picture is about three times life size. It could not have been made with the methods used in 1938. It was made in a fission chain reaction at the Oak Ridge National Laboratory.

The second synthetic element to be discovered was astatine with the atomic number 85.

Astatine could not have been made using the small cyclotron of 1936. But as soon as a bigger one became available it was possible to bombard bismuth with alpha particles of sufficient energy to penetrate the nucleus.

An alpha particle contains 2 protons and 2 neutrons. If the 2 protons could be put into the nucleus of a bismuth atom—which has 83 protons—then we make the astatine with 85 protons. The way to do this is as follows. Electrons are stripped from atoms of helium, leaving the helium nuclei, or alpha particles. The alpha particles are accelerated in the cyclotron, which starts them going around, faster and faster, until they ultimately strike the bismuth atoms in the target. When an alpha particle enters the nucleus of a bismuth atom, two neutrons are shot out and you are left with an atom of astatine.

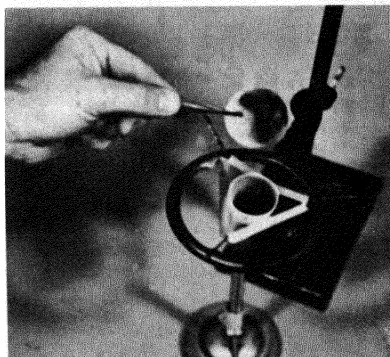
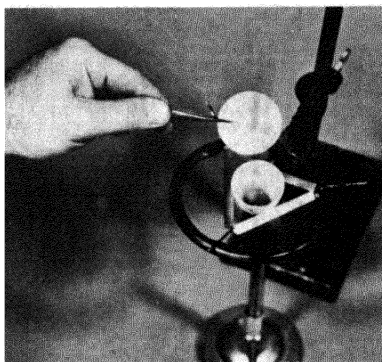
Then, of course, you have the problem of chemically identifying the astatine.

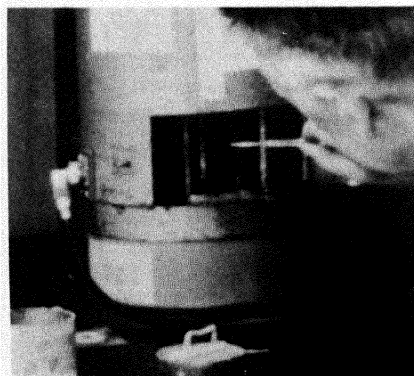
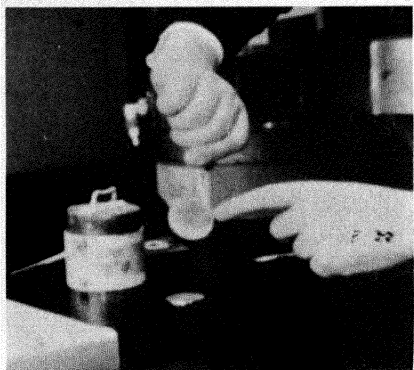
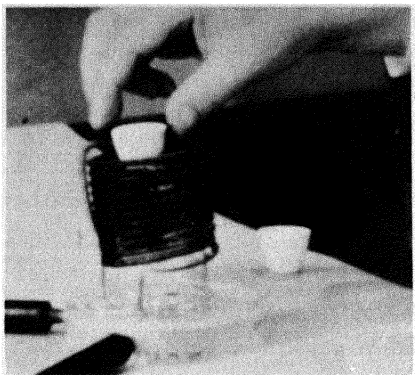
This is much easier than it was with the technetium. Astatine is very similar to iodine, which is volatile.

Astatine can be separated from bismuth by volatilizing it, although the discovery experiments performed in 1940, of course, were much more complex than this.

In the case of iodine, it can be put in a crucible and heated, and iodine vapor can be collected on the surface of a platinum plate.

We can do practically





the same thing with astatine. A piece of bismuth which has been bombarded with alpha particles, and which thus contains some astatine, is put into the crucible.

When the crucible is heated, the astatine will vaporize just as the iodine vaporized.

The astatine vapor is invisible, but since it is radioactive we can count its nuclear disintegrations with a Geiger-Mueller counter.

My colleagues in this astatine experiment were Dale R. Corson and K. R. MacKenzie, who were working in Berkeley at the time.

We named the element astatine because the halogens are named according to some property: chlorine, green; bromine, stinking; iodine, violet.

Astatine cannot be made into anything that

can be seen or detected in any way except by its radioactivity; for this reason it was called astatine, which means “unstable.”

Astatine is like the halogens in its chemical properties and like iodine in particular. In fact, before its discovery it was tentatively named eka-iodine. It even behaves physiologically like iodine and concentrates in the thyroid gland in animals and humans as iodine does.

EMILIO SEGRE

The fourth of the "missing elements" was number 61, which was named for Prometheus, the mythical Greek Titan who stole fire from the gods for the benefit of man.

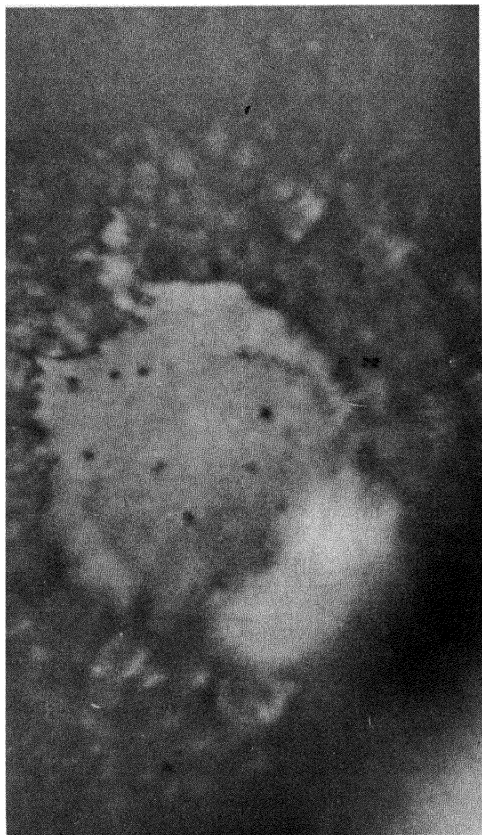
Promethium was discovered as a result of its artificial production in the atomic reactor rather than in the cyclotron. The identification was made in 1945 by three chemists at the Oak Ridge National Laboratory, Lawrence E. Glendenin,

J. A. Marinsky,
and Charles D.
Coryell.

It is a "rare earth" and therefore similar to the other "rare earth" or "lanthanide" elements.

Promethium nitrate is a rather undistinguished-looking powder. This actual-size photograph shows an appreciable fraction of the world's supply of this synthetic element.

The four "missing elements" are, for all practical purposes, missing



in nature because they are radioactive and thus unstable.

They are probably useless, or of limited use, from a practical point of view but of great interest from the point of view of atomic structure and nuclear structure.

BEYOND URANIUM

The discovery of the four “missing elements” completes the Periodic Table through uranium.

But this is only the start of the adventure of forging new forms of matter. The achievements which followed have been exciting triumphs over nature and have influenced profoundly the world we live in.

Beyond uranium are the highly radioactive elements with atomic numbers higher than 92. None of them exists naturally on earth except in extremely minute quantities in one or two cases. They can be produced only by synthetic means.

Actually, all elements heavier than lead and bismuth are radioactive and are constantly disintegrating. Eventually, lead and bismuth will be the heaviest natural elements on earth, for the heavier elements—polonium, radon, radium, actinium, thorium, protactinium, and uranium—will have disappeared at some date in the distant future.

The present existence of these elements is one of the indications of a finite age for the earth, about five billion years according to current estimates.

The elements beyond uranium are increasingly unstable. The half-life of the plutonium isotope Pu-239, for example, is 24,000 years; the half-life of mendelevium is about half an hour.

This is the Periodic Table as it looked twenty years ago, before World War II, when scientists first tried to produce elements beyond uranium. Technetium (Tc), promethium

H																	He	
Li	Be											B	C	N	O	F	Ne	
Na	Mg											Al	Si	P	S	Cl	A	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
Fr	Ra	Ac	Th	Pa	U	93	94	95	96	97	98	99	100					
		↓																
		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu		

(Pm), astatine (At), and francium (Fr) are included here although they actually were given their names later, and some were discovered at a later date.

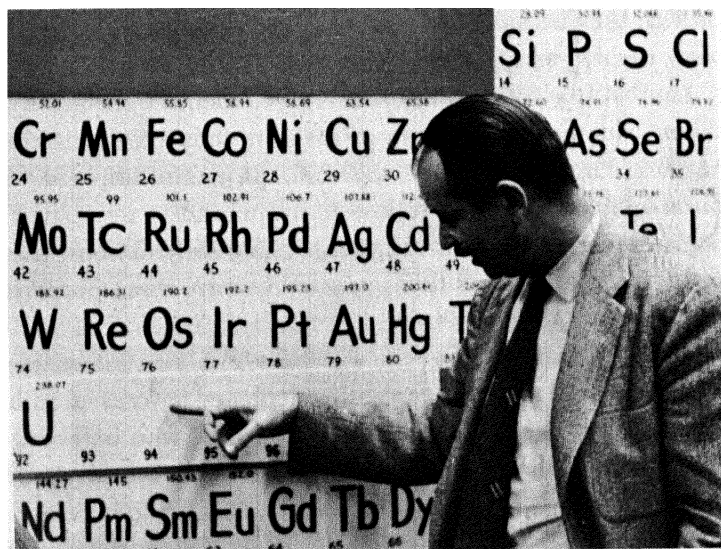
Lanthanum and the “rare earth” elements were fitted between barium and hafnium as they are today. But the three heaviest known elements—thorium, protactinium, and uranium—were believed to follow actinium and to be related to hafnium, tantalum, and tungsten, respectively. The next element, number 93, was thus expected to have some properties resembling those of rhenium. Similarly, elements 94 to 100 were expected to fit into the Periodic Table as shown here.

The first attempts to produce elements beyond uranium were made by Fermi, Segrè, and co-workers, who bombarded uranium with neutrons in Italy in 1934. They actually found a number of radioactive products; these seemed to have chemical properties that you might expect in such a “transuranium”

element as number 94 or 96, properties similar to those of elements listed directly above 94 and 96 on the table of that time.

However, subsequent work—especially the discovery of nuclear fission by Otto Hahn and F. S. Strassmann and their co-workers in 1938—showed that the interpretation of Fermi and Segrè was not correct. It revealed that radioactive products of the Fermi-Segrè experiments were actually radioactive isotopes of lighter elements, fission products like iodine or tin.

The discovery of an element with an atomic number higher than 92 came in 1940 as the result of the work of Edwin M. McMillan, now director of the Berkeley Radiation Laboratory, and Philip H. Abelson. They bombarded uranium with neutrons and made the first positive identification of element 93. This production of the first transuranium element will be described here by Dr. McMillan, who shared the 1951 Nobel Prize in chemistry with Dr. Seaborg for this and related discoveries.



NEPTUNIUM

The story of the first transuranium element goes back to the time when word of the great discovery of nuclear fission first came to this country.

Everyone was excited about this. Everybody tried to think of some simple experiment that they could do to find out something about it.

The particular one I thought of and performed was an experiment to measure how far the fragments into which the uranium atoms split will travel in matter.

To do this I made a thin layer of uranium oxide on a piece of paper. Next to this I placed very thin sheets of inert material to stop the fragments resulting when the uranium underwent fission.

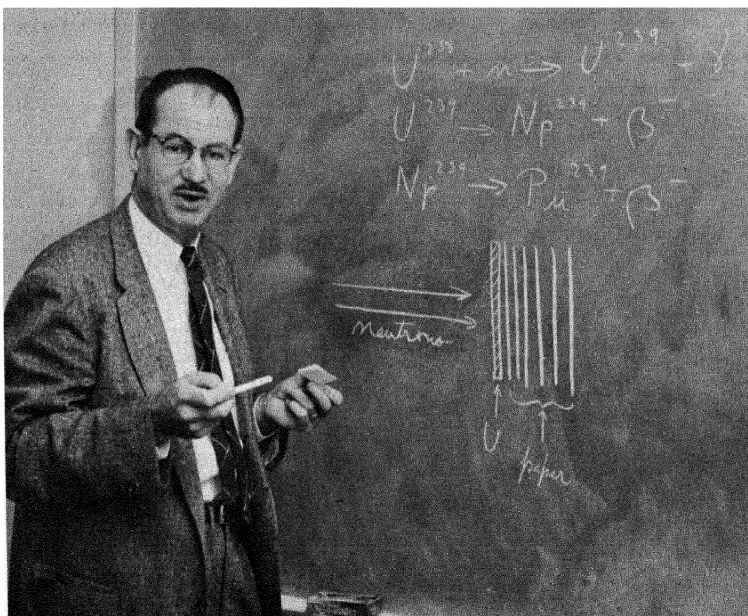
I used ordinary cigarette papers, the kind that people who roll their own use. I stacked them up in a sort of little book and put the uranium oxide on top.

This book of cigarette papers and uranium oxide I used as a target in the cyclotron. When it was exposed to bombardment by neutrons in the cyclotron, the fission fragments flew into the book of cigarette papers, stopping at various depths.

Then all one had to do was to take the cigarette papers apart and measure the radioactivity in each one by means of a Geiger counter.

This is about as simple an experiment as you can imagine. What turned out, of course, was that I got the result I was looking for. But it also turned out that a by-product was much more important than the initial experiment.

The by-product in this case was that the first piece of paper with the uranium itself on it also contained a radio-



activity which had a different half-life and different properties from those of the fission products in the other pieces of paper.

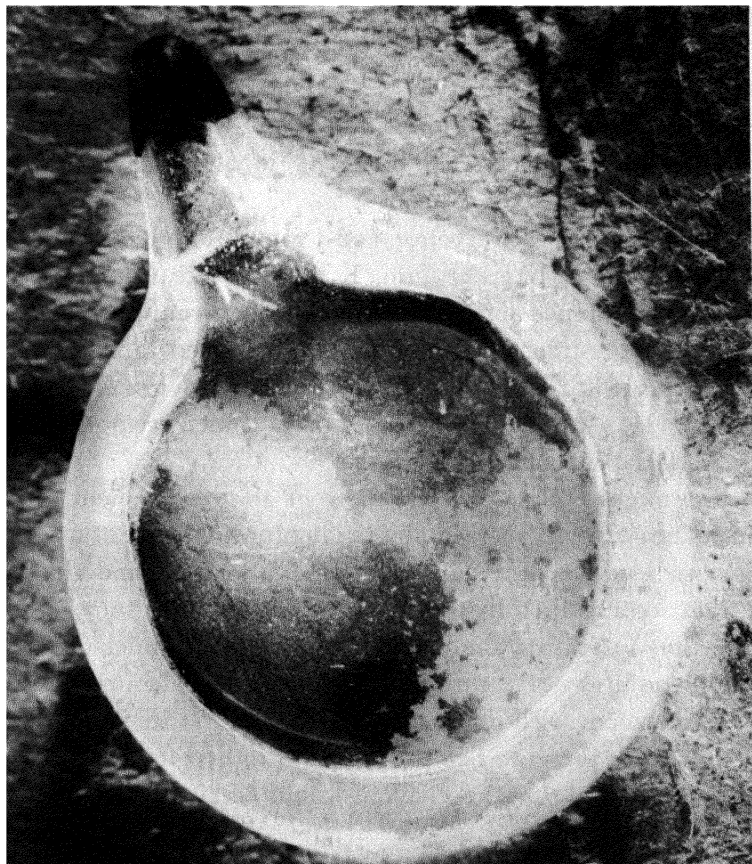
One tries to think: why should this be? Why should some of the radioactivities stay behind while the others fly out?

Well, you think that there may be another process which accounts for that. For example, suppose that a neutron is simply captured by a uranium atom and that the uranium then does not break apart. This process was in fact already known, and the resulting radioactive uranium was found on the piece of paper. But there was also another radioactivity with a different decay period that was not previously known, and that I suspected might belong to a new element, produced by the decay of the radioactive uranium.

Involved in following up this idea and trying to identify this other radioactivity was an old friend and colleague of

mine, Philip Abelson of the Carnegie Institution in Washington, who came to Berkeley for a summer vacation.

It turned out to be a rather arduous vacation because he joined me, and together we worked out the chemistry of the new radioactivity that stays behind and that is found on the first sheet of paper. We were able to show that it had chemical properties different from those of any other element known at that time. This was the element neptunium with the atomic number 93, shown here in a flask the size of a dime.



Neptunium was named after Neptune and was the second element to be named after a planet. Uranium had been named in 1789 in honor of the discovery of the planet Uranus eight years earlier.

Because of the unexpected chemical properties which we found for neptunium, this discovery showed that the Periodic Table of the Elements needed some changes.

Of course, it was known that there should be an element with one more proton than uranium. But it had been supposed that the properties of this new element would be like those of rhenium, following the usual rule of the Periodic Table as it was understood at that time. The Table at that time showed the element with the atomic number 93 directly below rhenium (as on page 129), and we expect to find similar properties among elements in the same vertical column.

As a matter of fact, Abelson and I found that neptunium is not in the least like rhenium. It is more like uranium. In fact, the difference between neptunium and uranium is only a rather subtle difference. And this fact soon led to a revision of the Periodic Table in the region of the heaviest elements.

After Dr. Abelson left Berkeley, I continued trying to find the next heavier element, number 94.

You might ask, why should there be another element present? The answer is very simple: the neptunium nucleus undergoes radioactive decay by emitting an electron, and this is always accompanied by the changing of a neutron into a proton, thus increasing the positive charge on the nucleus.

The addition of a proton to element 93, of course, gives us the element with the atomic number 94. So we knew that 94 had to be present. The problem was to find it.

We suspected that this element would decay by emitting an

alpha particle rather than an electron, and this made it a little hard to detect by the methods we were using at that time.

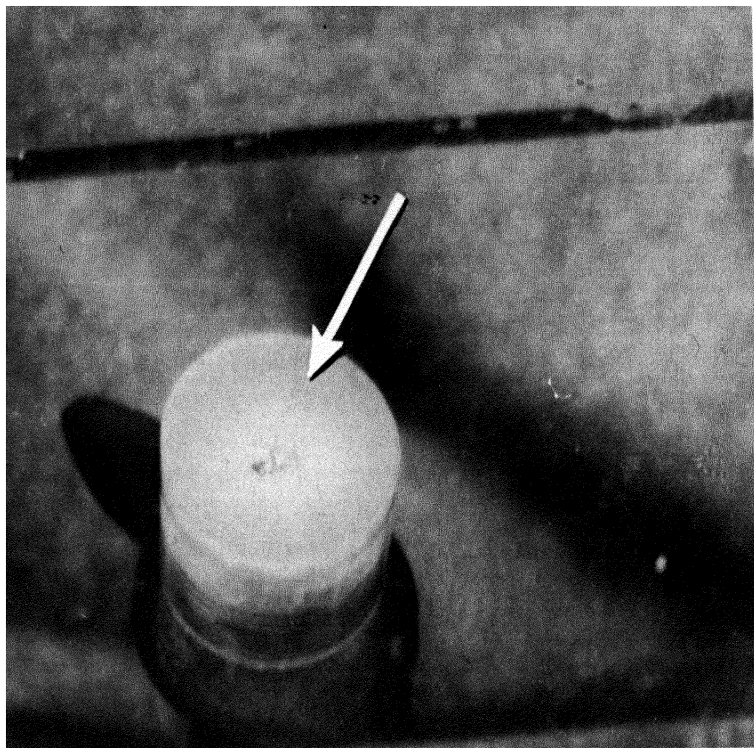
I found out a little about the chemistry involved, and I did convince myself that I had found some alpha particles. But we were about to get into war, and this work was not completed when I was called away to go into radar development.

The work was taken over then by Dr. Seaborg.

EDWIN M. McMILLAN

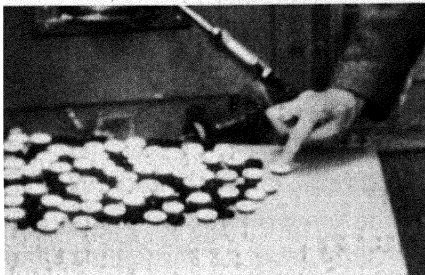
The work on element 94 was continued by Glenn T. Seaborg, Arthur C. Wahl, Joseph W. Kennedy, and Emilio Segre. Actually, the first positive identification of any isotope of the new element was made by bombarding uranium in the Berkeley 60-inch cyclotron with deuterons—hydrogen nuclei each containing a neutron as well as a proton—rather than with simple neutrons.

The element, which is the next element after neptunium, was named for Pluto, which is the planet beyond Neptune. Here is the first visible and weighable sample of plutonium ever made, a piece the size of a pinpoint, which is now mounted on a plastic cylinder as a souvenir.



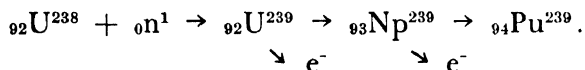
If we represent part of a uranium nucleus by a group of black checkers (protons) and white checkers (neutrons), we can get a crude picture of the nuclear reactions involved.

McMillan and Abelson added a neutron, leaving an unstable form of uranium which decayed when a neutron in the nucleus spontaneously emitted an electron and thus changed itself into a proton. This gave us element 93, neptunium.



Then another neutron emitted an electron and changed into a proton, creating element 94, plutonium.

In the technical language of nuclear physics, this series of reactions can be written:

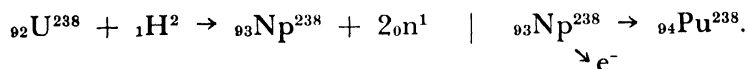


Uranium-238 absorbs a neutron, making the unstable heavy uranium isotope, U-239. The U-239 emits an electron and becomes neptunium-239. Then another electron is emitted and leaves us with plutonium-239.

The symbol for a neutron is written ${}_0\text{n}^1$ simply to indicate that the neutron has no charge but has an atomic weight of about one. The minus sign after the “e” indicates that we are losing one charge of negative electricity or, which is the same thing, gaining a charge of positive electricity.

The reaction for the first production of plutonium by

bombardment with deuterons would be written:



In this case, the deuteron or heavy hydrogen nucleus is indicated by the symbol ${}_1\text{H}^2$, showing that it has a charge of one due to its single proton but also contains a neutron, giving it an atomic weight of about two.

The most important plutonium isotope is plutonium-239, for Pu-239 has the property of undergoing nuclear fission when bombarded with slow neutrons. This means that it can be used for nuclear power and nuclear weapons.

The key experiment which showed that plutonium-239 undergoes fission when bombarded by slow neutrons was performed with a cyclotron at Berkeley early in 1941.

This experiment can be repeated on a small scale by using a mixture of radium and beryllium instead of a cyclotron as a source of neutrons. The results are detected on an oscilloscope connected to an ionization chamber. The ionization chamber—another detector of radiation like the Geiger counter—is an enclosure containing a positive electrode, a negative electrode and a gas; atoms of the gas can be ionized, or stripped of electrons, by charged particles entering the chamber. The resulting ions—or charged atoms—move to one of the electrodes, creating a current which can be amplified to register visually on the oscilloscope.

When the fissionable Pu-239 is placed in the ionization chamber, the alpha particles emitted by its radioactivity show as small pulses on the oscilloscope.

If the neutron source is now placed under the ionization chamber, an occasional large, bright pulse appears on the

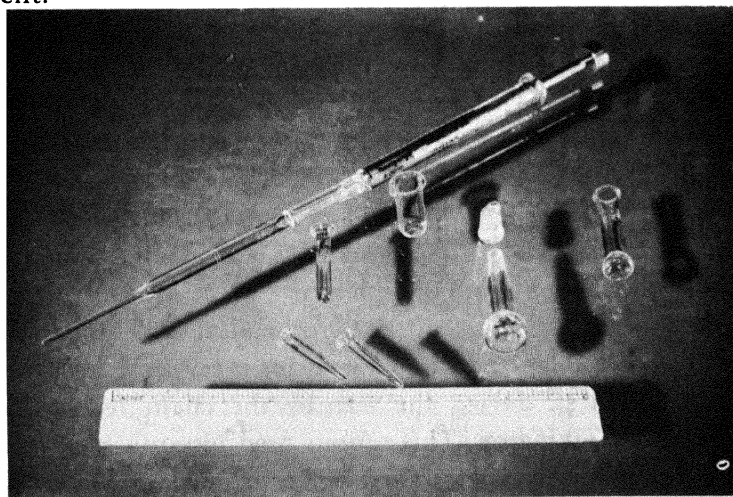
oscilloscope. These bursts of high energy indicate that fission due to bombardment by the slow neutrons is occurring.

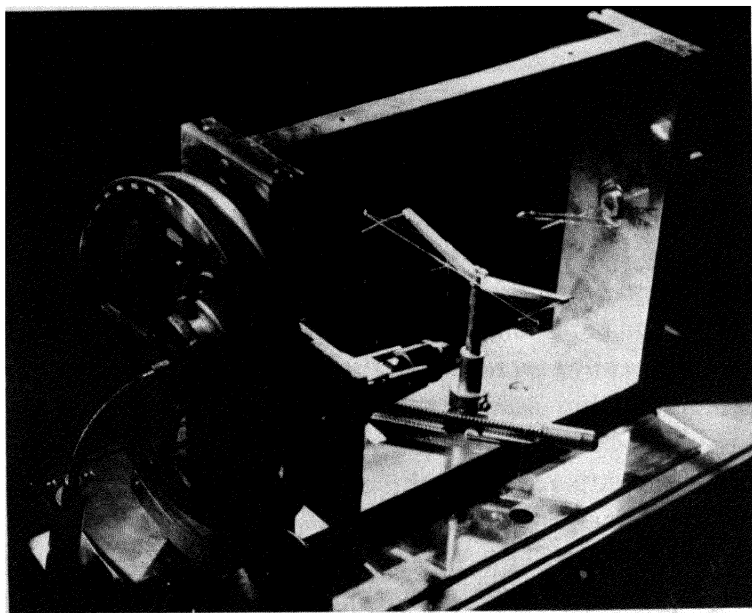
Once it was proved that plutonium-239 has this important property, the problem was to produce it in quantity.

For this purpose, chemists, physicists, and biologists were assembled at the famous wartime Metallurgical Laboratory at the University of Chicago. Here the physicists, led by the late Enrico Fermi, worked out the chain reaction for the mass production of plutonium from natural uranium and graphite.

The chemists worked out the process for separating the plutonium from uranium and from the very radioactive fission products which resulted from the chain reaction.

An interesting aspect of this chemical problem was that the key experiments had to be performed with almost invisible amounts of plutonium, amounts on the order of one ten-millionth of an ounce. This required dissolving the small amounts of material in very small amounts of a solution—a single drop, for instance. And this meant miniature lab equipment.





Balances and centrifuges as well as test tubes and flasks were fashioned to handle pinhead amounts of material. This ultra-microbalance has quartz fibers slimmer than a human hair for its crossbeams and pan holders, and the chemists sometimes spoke of weighing invisible samples on an invisible balance.

Using this ultra-microchemical equipment, Burris B. Cunningham and Louis B. Werner worked out the chemistry of plutonium. They also weighed the first visible sample of plutonium, which is shown on page 136. Its weight? Just under one ten-millionth of an ounce.

Experiments on this ultra-microchemical scale were performed to test the chemical process for separating the plutonium produced during the war by the chain reaction at Hanford, Washington. This process used bismuth phosphate as a carrier material, and it was conceived and worked out in

large part by Stanley G. Thompson.

Altogether there are three outstanding fissionable materials, or nuclear fuels: the "Big Three," plutonium-239, natural uranium-235, and the synthetic uranium isotope, U-233.

The plutonium-239, of course, is made by bombarding natural uranium-238 with neutrons in a nuclear chain reactor, or "pile" as it is sometimes called.

Uranium-233 is produced in a comparable manner by bombarding element 90, thorium, with neutrons.

The uranium-235 exists in ordinary natural uranium to the extent of one part in 140 and can be separated by complicated methods such as those that were worked out at Oak Ridge, Tennessee, during the war.

This brings us up to about 1944 in the hunt for new elements, and it also brings us to a major theoretical roadblock.

BREAKTHROUGH

At the Metallurgical Laboratory of the University of Chicago in 1944, Seaborg, Ralph James, Leon Morgan, and Albert Ghiorso began looking specifically for the next two elements, those with the atomic numbers 95 and 96.

Now the task of producing a new element is often not nearly as rigorous as the problem of discovering whether or not you really *have* produced it.

Seaborg and his colleagues conducted various nuclear experiments calculated to produce element 95 or element 96. Then they followed lengthy and complicated chemical procedures trying to isolate and identify microscopic traces of these new forms of matter. ~

When you are looking for something that's not even known to exist, you still have to have something to go on, some rough idea of how and where to look and what to look for. And these "rough ideas" determine how you design your experiments.

The Periodic Table of the Elements, if it is correct, can tell you what to look for.

Uranium, neptunium, and plutonium were considered to be sort of "cousins" in the Periodic Table as it was understood in 1944, but the family relationship was not clear. It was thought that elements 95 and 96 should be much like them in their chemical properties. Thus it was thought that these elements formed a "uranide" ("like uranium") group.

The Periodic Table of 1944, therefore, implied that the chemical properties of elements 95 and 96 should be very

Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt
Fr	Ra	Ac	Th	Pa	(U-96)				

U	Np	Pu	95	96
---	----	----	----	----

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd
----	----	----	----	----	----	----	----

much like those of neptunium and plutonium.

These assumptions proved to be wrong, and the experiments failed to lead to the discovery of elements 95 and 96.

Then Seaborg conceived the idea that perhaps the elements heavier than actinium were misplaced on the Periodic Table.

McMillan had discovered that neptunium was not at all like rhenium, as it had been expected that it should be. And now the undiscovered elements 95 and 96 apparently refused to fit the pattern indicated by the Periodic Table of 1944.

Seaborg's theory was that these elements heavier than actinium might constitute a second series similar to the series of "rare earth" or "lanthanide" elements; the "lanthanides" are chemically very similar and are usually listed in a separate row below the main part of the Periodic Table.

This would mean that all these heavier elements really

belonged with actinium—directly after radium on the Periodic Table—just as the known “lanthanides” all should fit in with lanthanum between barium (Ba) and hafnium (Hf).

The revised Periodic Table, then, listed the heaviest elements as a second “rare earth” series, and these heaviest elements—for which the name “actinide elements” was then suggested—were paired off with elements in the already-known lanthanide rare earth series.

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	95	96	97	98	99	100	101	102	103

The early members of the actinide series, up to about plutonium, have chemical properties related to those of the lanthanide elements although they are not quite the same; the actinides lose electrons more easily in chemical reactions to form higher “oxidation states.”

This new concept meant that elements 95 and 96 should have some properties in common with actinium and some in common with their rare earth “sisters,” europium and gadolinium.

When experiments were designed according to this new concept, elements 95 and 96 were soon discovered—that is, they were chemically identified.

Americium, No. 95, was named after the Americas as its rare earth “sister,” europium, was named after Europe. Curium, No. 96, was named for Marie and Pierre Curie since the corresponding lanthanide element, gadolinium, had been named after the Finnish rare earth chemist Johan Gadolin.

Using the same concept to predict the chemical properties of elements 97 and 98, Thompson, Ghiorso, Kenneth Street,

Jr., and Seaborg discovered these two new elements in Berkeley in 1949 and 1950.

They also were found to be like their rare earth counterparts.

Berkelium, element 97, was named after the city of Berkeley just as terbium previously had been named after the city of Ytterby, Sweden, where the rare earths were first found.

Californium, element 98, is similar to dysprosium, but it was simply named after the university and the state where the work was done. The discoverers of californium explained their choice in naming the element as follows:

“The best we can do is to point out, in recognition of the fact that dysprosium is named on the basis of a Greek word meaning ‘difficult to get at,’ that the searchers for another element a century ago found it difficult to get to California.”

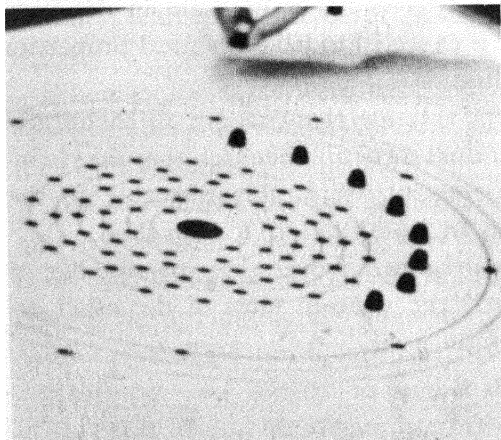
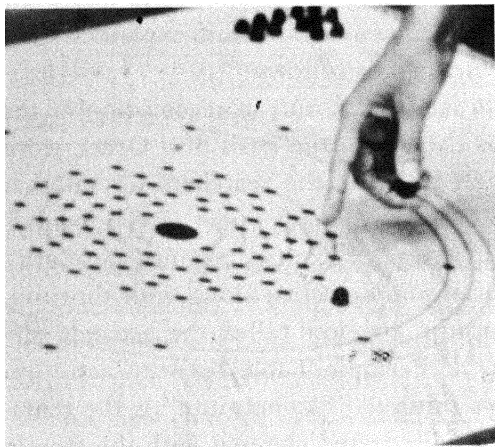
All these new discoveries, of course, verified Seaborg's theory, and the transuranium elements, along with thorium, protactinium and uranium, are now called the actinide elements. They all fit in the Periodic Table between actinium and the element “eka-hafnium.” “Eka-hafnium” is the tentative name given to the undiscovered element with the atomic number 104 which lies directly below hafnium in the Periodic Table and which is expected to have chemical properties similar to those of hafnium.

Now the fact that the actinide elements are all chemically similar to actinium reminds us of another fact, that the chemical properties of an element are largely traceable to the number of electrons in an atom's *outer* electron shell.

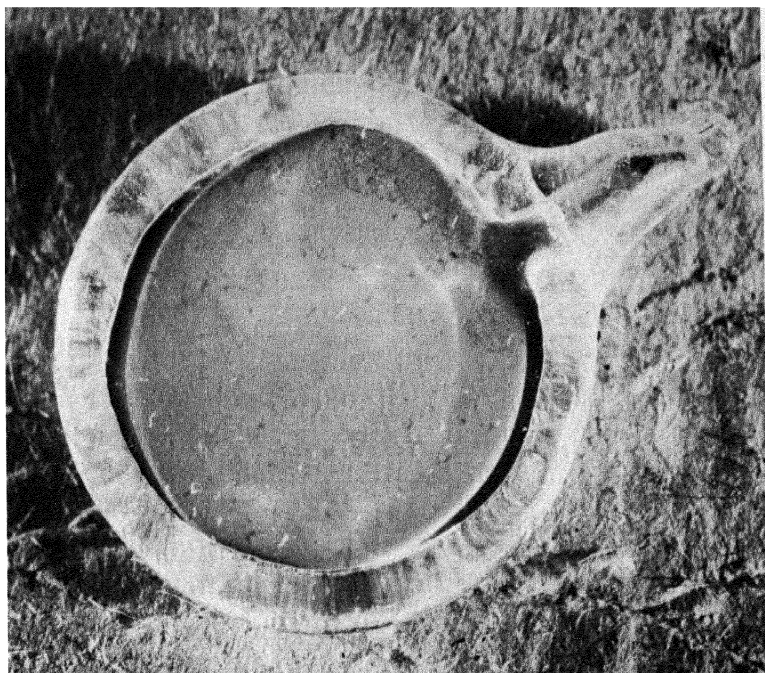
So it is no great surprise to discover that the number of outer electrons is nearly the same in most of the other actinides as it is in actinium itself.

Every element, of course, has a different and unique number of electrons. But most of the additional electrons required by successively heavier actinide elements are added in an inner ring of the atom, a ring of electrons known as the 5f orbit.

In a schematic model of an actinium atom we find a total of 89 electrons surrounding the nucleus and distributed among a number of different rings or orbits.



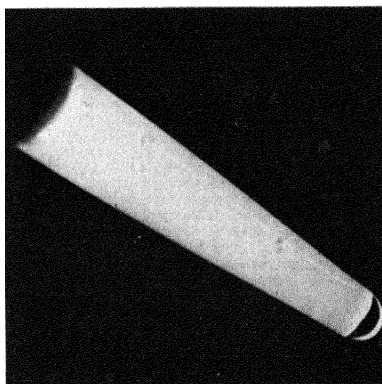
The additional electrons required by the next fourteen heavier elements are to be found in the inner, 5f orbit in most cases. By the time we come to the seventh element beyond actinium, which is element 96, 7 electrons have been added to this orbit; when we come to the fourteenth element, the undiscovered element 103, the 5f orbit has its full complement of 14 electrons and the actinide series of elements is completed.



Here is an oxide of americium in a vial about the size of a dime. This is a lot of americium, for elements beyond plutonium are increasingly unstable and difficult to prepare, and we have to deal with smaller and smaller quantities.

The tiny amount of curium in solution in this test tube is so radioactive that it was photographed by its own light.

Still heavier elements are often baked on to platinum plates in order to



detect them by their radioactivity because the quantities involved are too slight to be visible. In fact, at the present writing, no element heavier than californium has been isolated in a quantity large enough to see with the naked eye. The first weighable amount of californium—about one billionth of an ounce—was isolated in Berkeley in July, 1958.

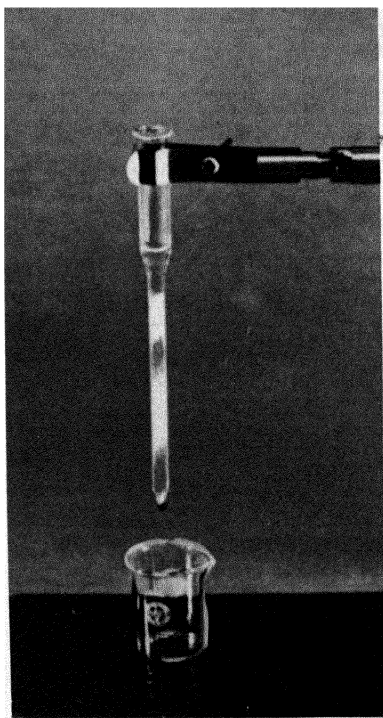
The discovery that the actinide elements are like the rare earth elements was most useful in identifying them.

The actinides, such as californium or berkelium, can be separated by a method which had been used successfully to separate the rare earth elements. It is known as the ion exchange absorption-elution method, a complicated name for something very simple. It can be demonstrated with other

elements such as cobalt and chromium. A mixture of the two is absorbed at the top of a column of an organic substance, a resin.

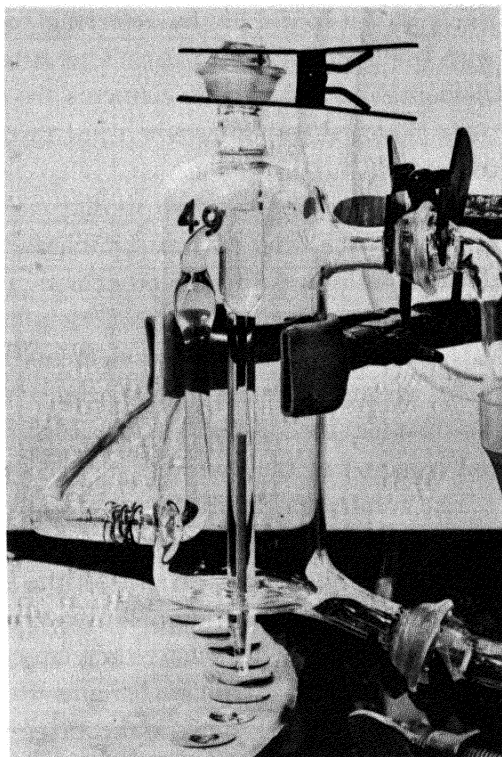
Then we pour in a solution that dissolves the mixture and washes it through the column.

One of the elements in the mixture washes through faster than the other and can be collected separately when it drips off at the bottom. In this cobalt-chromium separation you can see one dark band about a third of the way down the column and the other almost at the bottom.



The process works in the same way with the actinide elements although they are radioactive and often present in quantities too small to be seen.

The actual apparatus used to separate berkelium from californium, for example, is a dark column of resin surrounded by a glass jacket which serves merely to contain a vapor that warms the column so the



process will go on faster at a high temperature.

The solution containing the wanted element comes off drop by drop at the bottom and is collected on little platinum plates lying along the rim of a turntable.

One element comes off in the first few drops. Then a second element comes off in the next few drops, and so on. If these drops are collected separately, one by one, then the elements are, of course, separated from one another.

This is the way that all of the actinide elements heavier than curium were chemically identified in their discovery experiments.

We can even predict which drops a new actinide element is

likely to be found in by referring to similar experiments with the lanthanide elements. This is because the lanthanide elements and the actinide elements are remarkably similar in their chemical responses when put through an ion exchange column.

Suppose we perform two similar experiments, one with a mixture of the nine heaviest lanthanide elements (europium through lutetium) and the other with a mixture of the heaviest actinide elements (from americium on up).

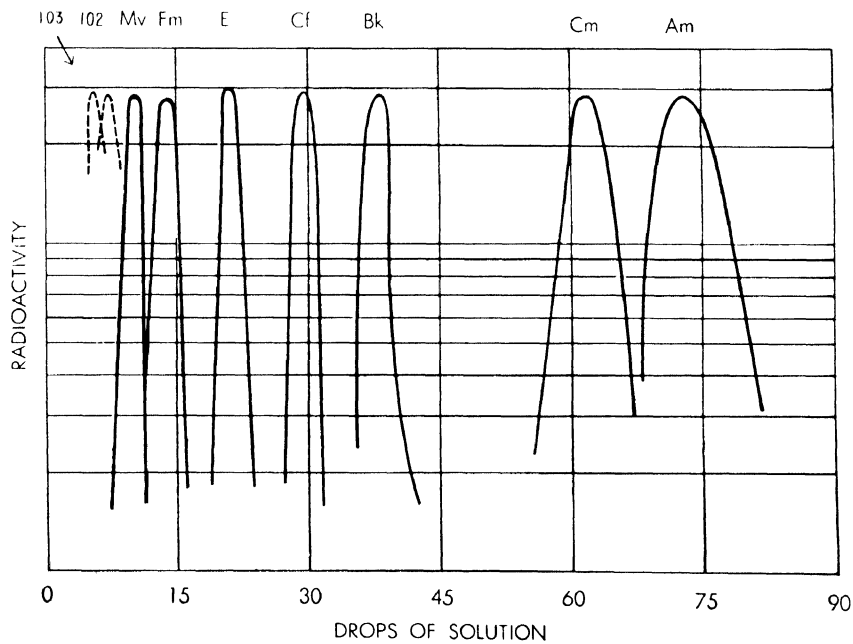
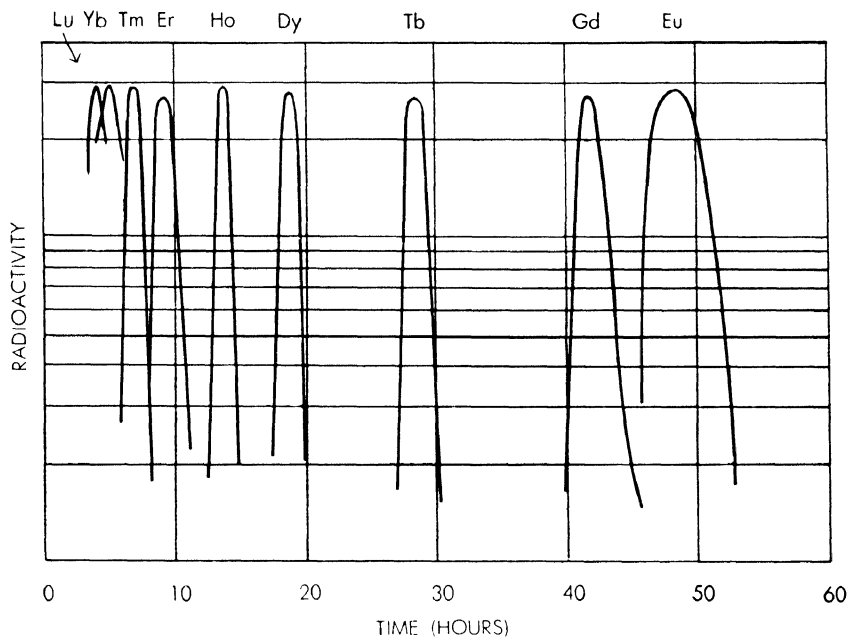
If we put the lanthanide elements and the actinide elements through separate ion exchange columns, we can record the time it takes for each element to run through the column and come off at the bottom.

The results appear in the graphs on the opposite page. In the upper graph, the first element to traverse the column and reach the bottom is the heaviest of the lanthanides, lutetium. It is followed by ytterbium, thulium, erbium, and so forth.

In the lower graph, where the elapsed time is recorded in terms of individual drops, elements 103 and 102 would appear first if they existed in the original mixture; they are shown in dotted lines to indicate their predicted positions. In actual experiments, which have not been conducted with element 102, mendelevium appears first, followed by fermium, einsteinium, and so on.

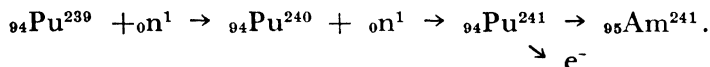
In every case, the time it takes for a particular lanthanide element to appear at the bottom of the column is echoed by the time it takes for the corresponding actinide element to traverse its column. For example, there is the same wide gap between the appearance of terbium and the appearance of gadolinium as there is between that of their "sister" elements, berkelium and curium.

This completes the story of the first six transuranium



elements except for noting the rather straightforward nuclear reactions involved in the production of americium, curium, berkelium, and californium.

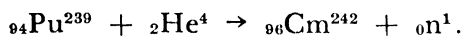
Americium was made by absorbing neutrons in plutonium.



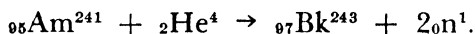
The plutonium absorbs a neutron, becoming heavier, and then absorbs another neutron. Then the plutonium nucleus emits an electron, changing one of its neutrons into a proton and, of course, resulting in the new element, americium.

The next three heavier elements were all made by bombardment with helium nuclei. In each case the helium nucleus, with its 2 protons and 2 neutrons, joins with a heavy nucleus. This increases the positive charge on the heavy nucleus by two, and a neutron or two boils off in the process.

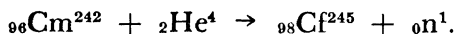
Curium was first made by bombarding plutonium-239.



The production of berkelium involved the bombardment of americium.



And curium was transmuted into californium in a similar way.



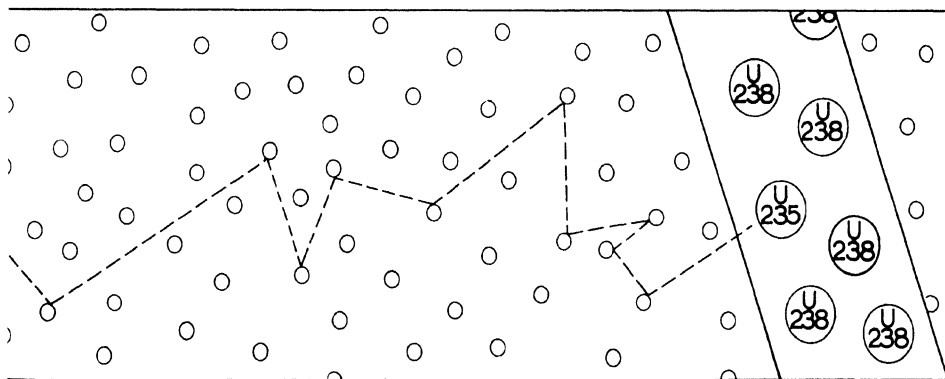
The discovery of nuclear fission has overshadowed all other recent developments in nuclear chemistry and physics.

The important fissionable isotopes of plutonium and uranium—Pu-239, U-235, and U-238—can be used in controlled chain reactions to produce atomic energy.

The heart of an atomic reactor or atomic pile can be a group of uranium fuel elements. These are made from natural uranium, which contains two kinds of atoms, uranium-235 which will undergo fission and uranium-238 which will not.

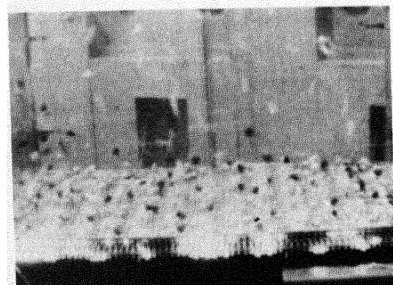
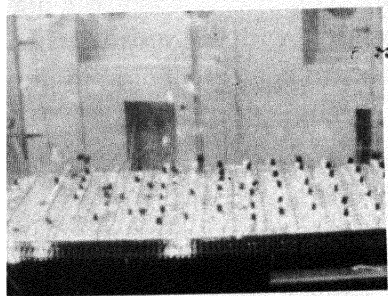
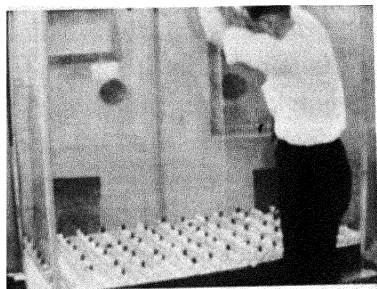
When a U-235 nucleus is struck by a neutron, it undergoes fission and breaks apart; each such nucleus thus produces two radioactive fission products. The fission also produces two or three more neutrons which fly out and keep the chain reaction going by colliding with other U-235 nuclei.

However, the neutrons must be slowed down because they are more effective that way. This slowing down is accomplished by a “moderator” such as graphite. A schematic diagram would show a neutron colliding with successive carbon atoms and finally entering the nucleus of an atom of uranium-235.



This causes the U-235 nucleus to split and release two or three new neutrons in the process. These neutrons follow a path similar to that of the first neutron and usually are captured by U-235 nuclei, causing fission and the release of still more neutrons.

The resulting chain reaction can be illustrated by a mechanical demonstration which is a lot of fun to watch. A large



number of mousetraps are set and two or three corks are placed on the spring "trap" of each one. When a single cork—representing a slow neutron—is dropped from above, it will spring one mousetrap, releasing two or three more corks which in turn set off more mousetraps. This chain reaction progresses with increasing speed until almost all of the mousetraps have gone off.

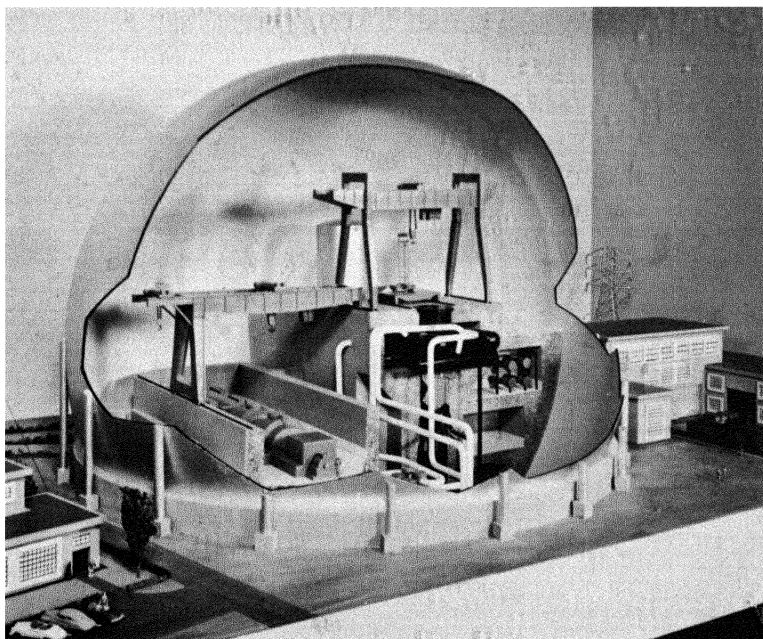
In this particular demonstration each mousetrap is wired to a small light which goes on when the mousetrap is sprung.

The mousetrap is not a very good analogy with a U-235 nucleus because it stays in one piece whereas a fissioning U-235 nucleus

breaks into two parts. These two major fragments of the split nucleus travel at high speed through the rest of the uranium in the fuel element. As they move through the fuel element, they are slowed down and give out energy in the form of heat just as slowing down an automobile causes the brakes to heat up. And this is the way we get useful heat from a chain reaction in an atomic pile.

The U-235 chain reaction is harnessed for commercial power as illustrated in this model of a nuclear-electric power plant near Chicago. Enclosed in concrete beneath the upper crane is the uranium reactor with its vertical rods which are the fuel elements.

The moderator in this particular plant is ordinary water. Heat from the nuclear fission of uranium causes the water to boil. The resulting steam is carried to conventional steam



turbines which turn the generators and produce electricity.

Essentially, the energy source is a nuclear reactor, and the rest of the plant consists of conventional machinery.

An atomic pile also creates an important by-product: more nuclear fuel.

Some of the neutrons released in the controlled chain reaction strike the nuclei of non-fissionable U-238 atoms. In this case, the U-238 captures the neutron and becomes a heavier uranium isotope, U-239, which eventually decays to produce plutonium-239.

This plutonium-239 can be chemically separated and used as a nuclear fuel in another reactor.

Similarly, the nuclear fuel uranium-233 can be made when ordinary thorium. (Th-232) captures a neutron to produce thorium-233; the thorium-233 decays to produce uranium-233.

DISCOVERY IN A MUSHROOM CLOUD

The discovery of elements 99 and 100 began in the South Pacific. In November 1952, there was a great explosion, the first full-scale thermo-nuclear explosion initiated as a chain reaction.

The explosion bombed a mile-wide crater in the island site, and a radioactive cloud rose ten miles and reached a diameter



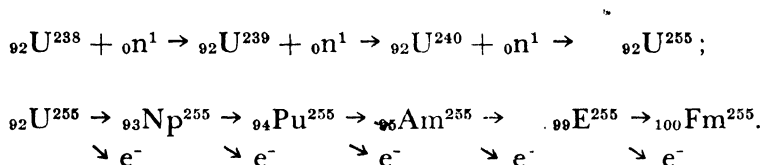
of about a hundred miles.

Radio-controlled drone planes flew through this cloud and collected samples of it for laboratory analysis so that what had happened during the explosion could be determined.

What actually was found in the laboratory proved to be a very unusual scientific accident. The uranium which had been included in the thermo-nuclear device was found to have captured very many neutrons, as many as 17. The normal weight of uranium—238—was thus increased to as much as 255.

This extremely heavy uranium isotope underwent radioactive decay, forming successively heavier transuranium elements, including isotopes of elements 99 and 100.

This series of reactions can be written in abbreviated form as follows:



The two new elements were named in honor of Albert Einstein and Enrico Fermi, einsteinium for element 99 and fermium for element 100. They were first found by isolating them from the enormous amount of debris produced by the explosion.

At first it was necessary to process by chemical methods the material obtained from filter papers which had been flown through the radioactive cloud by the drone planes.

But to get larger amounts of the new elements, the scientists on the project had to process tons of coral collected at the site of the explosion.

The discovery of elements 99 and 100, like so much of the research in the 1950's, was the result of teamwork. The scientists involved included Ghiorso, Thompson, G. H. Higgins, and Seaborg from the University of California Radiation Laboratory, nine researchers at the Argonne National Laboratory—M. H. Studier, P. R. Fields, S. M. Fried, H. Diamond, J. F. Mech, G. L. Pyle, J. R. Huizenga, A. Hirsch, and W. M. Manning—and three others at the Los Alamos Scientific Laboratory—C. I. Browne, H. L. Smith, and R. W. Spence.

There are other methods of preparing these same isotopes of einsteinium and fermium, and one of these involves the use of the chain-reacting atomic pile which produces a high intensity of bombarding neutrons.

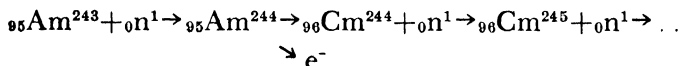
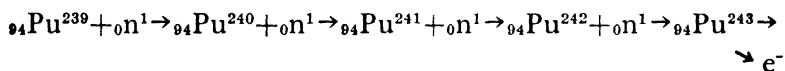
The specimens to be bombarded are placed in the center of a pile such as the "Materials Testing Reactor" at Arco, Idaho, a model of which is shown here.

The material to be irradiated is in the form of "napkin rings," each consisting of plutonium metal mixed with another metal (a combination of metals known as an alloy) and covered by an aluminum jacket.

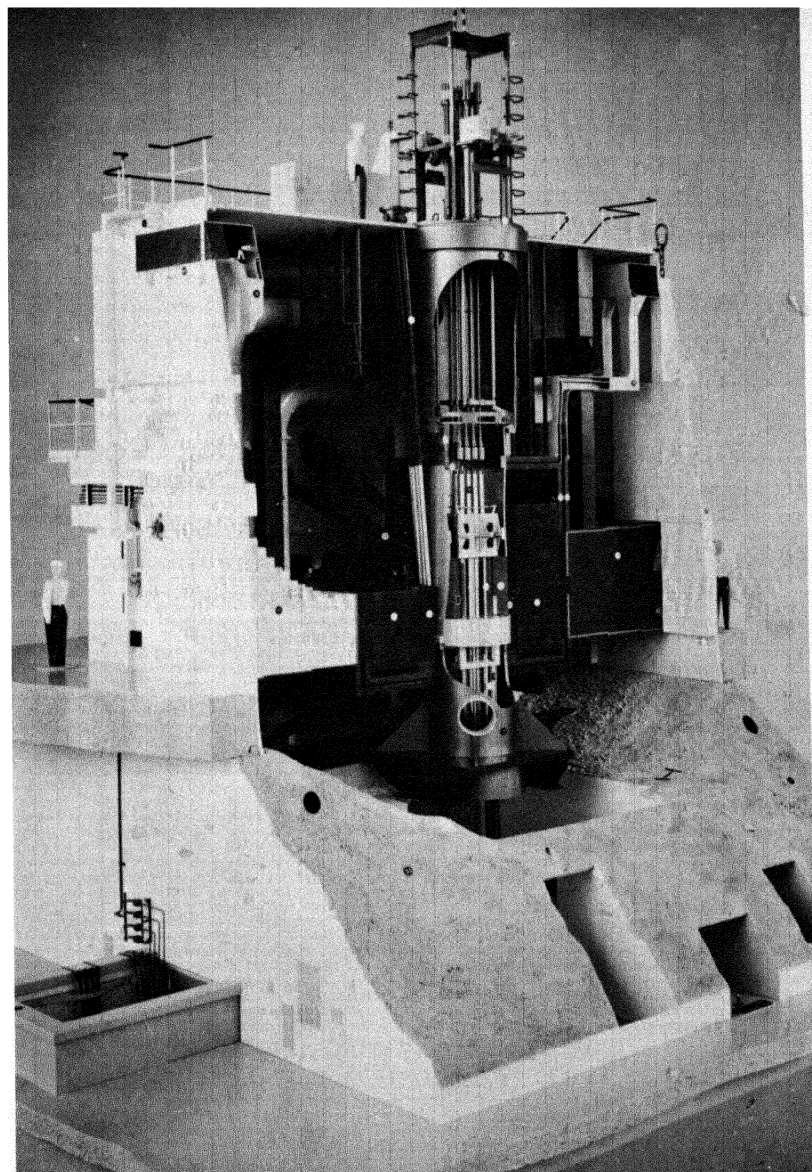
The plutonium alloy is in the shape of an ordinary table napkin ring in order to make it possible to remove the heat developed in the nuclear fission reaction by a water-cooling process.

In the reactor, some of the plutonium captures neutrons and decays into americium, which captures a neutron and decays into curium, and so on:

These reactions can be written:



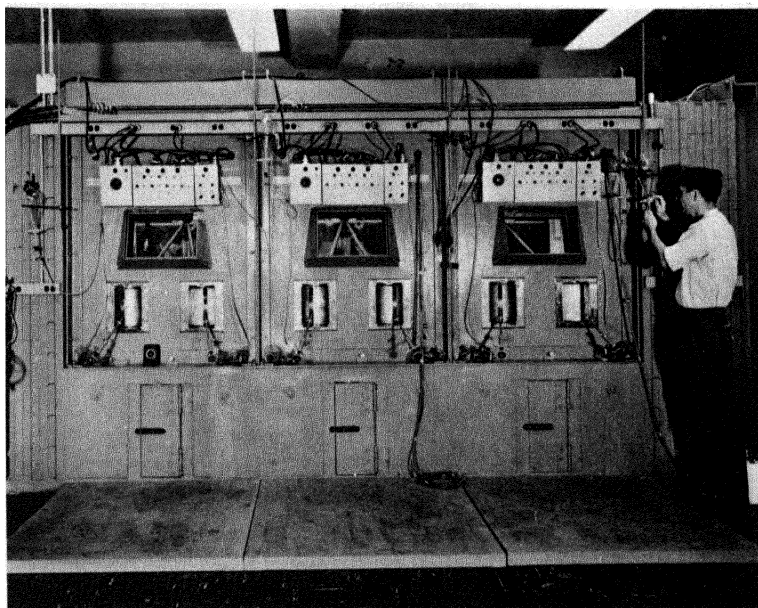
The basic difference between the two methods of producing

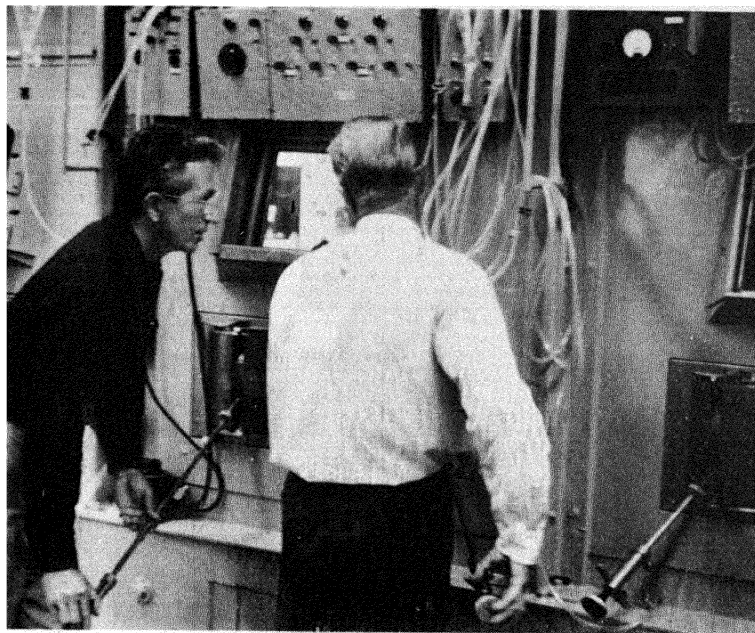


these heavy elements is one of time. In the case of uranium in a thermo-nuclear explosion, the reactions proceed in only a few millionths of a second. In the case of the plutonium napkin rings, the "fattening up" of the plutonium nucleus takes place over a period as long as two years or more, so you might say the isotopes are made from plutonium the long, slow, natural way.

In addition to the desired transuranium elements, a tremendous number of fission products is produced by the fission reaction on plutonium itself. These radioactive fission products are isotopes of a large number of elements of considerably lower atomic weight than uranium. Their presence means that there is a great and hazardous amount of radioactivity to be dealt with.

This calls for construction of a special "hot" lab such as this "cave room" at the Berkeley Radiation Laboratory.



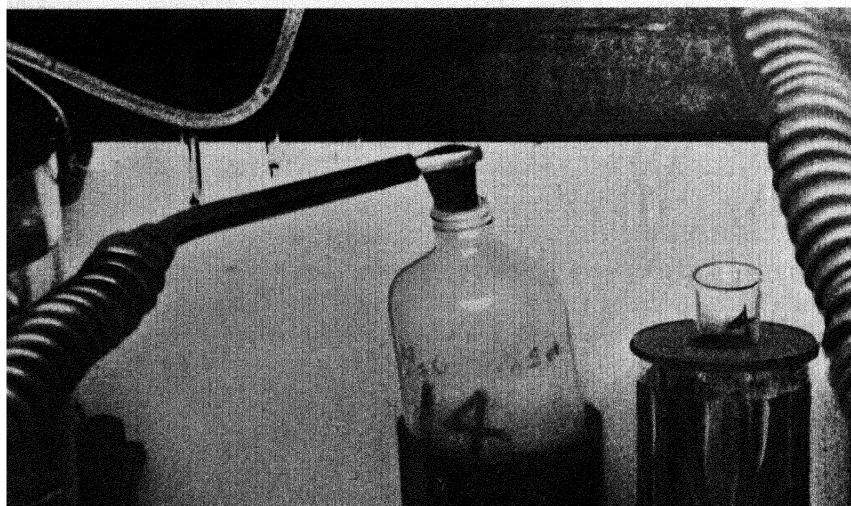
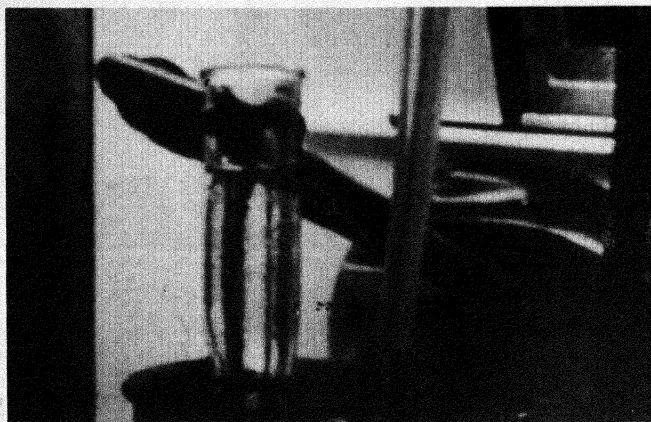


Here the necessary chemistry is performed behind heavy shielding by remote-control tongs operated from the outside.

The "cave" contains three separate metal boxes in which the dangerous material is handled by the two-fingered tongs. The boxes are enclosed in a six-inch lead casing, and each has a viewing window made of high-density "lead-glass" nine inches thick.

The boxes are airtight, have their own air-conditioning system, and maintain a slight vacuum so that material in the box will not escape even if there should be a slight leak.

The tongs, which are a mechanical extension of an experienced operator's arm and hand, are extremely versatile. They can handle test tubes, put corks on bottles, pour solutions, operate a spotlight or a pair of tweezers, and even mop up a spilled solution with tissues available inside the box.



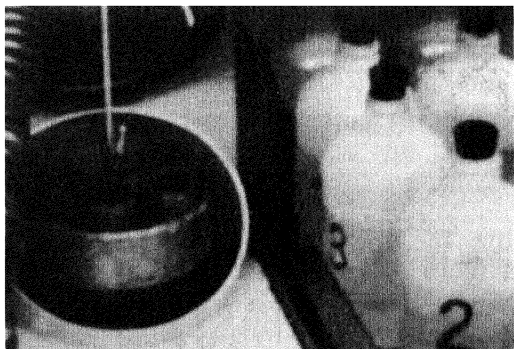
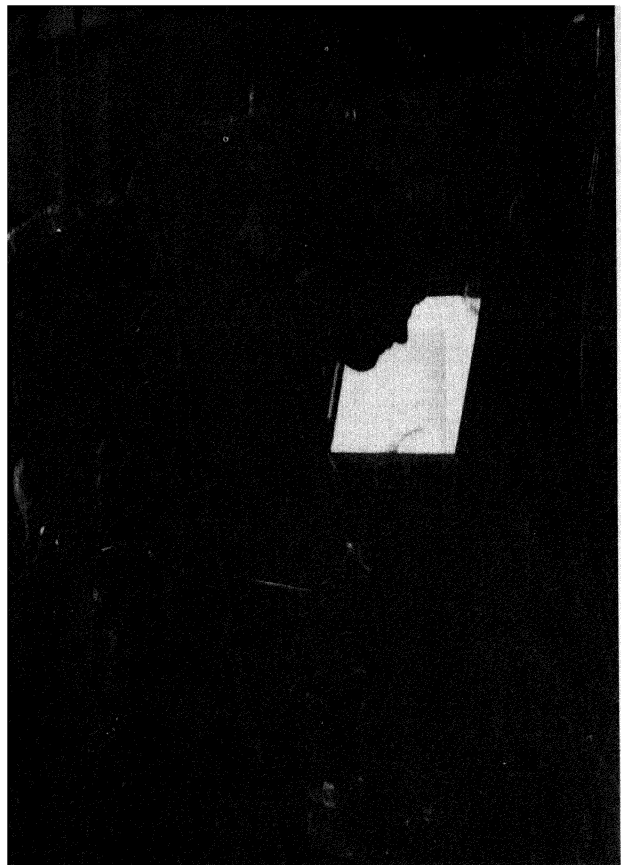
The purpose of all this equipment is to isolate tiny quantities of the rare synthetic elements like americium, berkelium, and einsteinium. They are present in the irradiated plutonium napkin rings along with the great quantity of highly radioactive fission products.

A "mass production" run of the heavy synthetic elements took place in Berkeley in October 1956. The chemical steps involved had been worked out in detail, and chemists had practiced with the tongs for months, using dummy napkin rings.

Ten "hot" napkin rings had been "cooked" with neutrons in the Arco reactor for more than two years. They were flown to Berkeley in a heavy lead container which was wheeled into the heavily shielded "cave" and inched into place directly beneath a port in the bottom of one of the metal chemical boxes.

Then a great lead door at the back of the "cave" was closed and





the entire cave area sealed off.

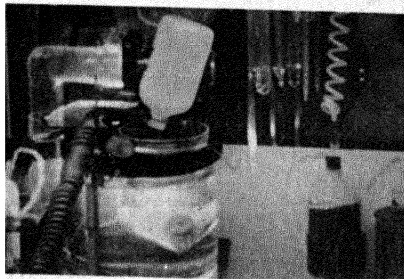
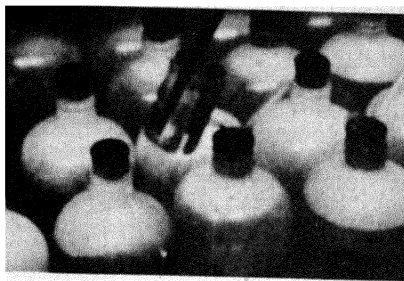
The first remote-controlled operation is that of opening the port in the bottom of the box. The "hot" napkin rings, worth thousands of times their weight in gold, are fished up out of the lead container beneath and drawn into the box.

The first ring is then picked up by the tongs and placed in an alkaline solution which will disintegrate and partially dissolve it. The resulting dark, murky liquid is siphoned into a polyethylene bottle and stored while the other nine napkin rings are dissolved,

The dark liquid is poured into glass cones and centrifuged at 1500 revolutions per minute. This throws the heavier elements, which are in the form of hydroxides, to the bottom of the cones.

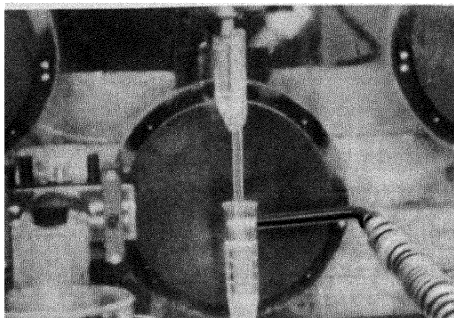
The chemical task here is to separate the heavy trans-uranium elements from the aluminum and from all the fission products formed during the bombardment of the plutonium.

These fission products include isotopes of nearly half of the elements existing on earth, all of them made from plutonium by nuclear transformations.



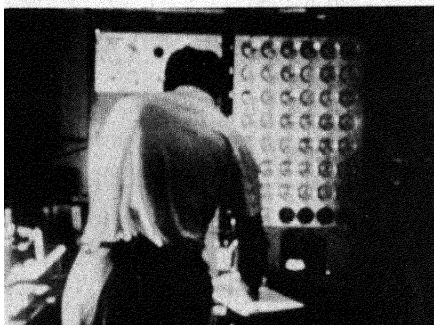
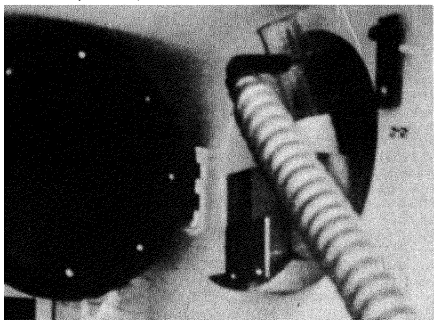
After centrifuging, the lighter liquid at the top of each cone is siphoned off.

This process is repeated a number of times to further purify the residue of



heavy elements and reduce the level of radiation, which is due mostly to the unwanted fission products.

The precipitated material containing the heavy elements is then dissolved and transferred into the neighboring chemical box by way of a sort of airlock between the two boxes.



The second box is used for column chemistry, which brings about the final separation of the various heavier elements as described on pages 148-151. The solution from the first box is forced through a column of organic material. The heavier elements remain in the column temporarily while many of the fission products which are still left pass through the column at a faster rate.

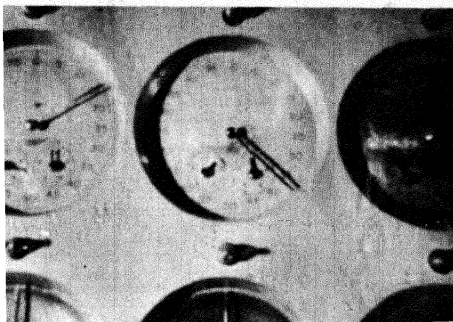
A "pulse analyzer" is

used to test a sample of the finished product after the sample has been dried onto a platinum "counting plate."

The pulse analyzer is an instrument which sorts out the radiations from different radioactive elements according to their respective energies. This, of course, makes it possible positively to identify the elements present in the sample.

The pulse analyzer is connected to dials or other recording instruments on a large panel in such a way that each separate dial will record the nuclear disintegrations of atoms of one particular element.

This dial is registering disintegrations of single atoms of einsteinium, element 99, which was isolated in the October 1956 napkin-ring operation.



Einsteinium was produced in a similar manner during a smaller-scale "run" the previous year, and it was subsequently bombarded with helium ions in the 60-inch cyclotron to produce the first atoms of element 101 ever made.

The discovery of mendelevium, element 101, demanded some fancy footwork on the part of the chemists involved because mendelevium has a half-life of only about half an hour. It was distinctly a matter of chemistry-on-the-run.

Mendelevium was produced and identified at Berkeley early in 1955 by Albert Ghiorso, Bernard G. Harvey, Gregory R. Choppin, Stanley G. Thompson, and Glenn T. Seaborg.

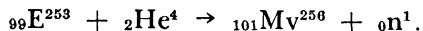
Its discovery—and the techniques involved in making something totally new under the sun—is described here by Ghiorso and Harvey (pictured below), Choppin and Thompson.



THE RACE FOR 17 ATOMS

The new element mendelevium was made by bombarding einsteinium, element 99, with helium nuclei.

The nuclear reaction is a simple one:

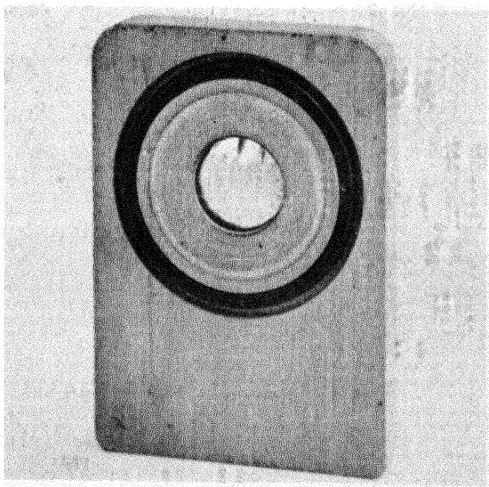


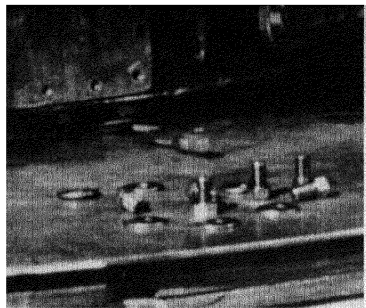
We accomplish this in the cyclotron where a beam of helium nuclei strikes a small target.

The target is a very thin circle of gold foil onto the back side of which has been electroplated an invisibly thin layer of einsteinium—not more than a few billion atoms.

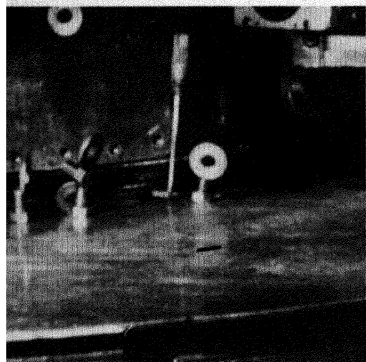
If any atoms of einsteinium are transmuted into mendelevium by the bombardment, they will be kicked out of the target foil by the impact of the helium nuclei. A second gold foil is placed directly behind the target to catch atoms of the new element as they are ejected from the target foil by the bombardment.

Both the target and the “catcher” are fitted into a holder which will be secured inside the cyclotron directly in the path of the bombarding particles.



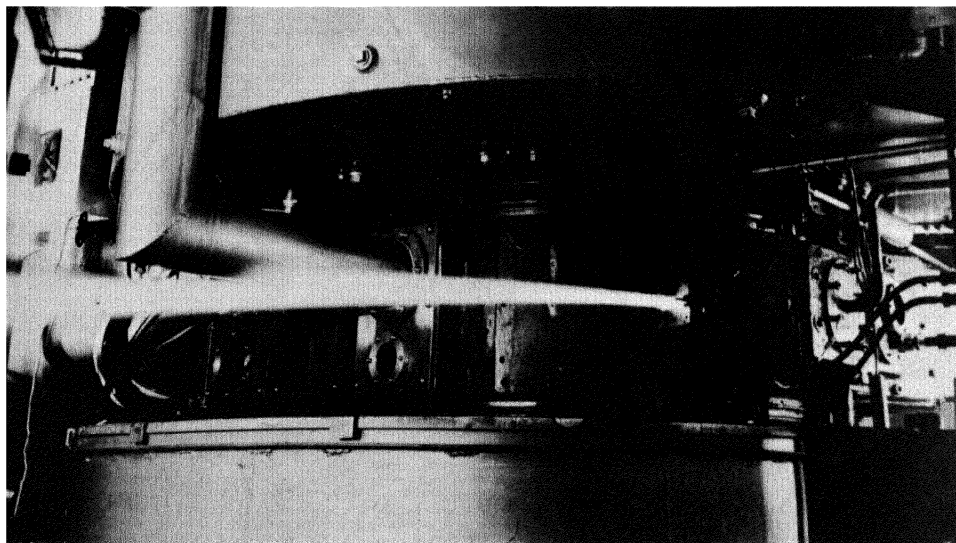


The fast-moving helium nuclei were supplied by the old, 60-inch cyclotron on the Berkeley campus of the University of California.



The cyclotron's magnet, which keeps the accelerated helium nuclei whirling in their spiral orbit, is strong enough to make a screwdriver stand on end or cause heavy iron washers to hang in the air.

We can see the powerful stream of helium nuclei as a thin blue beam of light if we allow it to escape through the target area out into the air. It can be photographed through the five-foot tank of water which acts as a viewing



window into the room housing the cyclotron. This is the beam which will strike the target and, with luck, add the 2 protons of the helium to the 99 protons of the einsteinium in the target to make mendelevium.

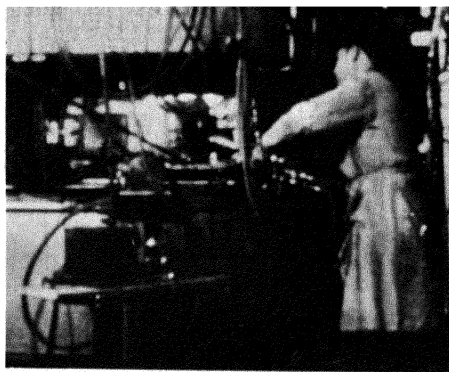
In the actual experiment—which we re-enacted to get these photographs—the entire cyclotron room was sealed off while the target was bombarded. Bernie and Al waited outside the water-filled door, a great tank of water on rollers.

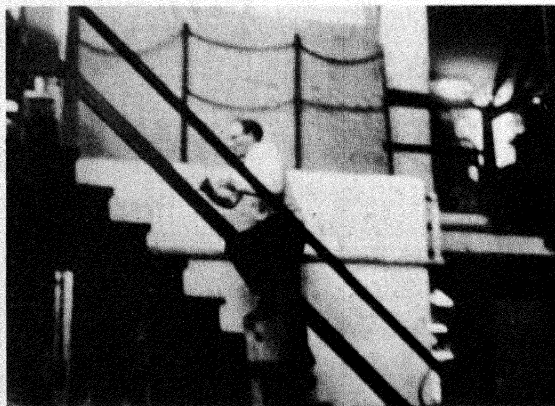
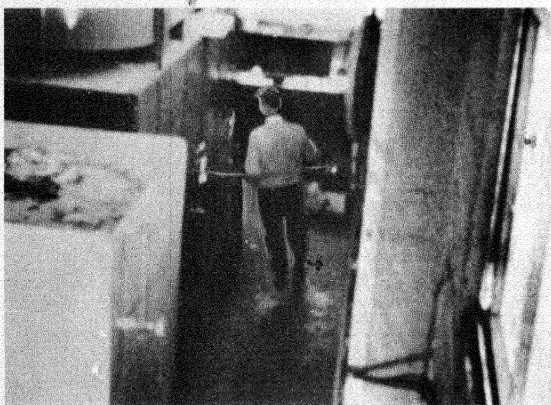
We were really waiting for the starting gun in a very unusual obstacle race. This first experiment was expected to produce just one atom of element 101, or perhaps two atoms. And this one or two atoms had to be separated from billions of atoms of einsteinium and identified in less than half an hour.

As soon as we got the all-clear signal, Bernie and Al pulled open the water-filled door and ran inside.

Al quickly took the “holder” from the machine. Bernie pulled off the second gold foil with a pair of tweezers and shoved it into a test tube.

Then he raced out through the narrow corridors and up a





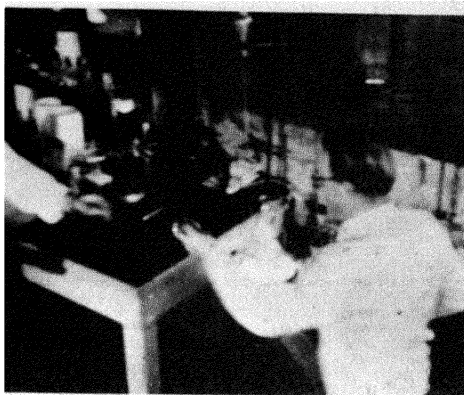
flight of steps to a temporary laboratory. In this makeshift lab, Bernie gave the foil to Gregory Choppin, who heated it in a solution to dissolve the gold.

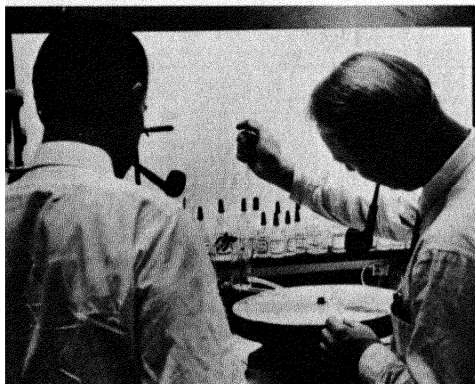
The result was a liquid containing the gold, some other miscellaneous elements, and perhaps a few atoms of the mendelevium we were after.

The rest of the necessary chemical steps had to be done a mile away, up on the "hill" at the Radiation Laboratory.

For the mad dash up the hill, Al was ready at the wheel of a get-away car just outside the cyclotron building.

We had—we hoped—a few atoms of element 101, and our job was to isolate and identify them before





they all disintegrated.

Mendelevium is so short-lived that half of any sample decays in about half an hour, becoming an isotope of fermium which in turn decays by spontaneous fission.

The precious drops were driven up to the nuclear chemistry building on the hill. Greg and Bernie raced into the lab where Stanley Thompson was waiting with the apparatus with which to separate the element 101 from einsteinium, gold, and anything else that might be present.

First the liquid was forced through an ion exchange column to get rid of the gold. The gold is held back in the column while a solution

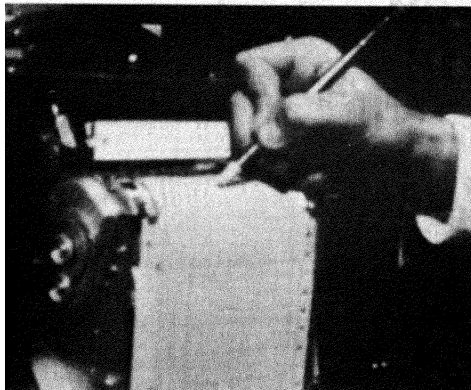
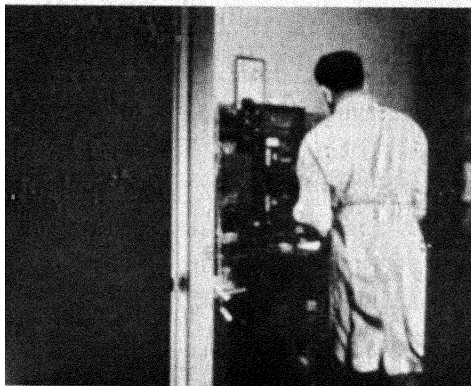
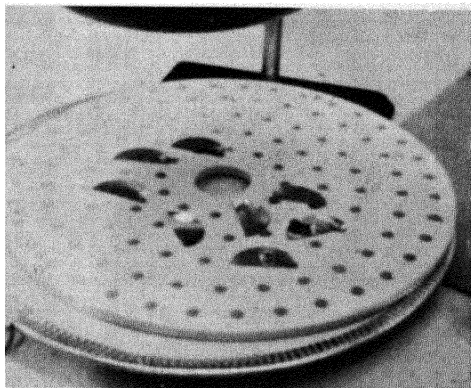
containing the mendelevium drips off at the bottom.

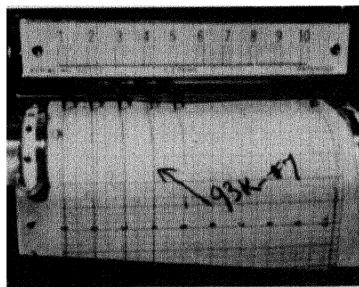
These drops were dried and redissolved. Then Stan forced them through a second column to isolate element 101 from any other elements still present.

The drops from the bottom of the column were caught, one by one, on little platinum plates which were then put under a heat lamp and dried.

The plates were taken into a counting room where Al put them in special counters, one to each counter.

If there were any mendelevium present in the particular drop being tested, it would show up as it decayed. Whenever an atom of the new element disinte-





grates, the subsequent exploding fission fragments cause a large burst of ionization. This pulse of current makes a pen on the recording chart jump.

It is typical of these elusive heavy elements that we cannot positively identify an atom until the moment that it ceases to be that element and disintegrates into something else. It's rather like the man who only counts his money as he spends it.

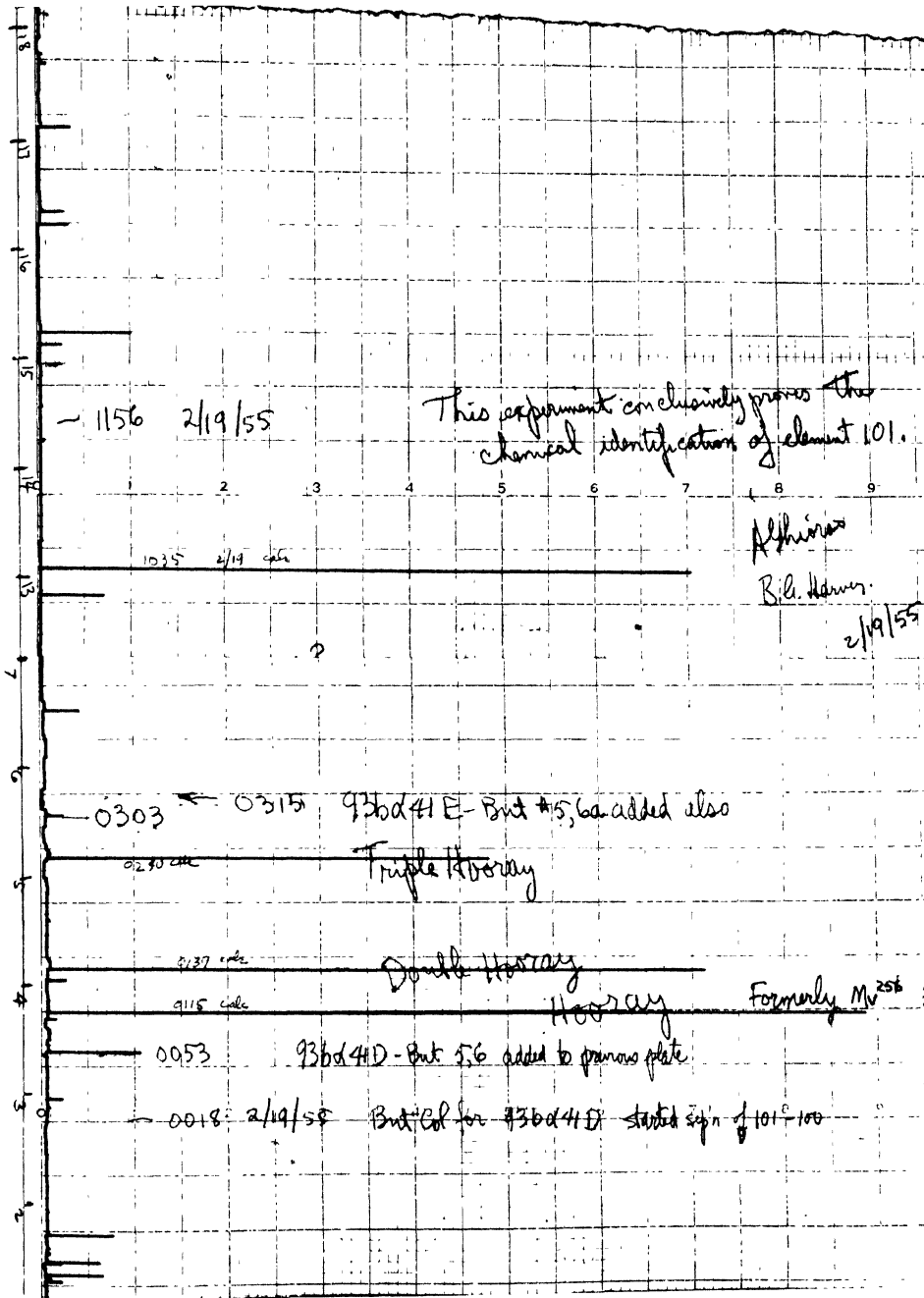
In the first experiment, we waited more than an hour before the pen shot to mid-scale and dropped back, marking a line that meant the disintegration of the first known atom of mendelevium.

Since this was quite an event at the Rad Lab, we connected a fire bell in the hallway to the counters so that the alarm would go off every time an atom of element 101 disintegrated. This was a most effective way of signaling the occurrence of a nuclear event, but quieter means of communication were soon substituted, following a suggestion put forward by the fire department.

We found only about one atom of mendelevium in each of our first experiments. We repeated the experiment perhaps a dozen times, and our grand total was seventeen atoms of the new element.

The recording chart from the first complete experiment shows the disintegrations of the first atoms of mendelevium ever made and completely identified.

ALBERT GHIORSO
BERNARD G. HARVEY
GREGORY R. CHOPPIN
STANLEY G. THOMPSON



- 1156 2/19/55

This experiment conclusively proves the
chemical identification of element 101.

Alphina

Bill. Haver.

2/19/55

1035 2/19/55

0303

0315

93bd41E - But #5, 6 added also

0.230 calc

Triple Hierarchy

0137 calc

Double Hierarchy

0115 calc

Hierarchy

Formerly Mv 256

0053

93bd41D - But 5, 6 added to previous plate

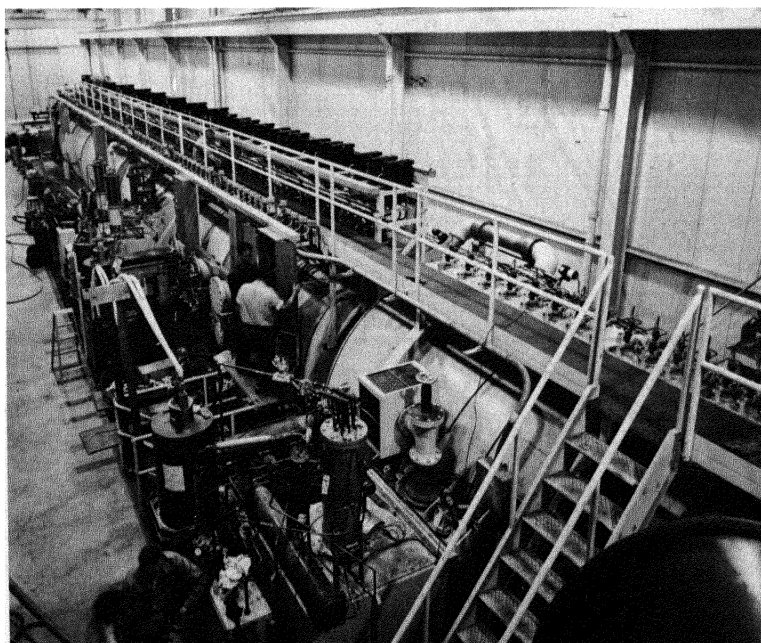
0018 2/19/55

But Cal for 93bd41E started sep'n of 101 & 100

The discovery of element 102 was reported at the Nobel Institute for Physics in Stockholm in 1957, but experiments to date attempting to confirm its discovery have not been successful. The element was produced and positively identified, however, at the U.C. Radiation Laboratory in Berkeley in April 1958. It was made by bombarding curium (containing 96 protons) with carbon ions (containing 6 protons).

The resulting isotope of element 102 had an atomic weight of 254 and a half-life of about three seconds.

The bombardment was effected in the new "Hilac," or heavy ion linear accelerator, in Berkeley. A linear accelerator, as the name implies, accelerates particles along a straightaway rather than around a circular course as in the cyclotron or bevatron.



THE UNDISCOVERED ELEMENTS

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	A
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac-103	(104)	(105)	(106)	(107)	(108)	(109)	(110)	(111)	(112)	(113)	(114)	(115)	(116)	(117)	(118)
Lanthanides		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
Actinides		Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	E	Fm	Mv	102	(103)	

Nuclear chemists obviously will continue trying to produce still more synthetic elements with atomic numbers greater than 102.

Element 103 should complete the actinide series of heavy rare earth elements. It should be analogous in its chemical properties to the rare earth lutetium.

Then we would jump to element 104, which should come after the actinide group and be a member of the family of elements including hafnium, zirconium, and titanium. Similarly, element 105 should be chemically related to tantalum, niobium, and vanadium. Heavier elements, through 118, should then fill in the rest of the same horizontal row on the Periodic Table.

Of course, it won't be possible to create that many elements, for each successively heavier element is more unstable.

But it should be possible to synthesize elements up to 104 or 105 and these elements probably will have half-lives long enough for them to be chemically identified.

A few more elements might possibly be identified by more complicated methods.

How will these heavier synthetic elements be produced?

It seems most likely that heavy ions will be employed as the bombarding particles, as was done in the case of element 102.

Instead of using the familiar ions, or nuclei, of helium, with their 2 protons, we might use nitrogen ions. The 7 protons of nitrogen, added to curium with its 96 protons, for example, should give us element 103.

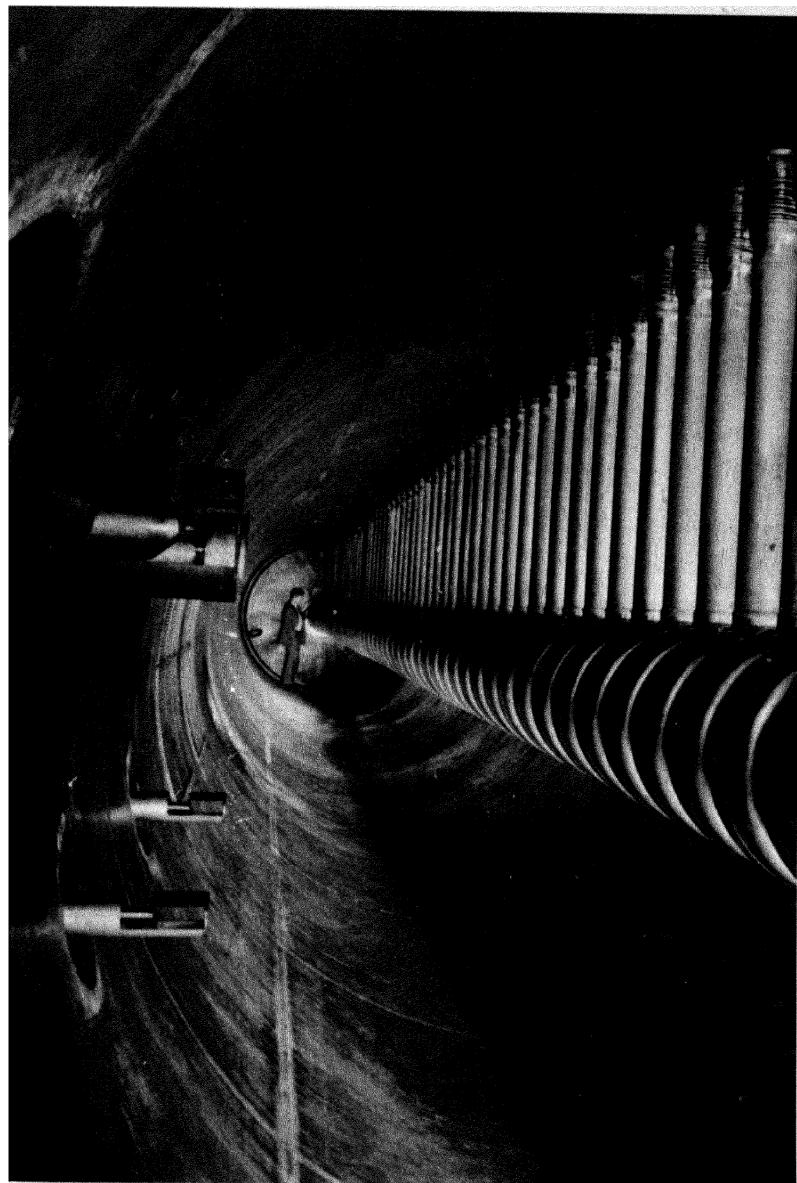
Or suppose we use neon. If we bombard plutonium with neon, then neon's 10 protons plus plutonium's 94 should produce element 104.

In a similar manner, curium's 96 protons plus carbon's 6 protons gave us our first atoms of element 102.

Accelerators designed specifically for the purpose of producing such heavy ions are being built at several laboratories. Berkeley's "Hilac," which can bombard targets with particles as heavy as neon ions, or even heavier ones, is just one example.

Such machines naturally call for vacuum tanks of considerable size, as this interior view of the "Hilac" illustrates. The vacuum tank encloses the course along which the bombarding particles are accelerated.

With such accelerators available, scientists hope to discover several more new elements and broaden their knowledge of the nature of atoms and atomic nuclei.



THE PLANET EARTH

For all the wonders we have learned to perform with the chemical elements, none of them is very impressive when compared with that remarkable arrangement of elements which we know as the planet earth.

The abundance of some elements and the scarcity of others, the very particular ways in which elements are mixed and compounded, the fact that the elements in our atmosphere cannot escape into space—such natural “accidents” have made our planet one of those rare fragments of the universe capable of supporting life.

Everything in this photograph—the clouds over Mexico, the Pacific Ocean along the horizon, and the rock of southern Texas below—is an improbable peculiarity in the universe.

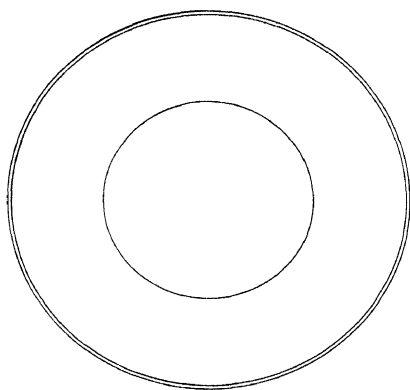
Oxygen, for example, accounts for but a fraction of one per cent of the substance of the universe, yet it comprises 89 per cent of the weight of clouds and water and 46 per cent of the solid crust of the earth.

The earth, of course, is entirely composed of the chemical elements, those fundamental building blocks of everything that exists. The millions of different chemical substances on earth are all made up of elements.

By examining the planet, we should be able to see how elements function in their natural context, how abundant or scarce various elements are on earth, how they are arranged, and how they sustain life.

We have learned a great deal about the interior of the earth by studying the earthquake waves which pass through it.





We've learned that the earth is divided essentially into three parts: a thin outer crust some twenty or more miles thick; an intermediate "mantle" of basaltic rock extending halfway to the center of the earth; and a central core.

We have to conjecture somewhat about the core. We can calculate its weight since we know the weight of

the whole earth as a result of its gravitational effects on other planetary bodies.

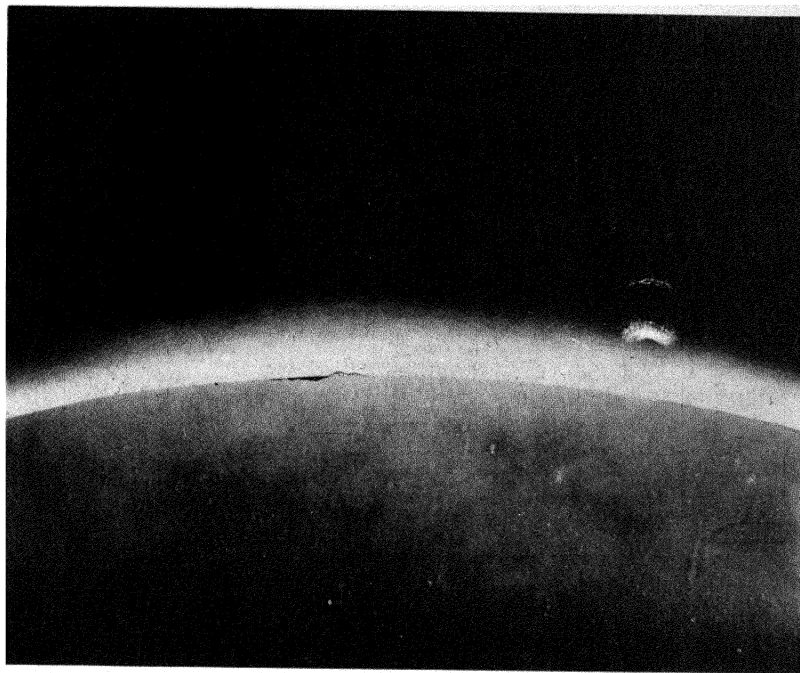
Earthquake waves tell us that the core is liquid, although this liquid core itself may contain a solid "inner core."

These known facts seem to fit well with the assumption that the core has the same composition as metallic meteorites, namely, iron and nickel.

The mantle seems to be composed of silicon and oxygen and a little iron. Its composition apparently is like that of the stony meteorites which bombard the earth from space.

Of far greater importance to us is the hard outer crust, or "lithosphere," and the oceans and atmosphere surrounding it. Man himself lives in this outer skin, bound to the surface. Although he has scaled Mt. Everest, he has managed to descend only some two miles down into the ocean, as indicated by the white dot in the drawing. On the left, the deepest mine is represented and on the right, the highest man has flown, as compared with the height of meteors and the auroras.

The average height of the first artificial earth satellites is still well within the frame of the drawing.



This is the environment of man, the part of the earth comparable with an apple's skin, a film embracing air and the surface of the land and sea. Sometimes it's even called the "biosphere" of the earth.

It is the unique arrangement of elements in the earth's crust which has led to life and which now sustains it.

Which elements are most common? Which elements are essential to life? How are they distributed?

Eight elements, led by oxygen and silicon, constitute more than 98.5 per cent of the earth's solid crust. Other familiar elements are scattered throughout the list, with radium and plutonium very near the bottom.

oxygen	46.60 %
silicon	27.72
aluminum	8.13
iron	5.00
calcium	3.63
sodium	2.83
potassium	2.59
magnesium	2.09
titanium	.44
phosphorus	.118
manganese	.100
fluorine	.06-.09
sulfur	.0520
carbon	.0320
chlorine	.0314
rubidium	.0310
strontium	.0300
barium	.0250
zirconium	.0220
chromium	.0200
vanadium	.0150
zinc	.0132
nickel	.0080
copper	.0070
tungsten	.0069
lithium	.0065
nitrogen	.00463
cerium	.00461
tin	.0040
yttrium	.0028
niobium	.0024
neodymium	.00239
cobalt	.0023

lanthanum	.00183 %
lead	.0016
gallium	.0015
thorium	.00115

less than 1/1,000% :

molybdenum
germanium
cesium
samarium
gadolinium
beryllium
praseodymium
arsenic
scandium
hafnium
dysprosium
uranium
thallium
boron
ytterbium
erbium
tantalum
bromine
holmium
europium
antimony

less than 1/10,000% :

terbium
lutetium

mercury	rhodium
iodine	rhenium
thulium	iridium
bismuth	less than 1/100,000,000% :
cadmium	
silver	neon
indium	krypton
	xenon
less than 1/100,000% :	ruthenium } { amounts
	osmium } { not known
selenium	hydrogen—rock analyses
argon	are variable.
palladium	
less than 1/1,000,000% :	less than 1 billionth % :
	radium
platinum	protactinium
gold	actinium
helium	polonium
tellurium	radon

neptunium and plutonium—tiny amounts, especially in uranium ores, from action of neutrons.

astatine and francium—tiny amounts from radioactive decay.

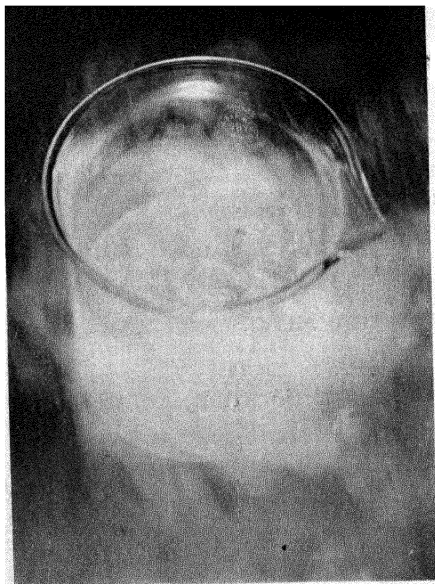
technetium and promethium—tiny amounts from action of neutrons.

americium, curium, berkelium, californium, einsteinium, fermium, mendelevium, and element 102 are essentially non-existent in nature.

It's rather startling to find that two elements, oxygen and silicon, account for three-quarters of the earth's crust. If we included water and air along with the hard crust, we would raise the percentage figures for oxygen, hydrogen, and nitrogen.

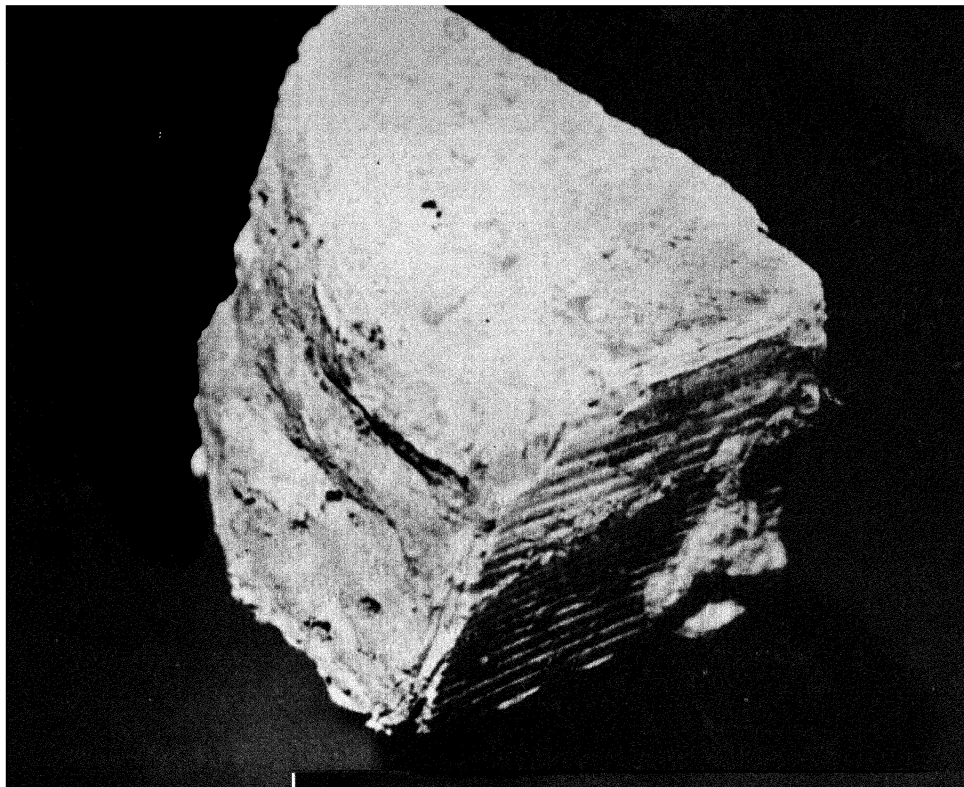
To include living matter and the minerals in the sea would scarcely affect our figures. The elements in living things and in the sea were ultimately derived from the earth's hard crust.

The obvious fact that emerges from our measuring of the quantities of various elements is that oxygen is the main constituent of man's environment. It is an absolute necessity in its free state as a gas; it is indispensable in the form of the liquid compound, water, and it occurs as a part of innumerable solid compounds.

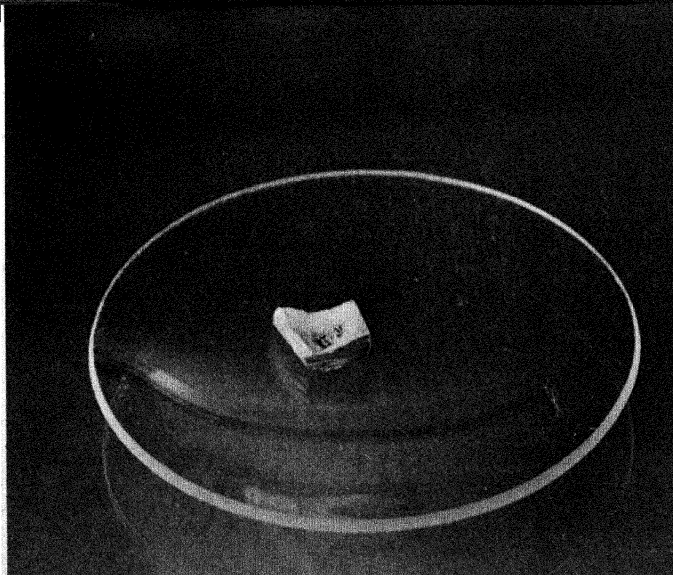


In contrast to oxygen, shown here in its liquid form, the element astatine is so rare that there is only a small fraction of an ounce of it in the entire hard crust of the earth.

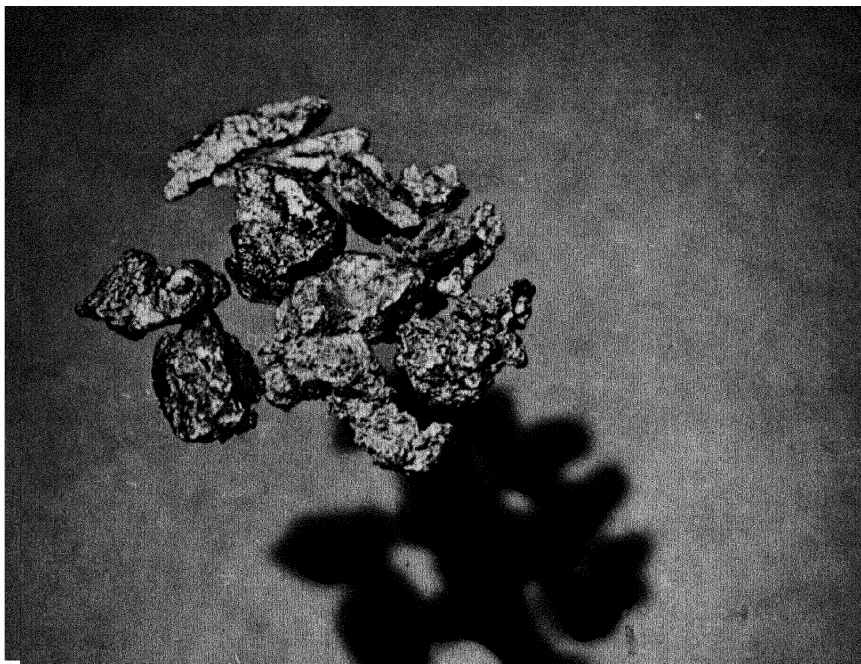
Samples of some of the other elements occurring naturally on the earth appear on the following pages. They are pictured in order of decreasing abundance in the earth's crust.



calcium

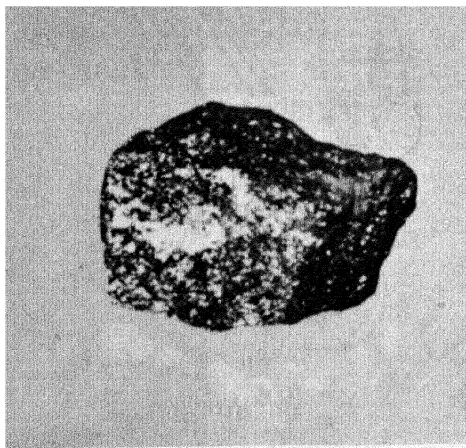
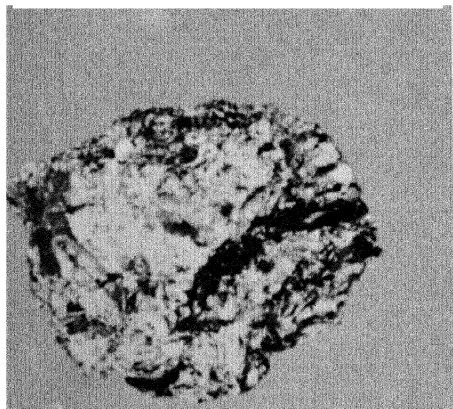


sodium

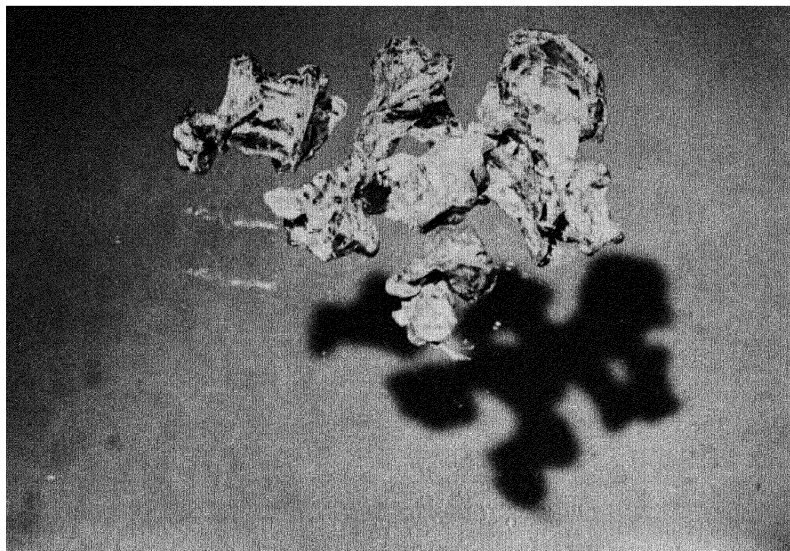


titanium

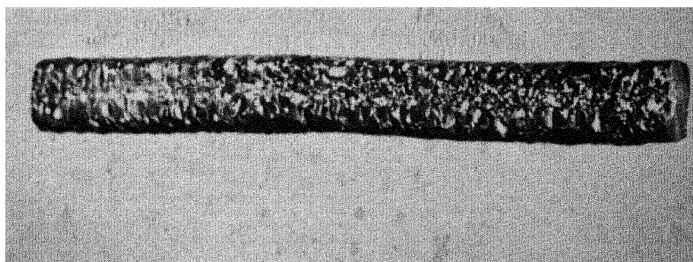
chromium



manganese

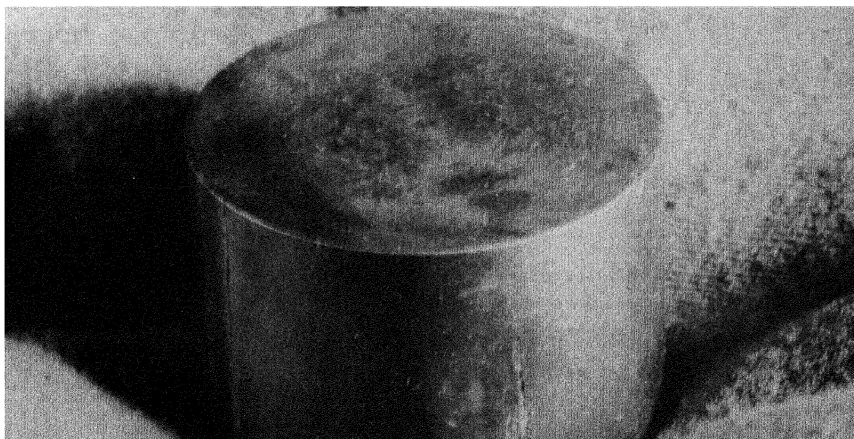


zinc

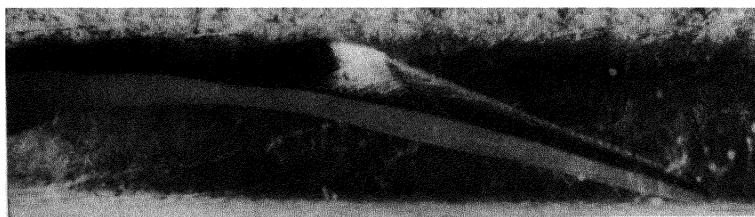


hafnium

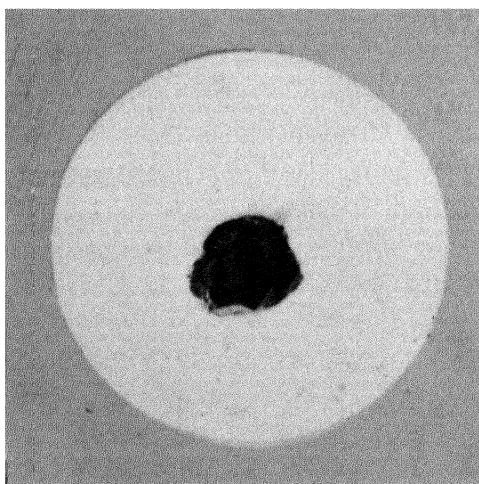
uranium



iodine



plutonium



neptunium

AIR

Perhaps the best way to approach the ninety-odd natural elements on earth is from space, where there is almost no matter at all.

When we get within a few thousand miles of the earth, we are likely to run into a few stray atoms, probably oxygen and nitrogen. Just where the "top" of our atmosphere lies is difficult to say, for it fades imperceptibly into the vacuum of space.

Sixty miles above the earth's surface we are well within the atmosphere, although it is about one million times thinner than at sea level. Here we find both molecules and single atoms of nitrogen and oxygen.

Of course these atoms have weight, and as a result they tend to settle towards the earth. In fact, these atoms and molecules actually would settle if they were not in constant motion, continually bumping into one another.

The entire atmosphere weighs something like 5,000,000-000,000,000 tons, about two million tons of air for every man and woman on earth. Half of this lies below 3.6 miles.

In a very real sense, the atmosphere blankets the earth, filtering the sun's radiation and holding in the sun's warmth during the night.

Air is a simple mixture of gases. Dry air is about 78 per cent nitrogen and 21 per cent oxygen. A cubic foot of air also contains about 16 cubic inches (less than one per cent) of the rare gas, argon, and three-fifths of a cubic inch of the compound, carbon dioxide, plus very small amounts of the

rare gases neon, helium, krypton, and xenon.

Air normally contains water vapor also, but the amount of water vapor varies.

Of all these components of air, oxygen is obviously the most important. We cannot breathe or make a fire without it.

Burning is the chemical reaction involved when oxygen combines with a combustible gas. A jet of air will burn just as well in an atmosphere of illuminating gas as a gas jet burns in air.

In a candle flame, the carbon and hydrogen from the melted wax combine with oxygen to make carbon dioxide and water. The light is due to hot, glowing particles of carbon. As we saw on page 60, a burning candle uses up the oxygen in the air and goes out as soon as the oxygen is consumed.

Free oxygen in the air comes from land and water plants and is made by the process of photosynthesis. Using the energy of sunlight, the green parts of the plants transform carbon dioxide and water into oxygen and carbohydrates like sugar, starch, and plant fibers.

Oxygen participates in all phases of this process, appearing in the gas, carbon dioxide, in the liquid, water, and in the solid carbohydrates. This continual interchange of oxygen emphasizes the close relationship between water—which is 89 per cent oxygen by weight—and air.

Oxygen also occurs in the air in liquid form, as fog, clouds, and rain, and in the water in gaseous form, the free oxygen which allows fish to breathe.

The behavior of oxygen is a fine example of how easily elements may combine to form substances which differ in the way they react and in the way they look, smell, and feel.

OCEAN

About half of the natural elements on earth can be found anywhere in the oceans that cover seven-tenths of the planet's face.

Most of these minerals have been washed into the oceans by the rivers of the world at a rate of billions of tons per year.

Just four cubic yards of sea water—enough to flood a small room to a depth of about a foot—contain the following amounts of various elements:

- 1.5 ounces of some 40 elements, including gold, copper, vanadium, and iodine.

- 1.5 ounces of strontium.

- 7.3 ounces of bromine.

- 11 ounces of bicarbonate of soda (sodium, hydrogen, carbon, and oxygen).

- 6 pounds of potassium sulfate (potassium, sulfur, and oxygen).

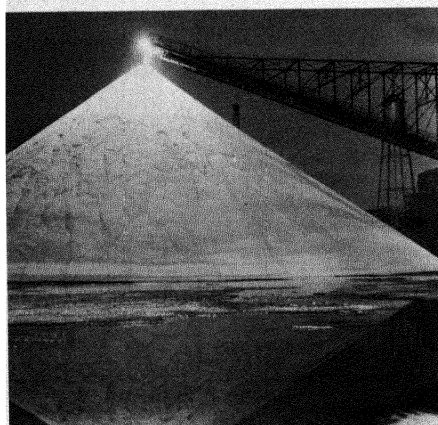
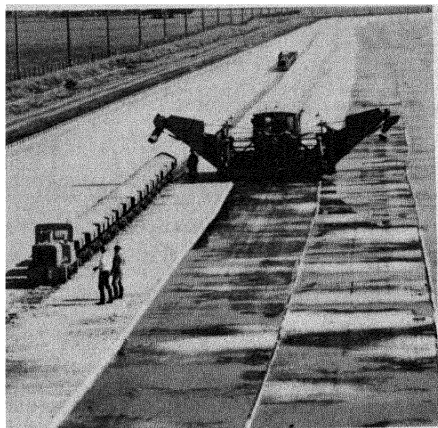
- 9.4 pounds of calcium sulfate.

- 36 pounds of magnesium salts containing nearly 9 pounds of magnesium.

- 186 pounds of ordinary table salt, sodium chloride.

With all these minerals available, it is not difficult to understand how the ocean can support some 8,000 species of plants and 200,000 species of animals.

The ocean is harvested not only for about thirty million tons of fish every year but also for elements such as iodine, bromine, and magnesium and, of course, for its salt.



Salt is harvested by letting sea water into shallow bays and allowing the sun to evaporate the water.

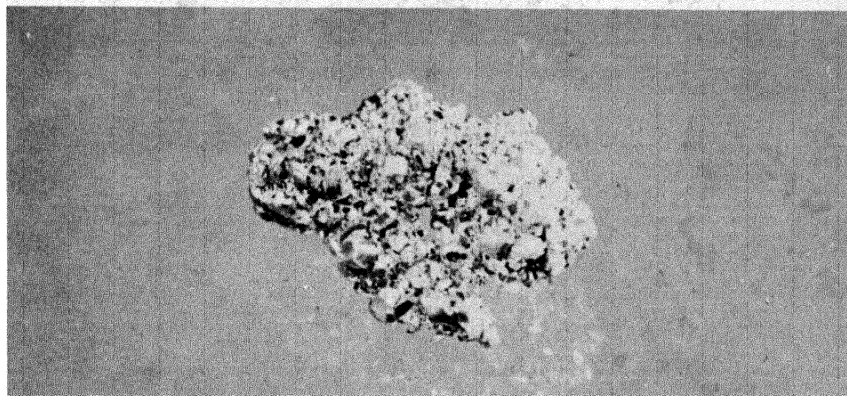
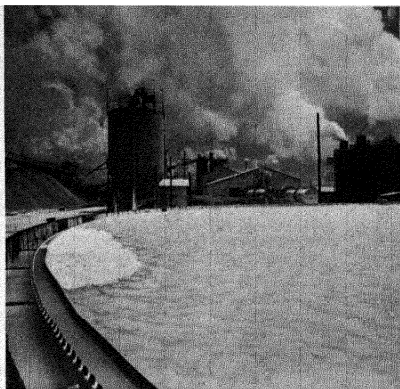
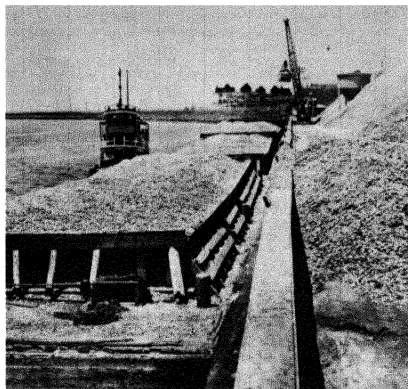
The salt is then carted to stockpiles and the area is flooded again with salt water.

Essentially this same method was known to the ancient Greeks, Romans, and Egyptians, and salt was obtained from the sea in China as early as 2200 B.C.

Mining the seas for magnesium, however, is a very recent industry, although all the magnesium metal produced in the United States today is obtained from the ocean.

Sea water, which contains about four million tons of magnesium per cubic mile, is mixed in great tanks with lime made from barge-loads of oyster shells.

The resulting milk of magnesia—which could indeed be used as a laxative—is converted to magnesium chloride. The metal is then separated from the chlorine by electrolysis. The final product is metallic magnesium, which is only one-fourth as heavy as iron and widely used in the aircraft industry.



The ocean is also rich in other elements, many of which will some day be harvested. It is in fact a veritable gold mine with some \$93,000,000 in gold per cubic mile; unfortunately, it would cost more than that to extract it.

Many of the elements in the ocean—the phosphorus and silicon in crustaceans' shells and the copper in the blood of lobsters, for example—are essential to sea life. If just a few of these vital elements were to disappear, it would ruin the fishing industry and lead to famine in many parts of the world.

THE SOLID CRUST

Although the solid earth itself—rock and soil—contains some ninety elements, all common rocks are composed of oxygen in combination with a limited number of other elements.

By far the most widespread of these oxygen compounds is silicon oxide, or silica (SiO_2), which we know in such various forms as sand, sandstone, quartz, flint, agate, and opal.

All other common rocks except limestone and dolomite consist of silicon oxide in combination with oxides of aluminum, iron, calcium, sodium, potassium, magnesium, or other elements.

The most abundant mineral not containing silicon is calcium carbonate (CaCO_3), which forms chalk, coral, marble, and limestone. Dolomite is a combination of calcium carbonate and magnesium carbonate.

Even artificial stone illustrates the central role of oxygen: Portland cement is made of the oxides of calcium, silicon, aluminum, iron, magnesium, and sulfur.

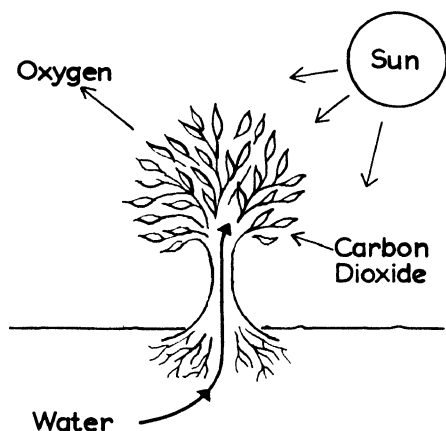
Less abundant elements are scattered throughout the earth, occasionally in a concentrated form as nuggets or as ore deposits.

The earth, like the air and the ocean, has a distinctive character which can be defined by abundance and kinds and combinations of elements involved. But if the three—earth, air and sea—were somehow walled off from one another, no life could exist on the planet. Living things depend utterly upon a constant interchange of elements among the three.

The sea, of course, is rich in elements washed down from the land by water which the atmosphere has carried from the

oceans and released as rain. More subtle but just as constant is the interchange of oxygen, carbon dioxide, and nitrogen.

The most important key in this interchange is the process of photosynthesis, which is accomplished by green plants using energy from sunlight. Carbon dioxide and water are transformed into free oxygen and carbohydrates. These compounds of carbon, hydrogen, and oxygen form the body of the plant and the ultimate source of all our food.



The escaping oxygen replenishes the supply of oxygen in the atmosphere.

Plants also need other elements to grow on, and they draw important elements like phosphorus, potassium, iron, and iodine from the soil or from the sea.

Nitrogen is particularly necessary to plant growth. But despite the fact that the air contains billions of tons of nitrogen, this element is chemically rather inactive and very few living things can use it directly.

The exception is a group of leguminous plants like clover and beans which can assimilate atmospheric nitrogen with the aid of nitrogen-fixing bacteria which grow on nodules on their roots. Such plants will “fix”—that is, transform into usable compounds—enough nitrogen for their own use and also will replenish the nitrogen removed from the soil by other crops.

Animals eat the plants and excrete nitrogen compounds

which fertilize the soil and also decompose to return free nitrogen to the air.

The supply of nitrogen to food crops is partly maintained by artificial fertilizers which man has learned to make by extracting nitrogen from the air and fixing it in the form of solid compounds.



Many elements are present in living matter in very small amounts, but they are nevertheless essential. Here, for

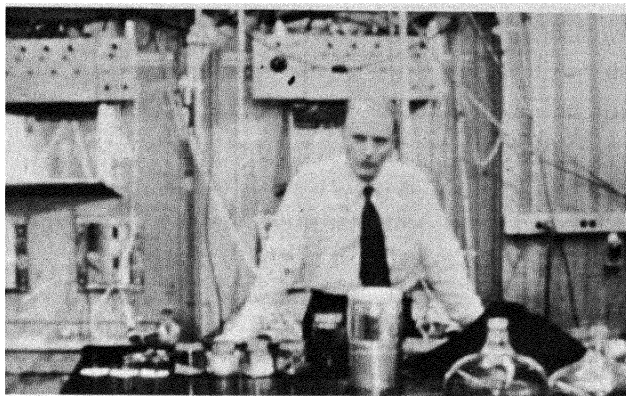
example, are the results of growing identical plants in soil without phosphorus and in soil containing a normal amount of phosphorus.

Lobsters and other lower forms of marine life need copper in their blood just as humans need iron. Kelp requires iodine and potassium, and the sea cucumber requires vanadium. Other living organisms must have such elements as zinc or sulfur or arsenic.

Then there is man himself.

What are little boys made of? Oxygen, carbon, and hydrogen for the most part, and over 60 per cent of it in the form of water.

A 150-pound man contains 97.5 pounds of oxygen, 27 of carbon, 15 of hydrogen, 4.5 of nitrogen, 3 of calcium, and 1.5 pounds of phosphorus, plus about 24 ounces of the following elements: 9 ounces of potassium, 6 of sulfur, 3.5 of sodium, 3.5 of chlorine, one of magnesium, one-tenth of an ounce of iron, and traces of iodine, fluorine, and silicon.



Here are two different versions of Senior Chemist Bernard G. Harvey, one in the more usual state and the other broken down into his various constituent elements.

The 13 gallons of water contain about 94 pounds of oxygen and 12 pounds of hydrogen. This leaves about 3 pounds each of oxygen and hydrogen in the form of other compounds.

The iron is a most important part of the hemoglobin in our blood and the iodine is important in the chemistry of the thyroid gland.

In man, of course, these elements are combined with one another in the form of many, many thousands of very complicated and wonderful compounds.

Reviewing the roles of the various elements on earth, we find oxygen to be outstanding in many respects. It is by far the most abundant element in man himself as well as in the solid crust of the earth and in the oceans. It is the most important element in the atmosphere, and it is essential to man in hundreds of different forms.

The seven next most abundant elements after oxygen account for most of the rocks in the earth's crust and most of the salts in the sea. All the other eighty-four natural elements together account for less than one and one-half per cent of the

weight of the earth's outer skin.

The relative abundance of these other elements, however, does not reflect their relative importance. Rubidium, a relatively unknown element, is four times as abundant as copper and a thousand times as plentiful as iodine.

On the other hand, some familiar names like neon, radium, and plutonium appear almost at the bottom of the list.

Gold, of course, is valued because it is rare, and yet there are twenty elements rarer than gold, not counting the synthetic elements. Titanium, the ninth most abundant element on earth, has only recently come into prominence as a useful material. This is true also of germanium, now used in transistors, and of many of the rare earth elements.

Niobium, which is more plentiful than lead, was considered useless until very recently, although it now promises to become important in the aeronautical industry because of its ability to withstand very high temperatures.

In the future we can expect other little-known elements to become important as uses for them are discovered in such diverse fields as metallurgy, medicine, agriculture, rocketry, and the development of nuclear reactors and thermo-nuclear and solar power sources.

So much for our earthly view of the elements.

As soon as we really broaden our viewpoint and look upon the universe as a whole, the earth proves to be an exception where the particular order of elements according to their abundance may be almost unique.

Hydrogen and helium are only two among the eighty-three elements which together account for just one per cent of the earth's crust. The universe as a whole, however, is 99 per cent hydrogen and helium, and all the other elements in the universe together make up the remaining one per cent.

Elements in Earth's Crust

O

Si

Al

Fe

Ca

Na

K

Mg

misc

Elements in The Universe

H

He



THE UNIVERSE

The elements in the universe are fiercely concentrated in the hot and lonely masses known as stars. And yet empty space itself, the vast and nearly absolute vacuum between the stars, weighs more than a hundred billion suns.

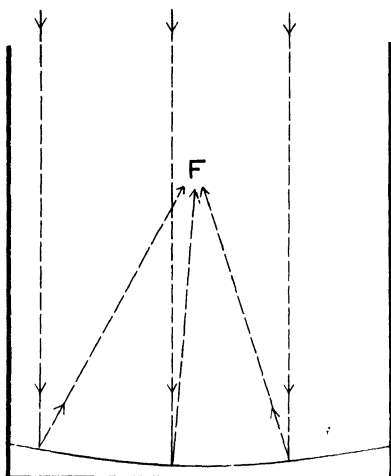
The study of the stars is the oldest of the sciences. The Babylonians observed them systematically 4,000 years before Christ. In Greece in 500 B.C., the Pythagoreans already thought of the earth as a sphere poised in space. And the concept of a solar system with the planets revolving about the sun was put forth about 265 B.C. by Aristarchus of Samos.

His theory was discredited, however, until some 1,750 years later when Nicolaus Copernicus explained the paths of the planets in terms of a sun-centered solar system. Even then people were outraged by the contention that the earth was not the center of the universe.

Modern astronomy was born with the invention of the telescope in Holland in 1608 and its immediate use and improvement in Italy by Galileo Galilei, who used it to discover the moons of Jupiter and the spots on the sun.

Sir Isaac Newton was dissatisfied with the limitations of Galileo's refracting telescope, which operates like a spyglass, and he invented the reflecting telescope.

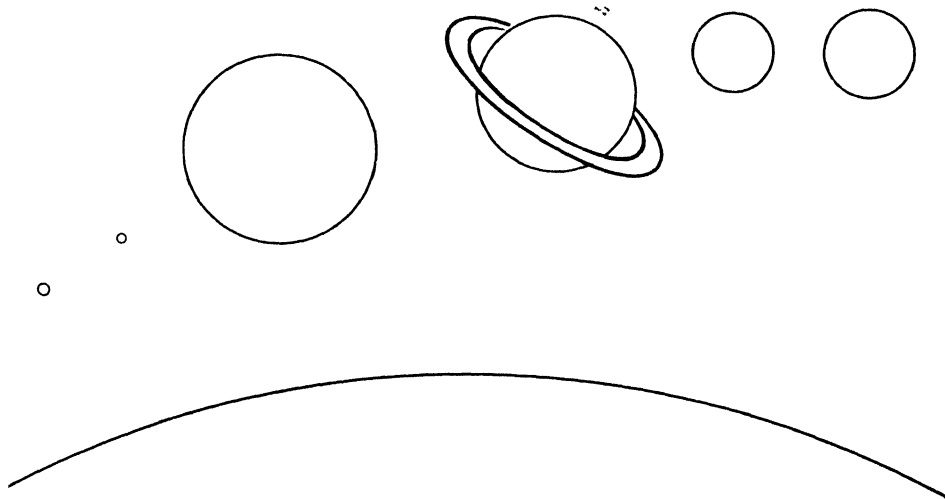
Modern reflecting telescopes range in size from small amateur instruments to the great Palomar telescope with its mirror measuring 200 inches across, but the principle is the same. Parallel rays of light from the stars are reflected to a single focal point by a parabolic mirror.



The star image can be observed directly at the point of focus, or it can be recorded there on a photographic plate. The image also can be reflected out at right angles by a flat mirror and viewed from the side of the telescope.

Repeated observations over the years have revealed nearly all we know about the universe, including the finest details of our own solar system.

Here are the relative sizes of the sun and the planets: Mercury, Venus, Earth, Mars, Jupiter, Saturn, Uranus,



Neptune, and Pluto.

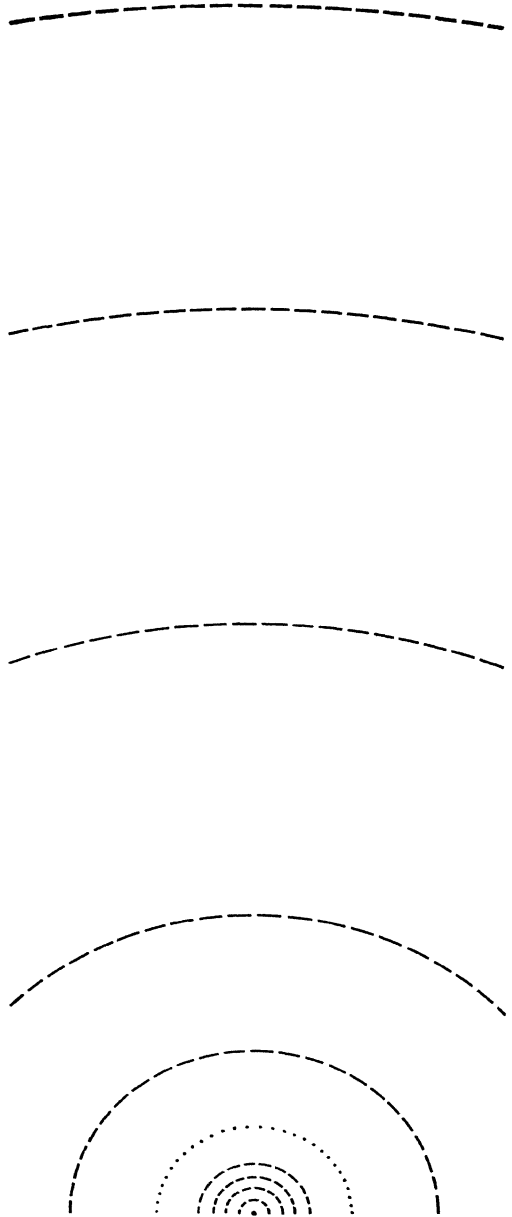
If the earth were the size of a period on this page, the sun would be three and one-half inches in diameter and some 31 feet away. On the same scale, Pluto would be a quarter of a mile from the sun and the nearest star some 1500 miles beyond.

The relative distances from the sun of the various planetary orbits are illustrated here, including the orbit of the asteroids (dotted line) between Mars and Jupiter.

The asteroids, more than 1300 of which have been counted, are apparently the debris of a tenth planet. The largest, Ceres, is 480 miles in diameter.

The arc at the top of the page represents Pluto's average distance from the sun; Pluto actually swings in as close as Neptune, following a very eccentric orbit.

As we should expect, the farther a planet is from the sun, the longer is its "year."



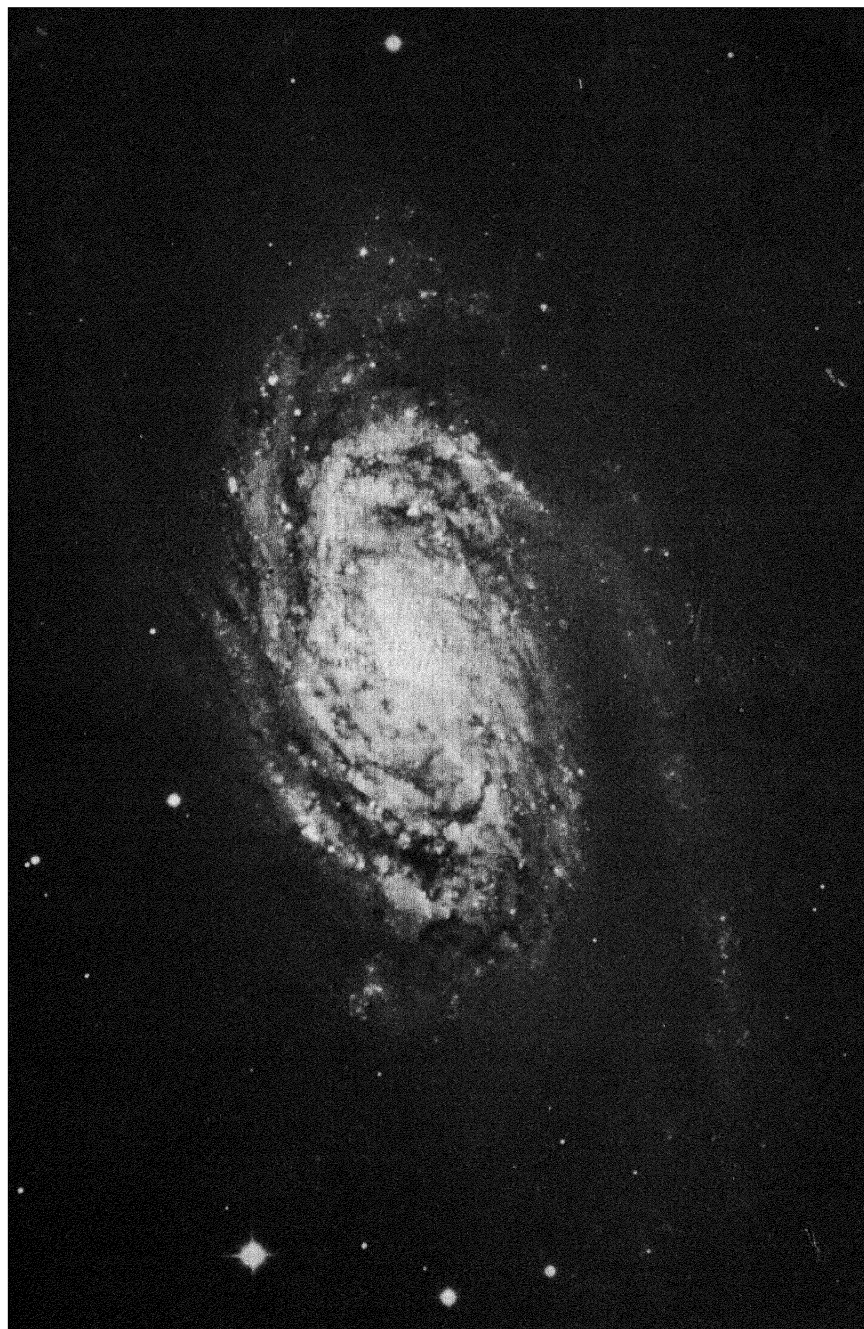
Mercury swings once around the sun every 88 days, Pluto once every 248 years.

The solar system is a very small part of a much larger complex, the local galaxy which we call the Milky Way. Our sun is one of some 100 billion stars in the galaxy and is perhaps 30,000 light-years from its center. The entire Milky Way is lens-shaped, with a radius of some 50,000 light-years. (A light-year, the distance travelled by light in a year at 186,000 miles a second, is about 5,860,000,000,000 miles.)

We on earth are on the inside of the galaxy looking out, and the thin veil of the Milky Way which we see stretched across our night sky is the view out toward the edges.

Other galaxies beyond our own give us an approximate picture of how the Milky Way must appear from the outside—like this spiral in Andromeda when seen edge-on, and like a familiar spiral in Leo when seen “from above.”





Between its billions and billions of stars, each galaxy contains vast stretches of wide open space. Matter exists in this near-vacuum but is so thinly distributed that there are only about 100 atoms per gallon of space. This might be contrasted with the number of atoms in a gallon of earthly air at sea level, roughly 203,000,000,000,000,000,000.

The extent of space is so great, however, that the weight of the sparsely scattered matter between the stars is about equal to the weight of a hundred billion suns.

Our own galaxy is a member of a small local group of galaxies which also includes the great spiral in the constellation Andromeda, shown on page 206.

The entire universe is now estimated to contain something on the order of 1,000,000,000,000 galaxies, many of them in clusters like this group in Coma Berenices, 40 million light-years away.



COSMIC METABOLISM

We are particularly interested in the chemical composition of the sun and stars and of the scattered gases and dust in space. Obviously we cannot reach out for a sample, so we have to identify the elements in space by indirect means. The tool we use is the spectroscope—with camera attached—which is fitted to the viewing end of a telescope.

The spectroscope, as we saw on page 69, spreads out the incoming light so that specific wavelengths or colors which make up the light may be seen and studied individually. This permits us to identify the particular kinds of atoms that emitted these colors, and we thus can discover what elements are present in certain stars or nebulae.

A spectroscope can record the spectral “fingerprints” of glowing sodium in the sun as surely as it can identify light from sodium particles tossed into a laboratory flame.

Clouds of cosmic gas and dust can be seen as dark shapes against a more distant background of stars, or they may shine by reflected starlight. The spectrum of such a bright gaseous nebula reveals both the light reflected from nearby stars and a number of specific wavelengths which are due to the gas having absorbed and then re-radiated this starlight.

Our own earthly atmosphere acts in a similar manner, allowing us to identify elements high in our own skies. Our atmosphere also absorbs and therefore blocks out a great many wavelengths of incoming light, particularly in the ultra-violet region, which complicates our interpretation of the spectral evidence we receive from outer space.

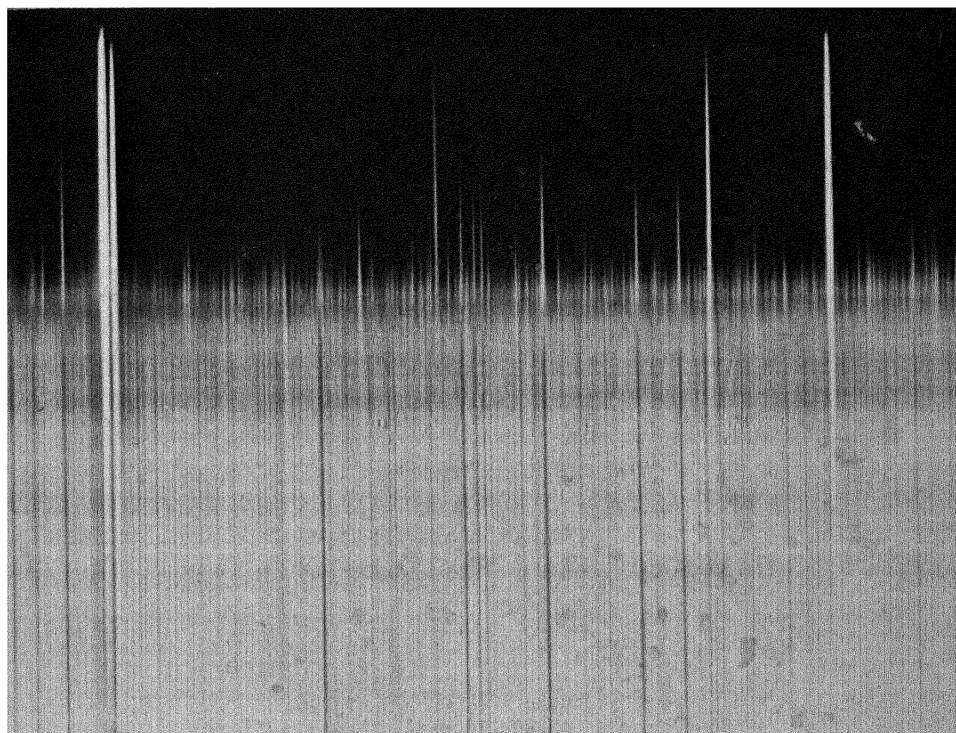
Here is part of the spectrum of the sun, a section ranging from a wavelength of about 3940 Ångström units on the left to 4130 on the right. (An "Ångström," or "Ångström unit," is a unit of length used to measure such things as the radius of an atom or the thickness of a film of oil; there are about 250 million Ångströms to an inch.)

The middle of this part of the spectrum marks approximately the limit of human vision. To the right are the extreme violet wavelengths seen by man. To the left are the invisible violet wavelengths at the beginning of the ultraviolet region which are still well within the sensitivity range of ordinary photographic film.

This spectrum was made with the telescope focused on the rim of the sun, and it shows a few bright "lines" standing out against the black vacuum of space beyond. These lines above the rim originate in the hot gases of the sun's turbulent atmosphere. The most prominent are caused by the light from glowing atoms of calcium, at 3968 Ångströms, and hydrogen (3970 Å), helium (4026 Å), iron (4045 Å), and strontium (4077Å). A second hydrogen line appears at 4101 Ångströms.

The lines of helium in the sun's spectrum were responsible for the discovery of this element in the sun twenty-seven years before it was known on earth.

Both emission lines and absorption lines are clearly visible. The lower part of the photograph contains dark absorption lines superimposed upon the bright spectrum of elements on the surface of the sun. The dark lines are caused by cooler elements above the sun's surface—that is, between the sun and our telescope—which absorb some particular wavelengths of the light radiated from the hot surface below.



The spectral analysis of radiations from the sun and stars, plus other evidence, has revealed a distribution of elements in the universe far different from that on earth.

The most abundant element in the universe is not oxygen, as on earth, but hydrogen, lightest and simplest of all the elements. The sun and stars are about 75 per cent hydrogen and 24 per cent helium. All the other elements make up the remaining one per cent.

Within this one per cent, the elements other than hydrogen and helium may be distributed roughly in the same order of frequency as they are on earth; we don't know at present.

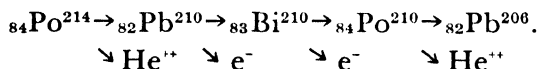
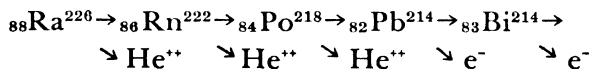
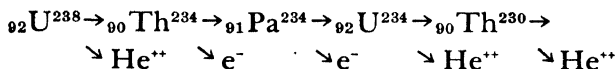
Our solar system was apparently created from the condensation of swirling gases, with the rotary motion conserved to give us the rotation of the planets and of the sun itself.

Some 90 per cent of this matter condensed in the sun and the rest was left over. About 99 per cent of this remainder was hydrogen and helium, which escaped. So about one one-thousandth of the original mass was finally left to form the planets.

Now we may ask, how old is the universe? How old are the elements that constitute the universe?

The most reliable clue we have is the age of our own planet. We can measure the earth's age rather easily and directly because natural uranium decays in a very regular manner.

Uranium undergoes natural radioactive decay, emitting an alpha particle, or helium nucleus, to become thorium-234. The thorium emits an electron and becomes protactinium. This nucleus continues to decay through a series of lighter and lighter isotopes of various elements until it finally reaches a stable state in the form of lead. The entire process involves fourteen distinct steps.



We know the decay rate of uranium; half of any given quantity changes into lead every four and one-half billion years.

If we take a sample of pitchblende, which contains radioactive uranium, we can measure the amount of uranium and the amount of the ultimate decay product, lead. From the ratio of lead to uranium we can determine the age of the ore; and from this information we can deduce the age of the earth.

Using this method, the age of the earth has been estimated to be about five billion years, and there is reason to believe that the universe is just a little older than this, perhaps five and one-half or six billion years.

The question of *how* the universe began is considerably more difficult to answer than that of *when* it began.

In the 1920's it was discovered that the universe is expanding in all directions, with galaxies retreating from one another like spots on the surface of a balloon which is growing constantly larger.

George Gamow of the University of Colorado suggests that the universe at the beginning was an incredibly dense core of radioactive neutrons.

The neutron—with a half-life of about 10 minutes—decays by emitting an electron, thus gaining the positive charge which

makes it a proton. Then the proton, or hydrogen nucleus, could capture more neutrons; some of these would in turn decay into protons within the nucleus, building up the nuclei of heavier and heavier elements.

Another theory which accounts for the creation of elements will be explained here by Otto Struve, Chairman of the Department of Astronomy of the University of California at Berkeley and former president of the International Astronomical Union.

The Gold Medal of the Royal Astronomical Society of London was awarded Dr. Struve in 1944. Much of his current research is centered on the study of the chemical composition of the universe.

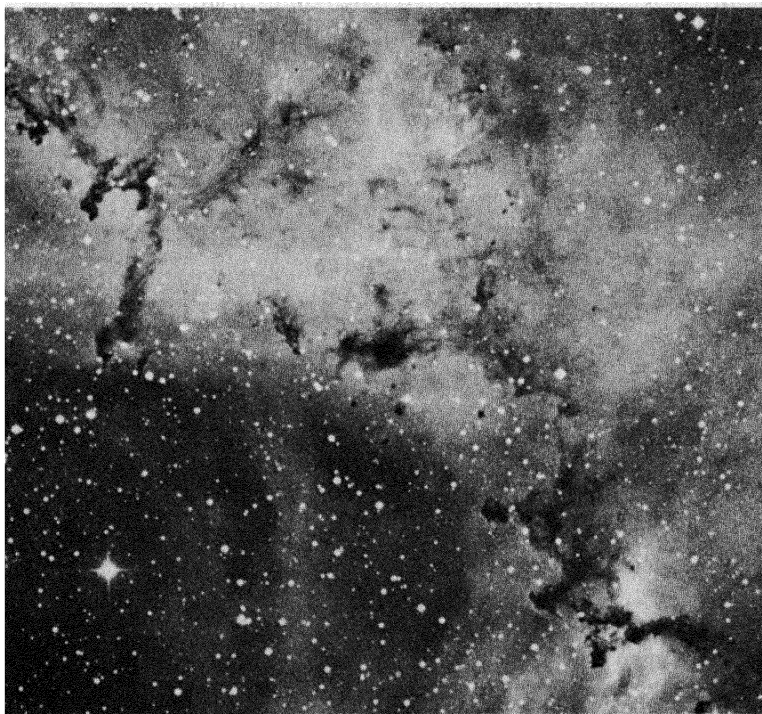




THE BIRTH OF THE UNIVERSE

This is the well-known "Horsehead" nebula in the constellation Orion. It is made mostly of dust, plus some cold hydrogen gas, and it stands out in silhouette against a more distant mass of hot, luminous hydrogen. It is similar to the great black rift and the indistinct dark clouds which obscure great sections of the veil of stars in the Milky Way.

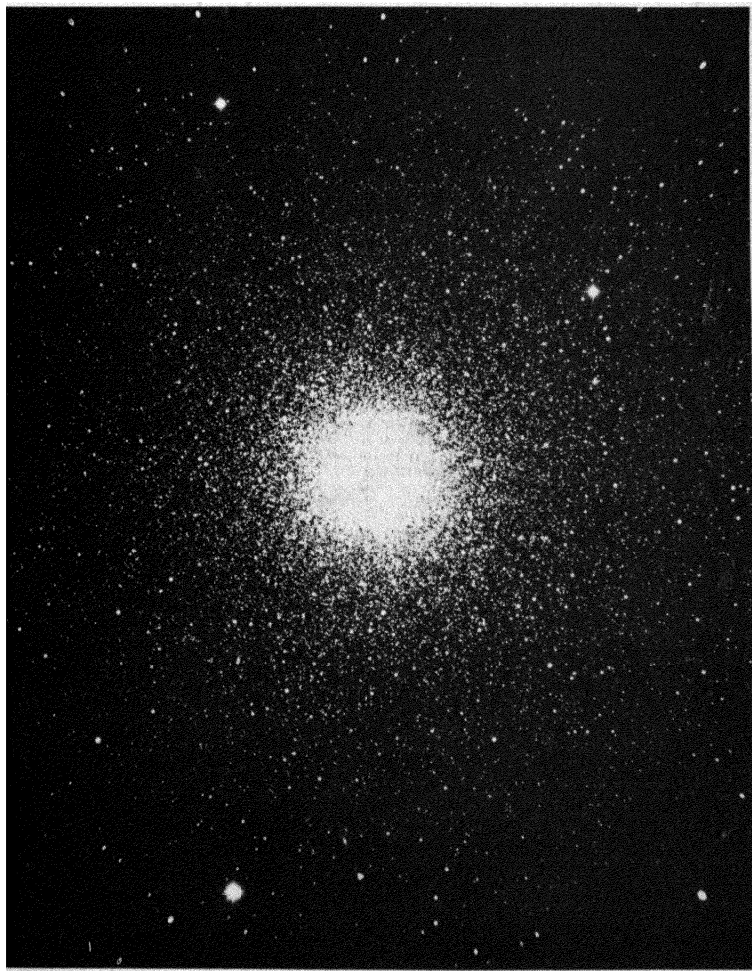
In another photograph of hydrogen gas far out in space, small black spots can be seen. These spots are globules of hydrogen, proto-stars that are now in the process of being formed.



The pair of photographs below actually chronicles the birth of two new stars. At the top is a group of three stars photographed in 1947. In the second picture, taken nearly eight years later, the lower star and the star on the left appear as doubles. The new stars had formed out of hydrogen gas during the intervening eight years.



By way of contrast, here is one of the oldest formations in the universe, a globular star cluster. It dates back five and a half or six billion years, and every star in the cluster was formed at that time. All the gas and dust originally present had condensed at that time into these 100,000 or so individual stars.



Now we come to the question of how the chemical elements are made in the universe.

Professor Gamow's theory attributes the formation of the elements to a condensed neutron gas existing perhaps six billion years ago. And the process of neutrons converting themselves into protons and building more and more complicated elements must have occurred during about the first half hour of the existence of the universe.

Here is a different theory.

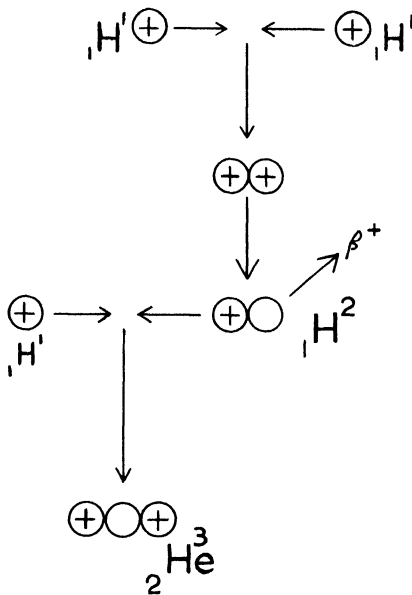
Suppose that the universe, to begin with, consisted entirely of hydrogen gas.

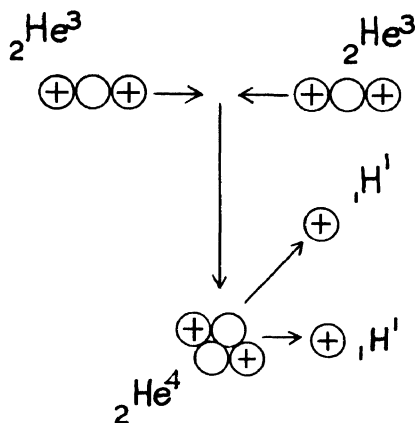
We can start with two hydrogen nuclei, or protons, which collide at very high velocity because of the great temperature of the star. They stick together, and one of the protons converts itself into a neutron by losing its charge in the form of a positron or "positive electron," indicated here by the Greek letter, beta, and a plus sign.

The resulting particle is a deuteron, or nucleus of heavy hydrogen, containing one proton and one neutron.

This deuteron can collide with another hydrogen proton, and when it does, it emits a gamma ray and forms a particle with an atomic mass of three.

We would call this a helium-3 particle because it contains two protons and a neutron and its mass is about equal to that of three single protons.





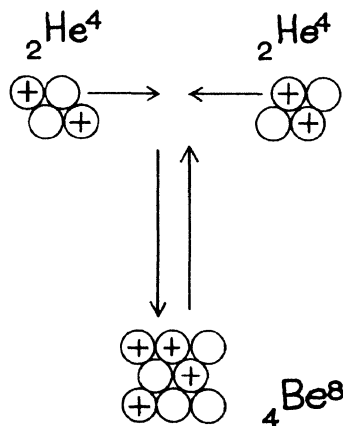
Helium-3 is a rare isotope, lighter than the ordinary helium-4, and it will not capture a proton.

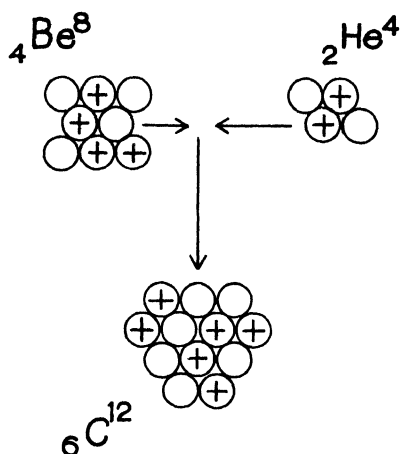
However, it can collide with another similar helium-3 particle. When it does, the two stick together, forming a common helium-4 nucleus and ejecting two hydrogen protons.

This process is the source of energy of the sun and of a great many stars. All the light which we get from the stars is produced essentially as a result of this nuclear process.

The process occurs when the temperature is about ten million degrees. But when the stars are contracting they grow very hot. And when the temperature goes up to about a hundred million degrees, then two alpha particles, or helium nuclei, will collide with tremendous velocity. They form a new particle containing the same number of protons and neutrons that were present in the two alpha particles. This new particle is an unstable isotope of beryllium which we do not find on earth because, if it is ever created on earth, it immediately separates again into two alpha particles.

But inside the stars at any given moment there is about one such beryllium-8 particle present for every billion alpha





particles, so a great many of them are continuously being produced in the stars.

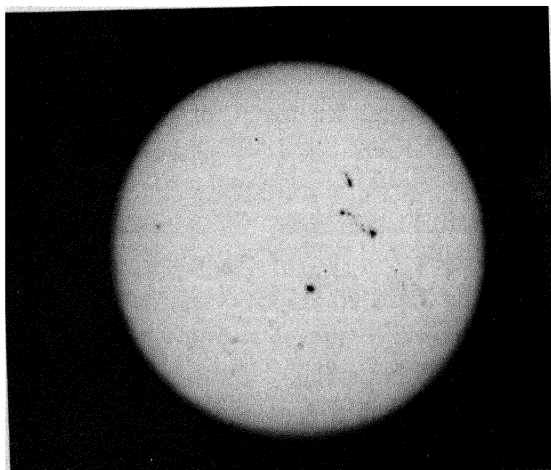
These beryllium-8 particles can collide with alpha particles—helium-4—to form the nucleus of the element carbon.

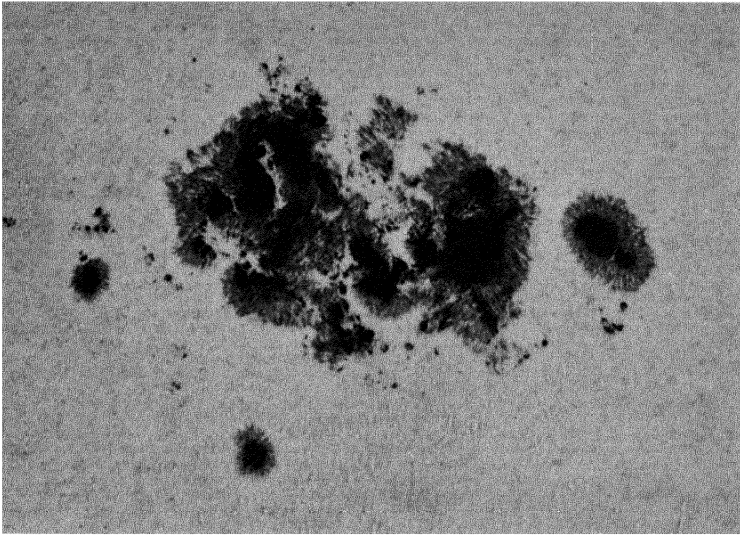
The carbon—ordinary carbon-12—can in turn capture an alpha particle to make oxygen-16. The oxygen then can capture another alpha particle to make neon-20, and so on.

In this way a great many heavy elements can be produced in very hot stars.

But we have not accounted for all the elements. We have not explained, for example, why there is lithium on the surface of the sun, or why there is technetium—a very unstable element on earth—on the surface of a great many stars.

In order to explain this, let us turn to the sun





Through even a very small telescope, sunspots show on the surface.

The sunspots move across the sun's face as the sun rotates, and when they are seen in profile at the edge of the sun they turn out to be tremendous eruptions which we call solar prominences. A prominence can be photographed—as on the following page—by using a coronagraph to artificially eclipse the sun's bright disk.

When we examine the spectrum of these sunspots, we find that they are actually great magnets, magnetic fields of enormous intensity built up in the sunspot in the course of a few hours or a few days.

And this is a situation we are familiar with on earth in the form of the betatron, an electron accelerator.

Electrons are fed into the betatron. Then a magnetic field is rapidly produced, and as the magnetic field increases the electrons are accelerated around a circular path and made to move with tremendous velocity.

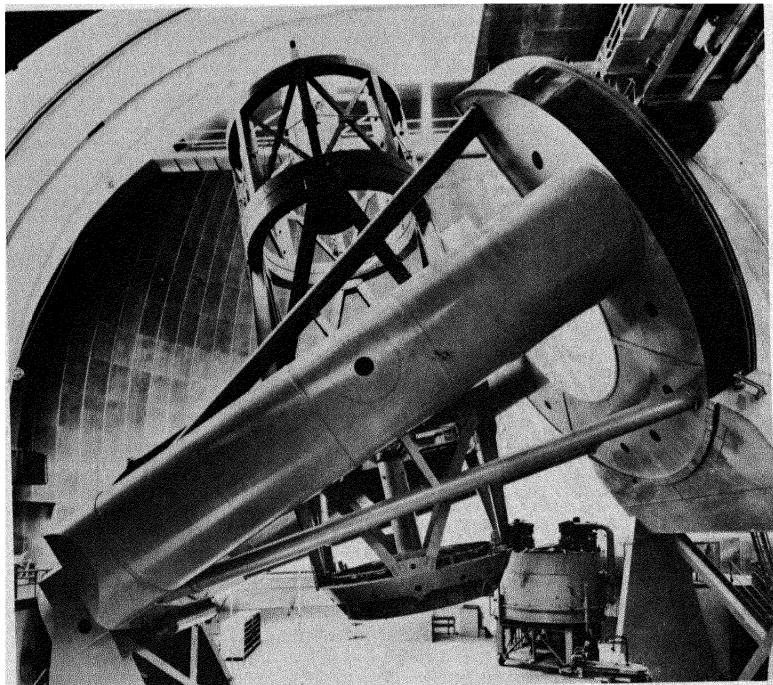


So the magnetic field in a sunspot is a natural particle-accelerator.

And in starspots, where the magnetic field may be many times greater than in sunspots, we expect particles of very high energy. These naturally accelerated particles can produce elements like lithium, beryllium, and boron which are very rare in the stars but which have been detected by analyzing starlight with a spectroscope.

OTTO STRUVE

The study of the structure and chemical composition of the universe proceeds today at an increasing tempo. There is more work to be done than our telescopes can handle. Research projects, for example, are programmed on a tight schedule far in advance for Mt. Palomar's great telescope with its 200-inch reflecting mirror.



Our newest tool is the radio telescope which gathers radiations of radio wavelengths rather than those of visible light. The world's largest telescopes are the 250-foot radio telescope in Manchester, England, and a 350-foot instrument in the U.S.S.R. The United States is designing a 400-foot reflector to be constructed near Sugar Grove, West Virginia.

We are discovering exciting parallels between man's experiments on earth and events in distant stars. For example, one of the heavy transuranium isotopes—californium-254, which decays by spontaneous fission with a half-life of 55 days—is now thought to account for the huge energy production in certain types of super-novae (exploding stars) whose intensity of light emission decays with a half-life of 55 days.

We are justly proud of great achievements of the twentieth century like the invention of particle-accelerators, control of nuclear fission, the potential control of fusion, and the transmutation of one element into another; yet we discover that particle-accelerators, fission, fusion, and transmutation have been going on in the sun and stars for several billions of years.

Science, of course, which began with simple observations and naive curiosity, always returns to nature to check its theories, to search for what was overlooked before, and to gain fresh insight into the character of the universe or the nature of man and his environment.

In the history of the Periodic Table of the Elements, for instance, successive versions reveal a multitude of bad guesses. But new and closer observations of natural events led each time to corrections and refinements so that the Periodic Table today stands as the great unifying principle which correlates the chemical properties of all the known elements and predicts those of elements as yet undiscovered.

APPENDIX

NAMING THE ELEMENTS

The elements were named according to no particular principle except that the majority of the metallic ones end conventionally with the suffix “-ium.”

Chemical symbols for the elements all derive directly from the names, although some of the symbols come from Latin or other alternate names no longer used.

The accepted standard for assigning symbols is based upon a proposal by the Swedish chemist, Baron Jöns Jakob Berzelius, in 1814. Berzelius suggested that the initial letter of the Latin name of the element be used as its chemical symbol. When names of two or more elements began with the same letter, the next distinctive letter of the name was to be added. This system was later extended for the derivation of symbols from non-Latin names assigned to elements.

Berzelius's proposal followed the comparatively recent realization that numerical calculations could be applied to chemical equations. The new symbols were simple and stood for definite atomic weights in the case of each element. They could be set in ordinary type, and they permitted the chemist to visualize chemical reactions simply and effectively.

The nine elements whose symbols derive from old Latin names are sodium (NAtrium), potassium (Kalium), iron (FErrum), copper (CUprum), silver (ArGentum), tin (StaNnum), antimony (StiBium), gold (AUrum), and lead (PlumBum).

The symbols for tungsten and mercury are derived from alternate names, Wolfram and HydrarGyrum.

In general, elements discovered before the seventeenth century bear their ancient names while those discovered after that time have been consciously named by their discoverers.

On the following pages is a list of the elements with their symbols. The symbol and name of each element is followed by the element's atomic number, its physical state at ordinary temperatures if it is not a solid, the date of its discovery (which admittedly is debatable in some cases), and the source of its name.

ORIGIN OF NAMES AND SYMBOLS

- Ac actinium, #89, 1899; named after the Greek "aktis" (ray) since it is radioactive.
- Al aluminum, #13, 1827; named for the aluminum compound, alum, in which it was discovered.
- Am americium, #95, 1944; named for the Americas.
- Sb antimony, #51, Middle Ages; from the Latin "antimonium"; possibly named because it is a tangible (metallic) substance which combines readily with other elements and thus is *anti* (the opposite of) *monium* (an abstract or isolated condition); old Latin name, "stibium," is the source of its symbol.
- A argon, #18, a gas, 1894; from the Greek "argon" (the lazy one).
- As arsenic, #33, Middle Ages; from the Greek "arsenikon" (yellow pigment). The Greeks once used an arsenic compound, arsenic trisulfide, as a pigment.
- At astatine, #85, probably a solid, 1940; from the Greek "astatos" (unstable).
- Ba barium, #56, 1808; named for the mineral barite, or heavy spar, in which it occurs; from the Greek "barys" (heavy).
- Bk berkelium, #97, 1949; named for Berkeley, California.
- Be beryllium, #4, 1797; found in the mineral beryl.
- Bi bismuth, #83, Middle Ages; from the German "Bismuth," possibly a corruption of the "*weisse masse*" (white mass) in which it was found.
- B boron, #5, 1808; from its compound, borax.
- Br bromine, #35, a liquid, 1825; from the Greek "bromos" (bad smell).

- Cd cadmium, #48, 1817; from the Latin “cadmia” (calamine) since it may be found along with calamine.
- Ca calcium, #20, 1808; from the Latin “calcis” (lime, or calcium oxide).
- Cf californium, #98, 1950; named for the state and for the university where the element was discovered.
- C carbon, #6, prehistoric; from the Latin “carbo” (charcoal).
- Ce cerium, #58, 1803; after the asteroid Ceres, discovered in 1801.
- Cs cesium, #55, 1860; from the Latin “caesius” (blue); cesium and rubidium were the first elements to be discovered—by Robert Bunsen and Gustav Kirchhoff—through their spectral lines; cesium is identified by its blue lines.
- Cl chlorine, #17, a gas, 1774; from the Greek “chloros” (light green); chlorine is greenish-yellow.
- Cr chromium, #24, 1797; from the Greek “chroma” (color) since it is used in pigments (e.g., chrome yellow).
- Co cobalt, #27, 1737; from the German “kobold” (goblin); the fact that certain ores yielded cobalt instead of expected metals like copper was attributed to goblins.
- Cu copper, #29, prehistoric; from the Latin “cuprum” or “cyprum”; the main source of Roman copper was once “from Cyprus.”
- Cm curium, #96, 1944; named after Marie and Pierre Curie.
- Dy dysprosium, #66, 1886; from the Greek “dysprositos” (hard to get at).

E	einsteinium, #99, 1952; named after Albert Einstein.
Er	erbium, #68, 1843; named after Ytterby, Sweden, where many of the rare earth ores were discovered.
Eu	europium, #63, 1901; named after Europe.
Fm	fermium, #100, 1953; named after Enrico Fermi.
F	fluorine, #9, a gas, 1886; named after the mineral fluor-spar, from the Latin "fluo" (flow) since fluorspar was used as a flux.
Fr	francium, #87, 1939; named for France.
Gd	gadolinium, #64, 1886; after the Finnish rare earth chemist, Johan Gadolin.
Ga	gallium, #31, 1875; named for France, or Gaul (Gallia); gallium is a solid below 85.55 degrees Fahrenheit.
Ge	germanium, #32, 1886; named after Germany.
Au	gold, #79, prehistoric; the symbol is from the old Roman name for gold, "aurum."
Hf	hafnium, #72, 1923; from Hafnia, Latin name for Copenhagen.
He	helium, #2, a gas, 1868; from the Greek "helios" (sun) since it was first discovered in the sun's spectrum.
Ho	holmium, #67, 1879; from Holmia, Latin for Stockholm.
H	hydrogen, #1, a gas, 1766; from the French "hydro-gène" (water-generator) because water is generated by its combustion.
In	indium, #49, 1863; from the Latin "indicum" (indigo); it was discovered with the spectroscope and its spectrum shows indigo lines.
I	iodine, #53, 1811; from the Greek "iodes" (violet-like).
Ir	iridium, #77, 1804; from the Latin "iris" (rainbow) because of the iridescence of some of its solutions.
Fe	iron, #26, prehistoric; the symbol is from the Roman name, "ferrum."

- Kr** krypton, #36, a gas, 1898; from the Greek “kryptos” (hidden).
- La** lanthanum, #57, 1839; from the Greek “lanthanein” (to lie hidden).
- Pb** lead, #82, prehistoric; the symbol is from the Latin name, “plumbum.”
- Li** lithium, #3, 1817; from the Greek word “lithos” (stone).
- Lu** lutetium, #71, 1907; from Lutetia, old Roman name for Paris.
- Mg** magnesium, #12, 1808; named after “Magnesia lithos” (the Magnesian stone), a white mineral found in the district of Magnesia in ancient Greece.
- Mn** manganese, #25, 1774; from the Italian “manganese,” a corruption of the Latin “magnesi-us” (magnetic).
- Mv** mendelevium, #101, 1955; after Dmitri Mendeleev.
- Hg** mercury, #80, a liquid, prehistoric; named after the planet; mercury is also known as quicksilver and as hydrargyrum, from the Greek “hydros” (water) and “argyros” (silver).
- Mo** molybdenum, #42, 1781; from the Greek “molybdos” (lead); molybdenum was found in what was thought to be lead ore.
- Nd** neodymium, #60, 1885; from the Greek “neo” (new) and “didymos” (twin); neodymium and praseodymium were separated from a material once called “didymium” and thought to be a single element similar to lanthanum.
- Ne** neon, #10, a gas, 1898; from the Greek “neos” (new).
- Np** neptunium, #93, 1940; after the planet Neptune.
- Ni** nickel, #28, 1751; from the German “Kupfernickel” (Old Nick’s copper).

- Nb niobium, #41, 1801; after Niobe, daughter of Tantalus in Greek mythology; niobium was once thought to be identical with tantalum, named after Tantalus, until about 1844. Its original and long-time alternate name was Columbium, with the symbol Cb.
- N nitrogen, #7, a gas, 1772; from the French “nitrogène” (generator of niter, which is sodium or potassium nitrate).
- Os osmium, #76, 1804; from the Greek “osme” (a smell).
- O oxygen, #8, a gas, 1771; from the French “oxygène” (generator of acid); it was once thought to be an essential component of acids.
- Pd palladium, #46, 1803; after the asteroid, Pallas, discovered in 1801.
- P phosphorus, #15, 1669; from the Greek “phosphoros” (light-bringing).
- Pt platinum, #78, 16th century; from the Spanish “platina” (silver).
- Pu plutonium, #94, 1940; after the planet Pluto.
- Po polonium, #84, 1898; discovered by the Curies and named for Marie Curie’s native country, Poland.
- K potassium, #19, 1807; from “potash,” which is potassium carbonate, found in wood ash; the symbol is from the Latin name, “kalium.”
- Pr praseodymium, #59, 1885; from the Greek “prasios” (green) and “didymos” (twin) because of its green salts and its early confusion with neodymium.
- Pm promethium, #61, 1945; after Prometheus, the Greek Titan who stole fire from heaven for man’s use.
- Pa protactinium, #91, 1917; from “proto” (first) and actinium, because it disintegrates into actinium.
- Ra radium, #88, 1898; from the Latin “radius” (ray) because it emits radioactive rays.

- Rn radon, #86, a gas, 1900; named after radium, with the suffix “-on” common to the inert gases; it is a product of radium and is itself radioactive. It was once called “niton” (shining) with the symbol Nt.
- Re rhenium, #75, 1925; after the Rhine (“Rhenus” in Latin).
- Rh rhodium, #45, 1803; from the Greek “rhodon” (rose) because some of its salts are rose-colored.
- Rb rubidium, #37, 1861; from the Latin “rubidus” (red); it was discovered with the spectroscope, and its spectrum shows red lines.
- Ru ruthenium, #44, 1844; after Russia (“Ruthenia” in Latin).
- Sm samarium, #62, 1879; from “samarskite,” a mineral named after a Russian mining official, Col. Samarski.
- Sc scandium, #21, 1879; after Scandinavia.
- Se selenium, #34, 1818; named for the moon (“selene” in Greek) because of its resemblance to tellurium, which was named for the earth.
- Si silicon, #14, 1824; from the Latin “silex” or “silicis” (flint); flint is silicon dioxide.
- Ag silver, #47, prehistoric; the symbol is from the Roman name, “argentum.”
- Na sodium, #11, 1807; from caustic soda, from which it is made; the symbol comes from the Latin “natrium.”
- Sr strontium, #38, 1808; from the mineral strontionite, named for Strontian, Scotland.
- S sulfur, #16, prehistoric; from the Latin name, “sulfur.”
- Ta tantalum, #73, 1802; named, because of the difficulty of isolating it, after Tantalus, son of Zeus and father of Niobe, who was punished by being forced to stand up to his chin in water which receded whenever he tried to drink.

- Tc technetium, #43, 1937; from the Greek “technetos” (artificial) because it was the first artificial element.
- Te tellurium, #52, 1783; named after the earth (“telluris” in Latin).
- Tb terbium, #65, 1843; after Ytterby, Sweden.
- Tl thallium, #81, 1861; after the Greek “thallos” (a young shoot) because of the bright green line in its spectrum.
- Th thorium, #90, 1828; after the mineral thorite, which is named for Thor, Norse god of thunder.
- Tm thulium, #69, 1879; after Thule, an old name for the Scandinavian northland.
- Sn tin, #50, prehistoric; the symbol is from the Latin name, “stannum.”
- Ti titanium, #22, 1791; after the Greek giants, the Titans.
- W tungsten, #74, 1783; after the Swedish “tung sten” (heavy stone); its symbol comes from an alternate name, wolfram.
- U uranium, #92, 1789; after the planet Uranus.
- V vanadium, #23, 1801; after Vanadis, Norse goddess of love and beauty.
- Xe xenon, #54, a gas, 1898; after the Greek “xenos” (stranger).
- Yb ytterbium, #70, 1878; after Ytterby, Sweden.
- Y yttrium, #39, 1794; after Ytterby, Sweden.
- Zn zinc, #30, 17th century; from the German “Zink.”
- Zr zirconium, #40, 1789; from the mineral zircon, in which it is found.
- 102 element #102, discovered in 1958, had not been named when this book was published. The element was reportedly discovered in 1957 and the name “nobelium” was suggested for it. However, experiments to date attempting to confirm this 1957 report have not been successful, and the name “nobelium” may not be kept.

1 H 1.0080								
3 Li 6.940	4 Be 9.013							
11 Na 22.991	12 Mg 24.32							
19 K 39.100	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.95	24 Cr 52.01	25 Mn 54.94	26 Fe 55.85	27 Co 58.94
37 Rb 85.48	38 Sr 87.63	39 Y 88.92	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc	44 Ru 101.1	45 Rh 102.91
55 Cs 132.91	56 Ba 137.36	57-71 La Series	72 Hf 178.50	73 Ta 180.95	74 W 183.86	75 Re 186.22	76 Os 190.2	77 Ir 192.2
87 Fr	88 Ra 226.03	89-103 Ac Series	(104)	(105)	(106)	(107)	(108)	

Lanthanide
Series

57 La 138.92	58 Ce 140.13	59 Pr 140.92	60 Nd 144.27	61 Pm	62 Sm 150.35	63 Eu 152.0
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Actinide
Series

89 Ac 227.04	90 Th 232.05	91 Pa 231.05	92 U 238.04	93 Np 237.06	94 Pu	95 Am
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								2 He 4.003
								10 Ne 20.183
								8 O 16.000
								7 N 14.008
								6 C 12.011
								5 B 10.82
								18 Ar 39.944
								17 Cl 35.457
								16 S 32.066
								15 P 30.975
								14 Si 28.09
								13 Al 26.98
28 Ni 58.71	29 Cu 63.54	30 Zn 65.38	31 Ga 69.72	32 Ge 72.60	33 As 74.91	34 Se 78.96	35 Br 79.916	36 Kr 83.80
46 Pd 106.4	47 Ag 107.880	48 Cd 112.41	49 In 114.82	50 Sn 118.70	51 Sb 121.76	52 Te 127.61	53 I 126.91	54 Xe 131.30
78 Pt 195.09	79 Au 197.0	80 Hg 200.61	81 Tl 204.39	82 Pb 207.21	83 Bi 208.99	84 Po	85 At	86 Rn

71 Lu 174.99	70 Yb 173.04	69 Tm 168.94	68 Er 167.27	67 Ho 164.94	66 Dy 162.51	65 Tb 158.93	64 Gd 157.26
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(103)	102	101 Mv	100 Fm	99 E	98 Cf	97 Bk	96 Cm
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FOR FURTHER READING

Among the good books written about various aspects of the chemical elements, the following are some of the most exciting:

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