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AT HOME AMONG THE ATOMS

BY THE SAME AUTHOR

YOUNG CHEMISTS & GREAT DISCOVERIES

AT HOME AMONG THE ATOMS

A First Book of Congenial Chemistry

By

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DEDICATED TO
SIR JAMES WALKER
PROFESSOR EMERITUS OF CHEMISTRY IN THE
UNIVERSITY OF EDINBURGH
AND TO THE MEMORY OF
ALEXANDER SMITH
LATE PROFESSOR OF CHEMISTRY
IN COLUMBIA UNIVERSITY, NEW YORK CITY,
TO WHOM, ABOVE ALL OTHERS, THE
AUTHOR'S THANKS ARE DUE FOR THE
INSPIRATION AND ASSISTANCE WHICH
THEY SO GENEROUSLY AFFORDED
HIM IN HIS EARLY CAREER

INTRODUCTION

I observe that of late chymistry begins, as indeed it deserves, to be cultivated by learned men who before despised it; and to be pretended to by many who never cultivated it, that they may be thought not to be ignorant of it.—ROBERT BOYLE, "The Sceptical Chymist."

THE title of this volume—"At Home Among the Atoms," and its sub-title—"A First Book of Congenial Chemistry"—have been carefully chosen, in order to indicate, as clearly as possible, its informal as well as its chemical character. The material things which make up this world are first of all classified into *substances* with definite and distinct properties. These substances, mostly quite complex, are next shown to be capable of dissection into a limited number of simple substances, or *elements*. Following upon this, attention is turned to *atoms* as the smallest individual particles of which all elements are composed; and finally the inner fabric of the atom itself, with its orderly array of *protons* and *electrons*, is revealed. All of this has been done in a frank and informal fashion, the author attempting to function throughout not as a remote and unapproachable specialist but rather as a host conducting an intimate friend for the first time through a beautiful (though difficult) piece of country, eager to exhibit and to explain to his visitor in detail every object of peculiar interest.

Many ponderous tomes, of course, have been written about the atom in recent years, while the popular press is saturated at the present time with light and airy summaries of the wonderful discoveries that have lately been made regarding the ultimate structure of matter. This book, however, is not just "another treatise on the atom," nor does it pretend to furnish predigested fodder for the benefit of idlers who merely wish to be entertained. It is a little difficult, indeed, to indicate in brief exactly what the author's purpose is in this volume, and probably the reader will not appreciate it fully until he has read through one or two chapters. Nevertheless, a few more words here on what is, and what is not, aimed at will be appropriate.

A casual glance at the table of contents will suffice to show that this book has *not* been written for the scientific expert. It is not even *intended* to serve, primarily, as a text-book for beginning students in chemistry classes, although it is possible that a few unorthodox teachers may be tempted to use it in this connection, and quite probable that more may see fit to recommend it to their pupils as an auxiliary reference book. The subject is developed logically and systematically, it is true, as befits an exact science, but no attempt has been made to cover every item included in the normal examination schedule, and the unconventional (often almost flippant) methods of presentation employed are not calculated to meet the approval of the hardened pedagogue.

On the other hand, as has already been stated, this volume is not designed to appeal to people who "think they think," or who wish to get the latest news about the atom at a single sitting in words of one syllable. The interest of the average citizen in the progress of chemistry at the present time is, in the opinion of the author, on a higher plane than this. There is a widespread desire on the part of the public *not merely to* *know but also to understand.* Intelligent men and women feel that they are missing something of vital importance if they do not keep up with recent advances in the science—advances which are so revolutionary in their nature that, in spite of their apparent remoteness from present practicality, they must eventually affect the whole political and economic life of the nation. The lessons of the Great War have not been forgotten. Then the very existence of our country hung upon the success with which its infant chemical industries could compete with the enemy output of high explosives and poison gases; now it is upon the application of chemistry to the problems of peace—dyestuffs, fertilizers, artificial silk, fuel conservation, new alloys, synthetic rubber, and hundreds of other household and industrial necessities—that our commercial prosperity depends. Far and wide the word has been spread that this is the Age of Chemistry; yet the sort of information regarding chemistry that the serious-minded searcher craves to find is almost unattainable.

The typical text-book on the subject is repellent to the layman; it stresses a mass of routine facts, most of which are of interest to the professional only. Popular articles, with their pretty pictures of protons and electrons, are all too likely to leave the reader either muddled or misinformed. It is really an impossible task to present the results of recent work on atomic structure to the non-chemist intelligently and intelligibly without leading up to the topic slowly and carefully by explaining the previous knowledge on which such work was based. The latest developments are necessarily also the most uncertain; even the experts are still at odds over many of the main points. How, then, can the amateur be expected to orient himself and to distinguish fact from fancy ?

To enable him to succeed in these ambitious but laudatory efforts constitutes, in brief, the principal purpose of this book. It does not presuppose any familiarity at all with the formal side of chemistry—that side, indeed, is suppressed to a minimum throughout—but it does insist on a degree of interest sufficient to incite the reader to co-operate faithfully in accompanying the author up each single slant of the trail which the main argument follows until the summit is reached. In order to promote this joint achievement, the path has been carefully cleared of all unnecessary obstructions, and no legitimate device for easing the ascent has been neglected. Homely analogies and alleged witticisms furnish temporary distractions from the abrupt precipices of strict fundamental theory which cannot be avoided, and a helping hand is stretched out to assist the climber at all difficult points.

To sum the matter up in a sentence, the writer has done his level best to keep the present volume essentially *human*, and if only a few people find as much enjoyment in reading it as he has found in writing it, he will consider his labours amply repaid. After all, chemistry *is*, as the author of *Creative Chemistry* has said, " the joyous science," and those who are lucky enough to be chemists should be only too happy to let their fellow-men know something of their good fortune. Nevertheless, the complaint of Robert Boyle in 1661 is still only too frequently justified : " A person anything versed in the writings of chymists cannot but discern by their obscure, ambiguous, and almost ænigmatical way of expressing what they pretend to teach, that they have no mind to be understood at all, but by *the sons of Art* (as they call them), nor to be understood even by these without difficulty and hazardous trials."

It is sincerely hoped that this book will escape the above accusation, and if the reception accorded by the public to this first volume is sufficiently encouraging, a second instalment dealing with the kinetic-molecular theory and discussing the preparation and industrial applications of typical chemical compounds will be published as soon as the author's other responsibilities permit.

JAMES KENDALL

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THE ALCHEMIST

Chapter I

THINGS IN GENERAL

The world is so full of a number of things,
I'm sure we should all be as happy as kings.
—Robert Louis Stevenson, *A Child's Garden of Verses.*

The trouble with these literary people is that they are so terribly indefinite. It is not our intention to enter at this point into a weighty discussion as to the truth of the suggestion that kings are happier than common people. A great deal of evidence might easily be brought forward to the contrary, leading off with the statement by Shakespeare,

Uneasy lies the head that wears a crown.

We do want to look more attentively, however, into the phrase " a number of things."

At once the question arises : What is this number ? The answer given by a non-scientific person, remembering the complexity of everyday life, undoubtedly would be that the number was far too large for any reasonable estimation and that one might as well let the matter rest. The scientist, nevertheless, has a peculiar sort of mind, and he refuses to run away from a situation of this character. When he encounters a large army of isolated objects, his instinctive reaction is first to try to divide the unwieldy mob into a few distinct types and then to examine each type separately and in detail. After all, the average man does the same thing unconsciously.

Suppose we put you in a familiar situation and see how *you* react.

Imagine that you are crossing a crowded street and become confused on account of the traffic. You might well be tempted to amend Stevenson's verse by muttering:

> The world is so full of a number of cars,
> I'm sure we should all be much happier in Mars.

Although it is quite likely that the canals of Mars are equally congested with motor-boats, so that you may be just as well off where you are. But, returning to your present predicament, do you *really* think of each single motor-car as one of a very large number? It must be granted that every car does carry a distinguishing number, and that this number is useful to remember in case a car does happen to strike you; but you rarely attempt to recall any particular car by its number otherwise. Rather you say indignantly, "I was nearly knocked down by an Austin Seven to-day," or proudly, "I almost wrecked a Rolls-Royce in Piccadilly this afternoon." All of the many millions of automobiles that render modern life hazardous, then, are included in a limited number of distinct makes, ranging from the lowly two-seater to the lordly Rolls-Royce.

This principle is exactly the one that the chemist follows with regard to "things." It is true that his task is more difficult since the field is much wider, but, as we shall see, he has gradually succeeded in dissecting all of the material objects which fall into the general scheme of things and in grouping their constituent parts into a restricted number of types, comparable with the number of different makes of motor-cars. In recent years he has gone still farther. He has managed to devise a sub-microscopical point of view under which the ultimate number of things which make up the whole universe has been reduced to two! It is our object in the ensuing pages to show, as simply as we can, how this has been done.

Before we actually begin, however, let us pause for a space in order to take this opportunity to anticipate two questions which are certain to arise in your mind sooner or later. The first is: "Why should normal men and women want to know anything about the inner workings of a subject which has the reputation of being so abstruse and unsavoury as chemistry?" The second is: "Why should a professor of chemistry feel constrained to cater to any such curiosity by attempting to write an informal and popular treatise on his chosen subject, disclosing its most sacred secrets to the view of the profane?"

These two questions are interconnected, and we may answer them at one and the same time. To some extent, indeed, they have already been covered in the Introduction, and our suspicion that the normal man or woman does nurse a desire to know something about chemistry is already vindicated by the fact that you have a copy of this book in your hands. Nevertheless, in order to fortify you to continue on the right path, let us put you in touch with a few opinions which have been expressed by various people of importance of past and present times on the points at issue. We shall not quote from chemists, since they are open to the accusation of being biased regarding the value of their science; neither shall we quote from kings, since they are exposed to too many minor distractions to supply us with material of worth. Presidents and friends of presidents, princes and belted earls, however, are not taboo, and we shall accordingly call upon several such notables, of British or of American origin, for international moral support. Summoning our witnesses in chronological order, the first of them shall be Thomas Jefferson, the second President of the United States, who wrote to his friend Dr. Ewell of Virginia in 1805 as follows:

> Of the importance of turning a knowledge of chemistry to household purposes, I have been long satisfied. The common herd of philosophers seem to write only for one another. The chemists have filled volumes on the composition of a thousand substances of no sort of importance to the purposes of life; while the arts of making bread, butter, cheese, vinegar, soap, beer, cider, etc., remain unexplained. Chaptal has lately given the chemistry of wine-making; the late Dr. Penington did the same as to bread, and promised to pursue the line of rendering his knowledge useful to common life; but death deprived us of his labours. Good treatises on these subjects should receive general approbation.

The only comment that we need to make upon this letter here is that before it is possible to apply a knowledge of chemistry successfully to household purposes, to home-brewing, or to anything else, it is clearly necessary to know a little about some of the main principles of the science. We shall, indeed, find it advisable to occupy the greater part of this volume with the development of such principles, leaving their household and industrial applications for later presentation. But we trust that we shall succeed in doing this in such a way as to avoid the charge that we are writing only for philosophical experts and not for thirsting amateurs.

John Adams, the direct successor of Jefferson in the presidency, may be called upon to testify next. Here is what he has to say on the matter:

> We are all Chymists from our cradles. All mankind are Chymists from their cradles to their graves. The Material Universe is a chymical experiment.

Nothing can be fairer than that, although, as we shall see later, Uncle John was not always so shrewd in his statements.

Now it is time for us to re-cross the Atlantic and listen to a few home-truths. In his presidential address before the British Association at Aberdeen in 1859 the Prince Consort, always an enthusiastic patron of science, said:

> The labours of the man of Science are at once the most humble and the loftiest which man can undertake. He only does what every little child does from its first awakening into life, and must do at every moment of its existence; and yet he aims at the gradual approximation to divine truth itself. To *arrange* and *classify* the general universe of knowledge is the first, and perhaps the most important, object and duty of science.

Coming down to more recent times, the Prince Consort's great-grandson endorsed these remarks, in a more practical way, when he, in his turn, presided as Prince of Wales over the same British Association for the Advancement of Science at Oxford in 1926:

> Scientific research, properly applied and carried out, is never wasted. From the purely materialistic point of view, scientific research is not a luxury; for the community it is probably the cheapest possible form of investment. In order, however, that the community may fully realize all that it owes, or all that it might owe, to the advancement of science

the channels of communication between research and the public mind have to be kept clear, maintained and widened. The non-scientific public is accustomed to view science as it might view a volcano ; prepared for the eruption of some new discovery from time to time, but accepting the effects of the eruption without realizing the processes which led up to it during the preceding period of quiescence.

What is said here of science in general is equally applicable, of course, to chemistry in particular. Indeed, the Earl of Balfour, in his Messel memorial lecture before the Society of Chemical Industry at London in the same year, stated as follows :

> Chemistry is not merely a great science among other sciences, but a science which permeates the whole of life. The greatest of all the aspects of scientific study is the discovery of how the world we live in is constructed, what its real intimate character is.

Hear now, finally, what Herbert Hoover said in 1926, urging the establishment of a national fund in the United States for the support of research in pure science :

> No one can estimate the value to the world of an investigator like Faraday or Pasteur or Millikan. The assets of our whole banking community to-day do not total the values which these men have added to the world's wealth. . . .
>
> And the more one observes the more clearly does he see that it is in the soil of pure science that are found the origins of all our modern industry and commerce. In fact, our civilization and our large populations are wholly builded upon our scientific discoveries. . . .
>
> The progress of civilization, as all clear-thinking historians recognize, depends in large degree upon " The increase and diffusion of knowledge among men." Our nation must recognize that its future is not merely a question of applying

present-day science to the development of our industries, or to reducing the cost of living, or to eradicating disease and multiplying our harvests, or even to increasing the general diffusion of knowledge. We must add to knowledge, both for the intellectual and spiritual satisfaction that comes from widening the range of human understanding and for the direct practical utilization of these fundamental discoveries. If we would command the advance of our material and, to a considerable degree, of our spiritual life, we must maintain this earnest and organized search for truth.

These authoritative utterances ought to be sufficient to convince you that it is worth while going on to learn what we have to say in the rest of this book. We cannot refrain, however, from inflicting upon you one more extract, from a book written by Jefferson's friend Thomas Ewell in 1806, which should clear your mind completely of all wonder as to why we find ourselves compelled to put before you a volume of this character.

No individual can now do too much in promoting the study of chemistry. Ye free agents! Ye guardians of the young! Can you allow those under your care to neglect learning the principles of this all-important science? What then will you say, when arraigned at the bar of justice, before a Creator and an assembled Universe, for neglect of your duty! Your hoary locks will not cover you! the number of the accused will naught extenuate—and in vain will you deny the charge! The children of successive generations will rise up around you! In the face of Heaven they will bitterly complain of the beauties to which they were insensible! While humanity, bewailing her misfortune, will recount the number of blows she received, and the number of denials made to her call, because these creatures could not perceive the pleasures of her paths!

Fellow citizens, I really feel incapable of properly stating the beauty, the connection, and the importance of the science of Chemistry. For you to know these you must know the science itself. This understood, your souls would feast on the eternal beauties of nature, while you were inhaling the grateful incense arising from your services to the cause of humanity.

Now that you have read through this fervid oration, you will realize that, after all, we are acting merely in self-defence.

Chapter II

KINDS OF THINGS

> "The time has come," the Walrus said,
> "To talk of many things ;
> Of shoes—and ships—and sealing-wax—
> Of cabbages—and kings—
> And if the sea is boiling hot—
> And whether pigs have wings."
> —LEWIS CARROLL, *Through the Looking-Glass.*

IT seems as if we are unable to keep away from royalty in our quotations. Probably kings are not of such special importance in the general scheme of the universe as our poetical friends would have us believe ; they merely furnish a convenient rhyme. The interesting fact is that we have already begun to get a basis of division when we know how to distinguish clearly a king from a cabbage. It will be useful for us to linger for a moment and see how we can be quite sure which is which. Both have heads, and both are liable to lose their heads, so that the recital of this, or any other, point of similarity does not assist us very much. Points of dissimilarity, however, are many and obvious, and it is by remembering the most striking of these differences that we may avoid committing the unpardonable blunder of mistaking a king for a cabbage at a social function.

Let us return once more to the motor-car question. How do we recognize one make of car from another ? Each make has certain points in common with all others ; it has four wheels, a chassis, an engine, brakes, a steeringwheel, etc. On the other hand, each make has certain features which distinguish it from all others; the Franklin, for instance, has an air-cooled engine and no water jacket. Some cars do not even have to be seen to be recognized: we can tell them by their rattle. We *describe* any particular type of motor-car by mentioning its peculiar characteristics. We say, in brief, what it *is* and what it *does*.

Just this same plan is followed, somewhat more systematically and with regard to all kinds of things, by the chemist. He distinguishes things by describing each type in detail, mentioning its peculiarities and saying what it is and what it does. It is clearly not so important to emphasize matters which are common to many types, or to describe what a thing does not do. The statement that pigs have four feet is not very much more useful than the statement that pigs have no wings. Many other animals have four feet, but a pig alone has a grunt and gives us bacon.

Since the chemist has a peculiar sort of mind, as we mentioned before, he is far more critical in his definition of a "thing" than is the ordinary mortal. A shoe is not a thing to a chemist, neither is a ship, much less a cabbage or a king. Each of these is altogether too complicated in its structure to be described concisely and accurately, which is what the chemist insists upon doing. An object is not a single " thing " to a chemist unless it is of like material in all its parts, and none of the examples adduced by our friend the Walrus come under this category. The greater portion of most shoes is made of leather, it is true, but the heels may contain rubber or wood, while the stitching, the lining, and the nails all add to the complexity of the assembled article. Cinderella's slipper, if the story-books are to be believed, was made of glass, although the ingenuity of modern research (from which nothing is immune) would try to persuade us that the translator from the original French merely made a mistake and confused *en verre* (glass) with *en vair* (fur)—a material more comfortable for dancing, even if less picturesque. Be that as it may, the fact remains that shoes do differ tremendously in their composition, and so do ships, cabbages, and kings in greater or lesser degrees. There is no standard, in any case, to which all examples must of necessity comply.

Indeed, it is rather a difficult matter to find any kind of thing with which we are familiar in our everyday life that does warrant a perfectly definite description. The trouble is that one kind of thing is so likely to get contaminated with another kind of thing, and while " to the pure all things are pure," to the chemist almost everything is impure. Among the few necessities of common existence which do approach the standard of what the advertisements would call the acme of purity, we may mention water, sugar, and salt. The water which comes to us in our homes to-day has been most rigorously purified, and yet water is so excellent a solvent, it takes up traces of materials with which it comes in contact so easily, that none of us (not even the most ardent prohibitionist) has ever tasted perfectly pure water. We should probably be very much disappointed if we could, since water which has been freshly distilled and thereby deprived of the bulk of its dissolved air is terribly flat and unpalatable. The sugar that we use nowadays is infinitely superior in purity to the crude " molasses " of earlier generations, but it still is not quite what the chemist would call " 100 per cent sucrose." The fact that our salt is liable to get sticky and cake together in damp weather is due to the presence of a minute trace of an impurity—magnesium chloride—which has not been entirely removed from the true salt, or sodium chloride (we shall explain these chemical names later), in the process of manufacture.

This lack of absolute perfection, however, ought not to depress us unduly. After all, from a practical standpoint, each of us is well qualified to give a description of " things " such as water, sugar, and salt which will satisfy all reasonable requirements. We can distinguish water from alcohol without difficulty, just as we can distinguish salt from sugar, by differences in taste. The chemist may prefer to quote other points of diversity—for example, the fact that water freezes in cold weather, while alcohol does not (hence the use of the latter liquid in cheap thermometers); or that a solution of salt in water allows an electric current to pass through it, while a solution of sugar does not; or that a perfect crystal of " rock salt " has no resemblance to a perfect crystal of " rock sugar," the faces being set at different angles. When he makes such distinctions, he is not doing anything mysterious or remarkable. He is merely selecting those characteristics of each " thing " which are most suitable for his purpose in defining it. He is satisfied only when he has discovered enough of such characteristics to be sure that he is not confusing one particular " kind of thing " with any other.

Let us now take up one or two more examples. Suppose that we, being either rich or romantic, have purchased a thing called a diamond. What characteristics does our diamond possess which are common to all things called diamonds, but which are not shared by other kinds of things? Obviously the shape and size are not matters of particular importance, since various specimens are cut in diverse shapes and sizes (*Fig. 1*). Slight differences in colour do exist, it may be admitted, but this is a point of special interest only to the jeweller as affecting the value of a given specimen, and since unusual tints are in all probability due to minute traces of impurities, we do not need to worry ourselves about them here. The really essential marks of a diamond are, first of all, its extreme hardness, and secondly, its extraordinary power of scattering light. The first of these distinguishing features (or *properties*, as the chemist calls them) we can test by attempting to scratch a diamond with other hard objects, or by attempting to scratch other hard objects with a fragment of diamond. We find that the former is impossible, while the latter is done easily. By repeating similar tests with other kinds of things (or *substances*, as the chemist calls them) we can establish a scale of hardness which is headed by the diamond and in which other hard minerals, such as corundum, topaz, quartz, and feldspar, follow. In a similar way it can be shown by experiment that the brilliancy peculiar to the diamond is due to its abnormal ability to bend or refract the rays of light which enter it, and the special shapes

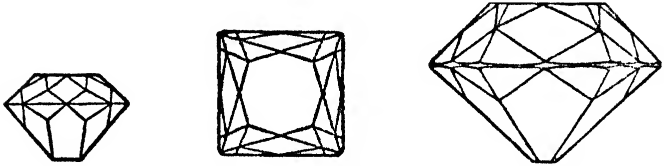

FIG. 1.—Cut Diamonds.

The largest brilliant represents the Regent, or Pitt, diamond, in its original size.

into which diamonds are cut are carefully devised so as to use this property to its limit. Imitations can readily be distinguished from the real article by their inferiority in hardness and "refractive index," and so these properties are of exceptional usefulness in deciding whether a diamond is genuine or not. Other properties of the diamond will be mentioned later.

Gold is another substance or kind of thing which may be discussed while we are in an affluent mood, since it also possesses some very obvious distinguishing features or properties. "All is not gold that glitters," and inasmuch as pure gold is too soft to resist rough usage, most gold coins and articles of jewellery contain other metals (generally silver) mixed with the gold to make them more durable. The proportion of gold in such mixtures, or alloys, is represented in terms of "carats," pure gold being 24 carat, and British sovereigns (which contain one-twelfth their weight of copper) 22 carat. Nuggets of quite pure gold are, however, found in nature (*Fig. 2*). When a chemist examines such a nugget, its size and shape are casual points which he, *as a chemist*, does not trouble to consider. The chemist concentrates his attention upon those properties which are peculiar, or specific, to gold, such as its yellow colour, its resistance to corrosion, its malleability, and its heaviness. Gold leaf can be hammered down to a thickness of less than one-half of a millionth of an inch for the purpose of ornamenting book edges. One of the earliest experiments recorded in history is based upon the high "specific

FIG. 2.—Native Gold.

gravity" (or relative weight) of pure gold as compared with that of a baser alloy. Hiero, King of Syracuse, suspected that a quantity of silver had been substituted into a gold crown which he had ordered to be made for him, and asked his pet inventor, Archimedes, to decide the question. The crown weighed just as much as the original bar of gold from which it was fabricated, but still the king had a feeling that there was something fishy somewhere. Archimedes was puzzled how to proceed until one Saturday night, stepping into a full bath, he observed the resulting overflow of water, and was struck by the idea that, if the crown had been adulterated with a lighter metal such as silver, it ought to displace a larger volume of water than its equal weight of pure gold would do. He was so delighted with his discovery that he ran out into the street exactly as he was, shouting "Eureka! Eureka!" (I have found it! I have found it!). Whether he was arrested for his exuberance history does not relate, but when the actual test was made the crown *did* prove to be partly silver; and the goldsmith, at any rate, was sentenced to a term in prison.

Unfortunately, all minerals are not so simple (or so valuable) as diamond and gold. For example, if we examine a piece of granite, we find that it is spotted, and quite obviously all parts of it are *not* alike in their properties. Now the chemist is not concerned with the source of granite, or with the conditions of its formation in the earth; these are questions for the geologist. Neither is he concerned with the fact that it is very ornamental, and an expensive but durable material for building purposes; these are questions for the constructor. The chemist, looking at a piece of granite very closely, sees at once that it is a mixture of several different kinds of things, and, if a piece of granite is broken up, three distinct substances can be picked out from the fragments. Some of the fragments are very hard and are quite transparent, like glass. Larger masses which have exactly the same properties as these fragments are found in nature under the name of *quartz* (*Fig. 3*). The most interesting property of this substance is its transparency to ultra-violet light, and although pure quartz is very difficultly fusible, windows of "quartz glass" are now constructed for hospital use in the treatment of rickets. Other fragments from the granite are dark in colour and can be split up very easily into shimmering leaves or sheets, thinner than paper. These fragments are identical in all their properties with another substance found in nature called *mica* (*Fig. 4*). Large sheets of mica are employed, on account of their translucency and infusibility, in furnace windows and lamp-chimneys, and also, because of their insulating value, in electrical machinery. Smaller "spangles" are much used for decorative purposes—for example, to produce a frosted effect on toys and stage scenery. Finally, there are fragments in granite which are neither so transparent nor so hard as quartz and which do not split so easily as mica. These are recognized by the chemist to be identical in their properties with a third substance, *feldspar*

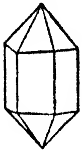

FIG. 3.—Quartz.

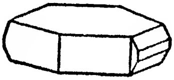

FIG. 4.—Mica.

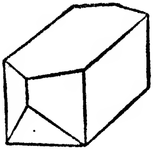

FIG. 5.—Feldspar.

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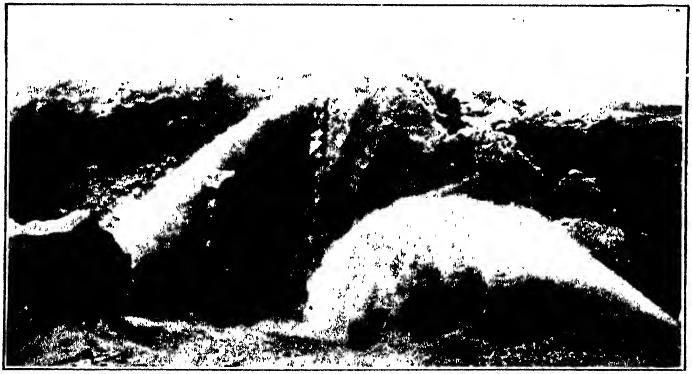
A MICA MINE IN INDIA



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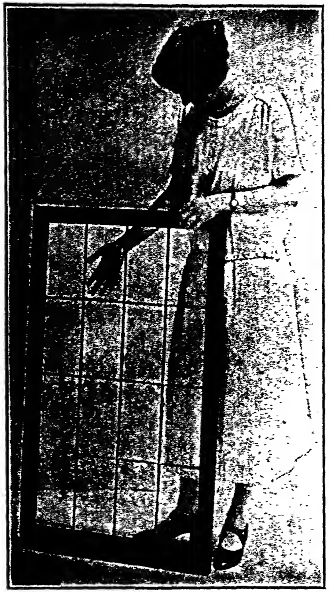
HYDRAULIC GOLD MINING



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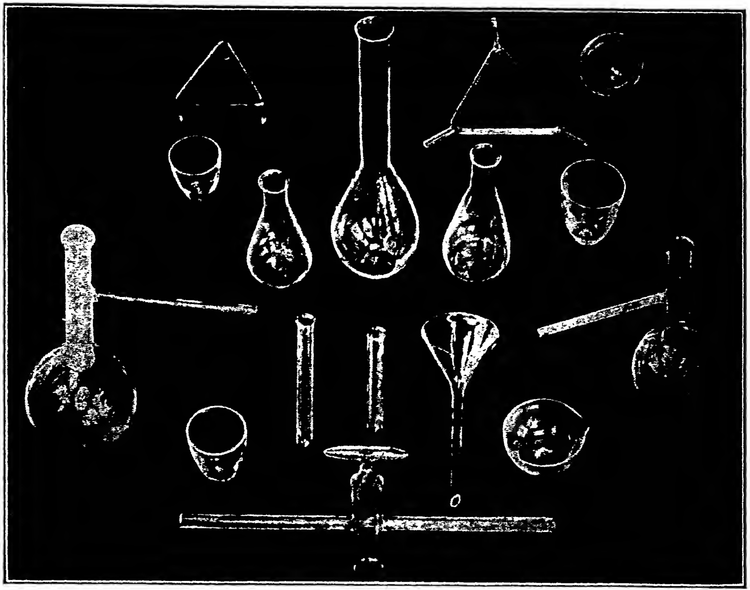
CLEAR FUSED QUARTZ WINDOW



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LABORATORY WARE OF CLEAR FUSED QUARTZ

(*Fig. 5*). Feldspar, unlike quartz and mica, melts very easily, and is consequently very largely used as a glaze for china and porcelain. Granite, then, is not a single kind of thing to the chemist. It is a mixture of three different substances—quartz, mica, and feldspar—and the only way in which we can give a complete description of granite is by stating the properties of these three different kinds of things.

This is, in fact, the first task of the chemist in his profession: to examine whatever samples are brought to his attention and to find out what substances they contain. Thus, when we suspect that a sample of water (or some other liquid) is not safe to drink, we take it to a chemist to be analysed, and after a thorough study of its properties he is able to inform us what impurities are present and what their effect on the human system would be. If we think that we have discovered a gold mine we send specimens of our glittering ore to a chemist, and by a few simple tests he can tell us at once whether it is indeed "native gold" or whether we are being deluded, as so many prospectors have been before us, by a deposit of pyrites, or "fools' gold" (*Fig. 6*).

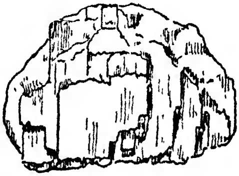

FIG. 6.—Fools' Gold.

While this is all very necessary and important, it does not represent the whole duty of the chemist. If the analysis and description of substances were the beginning and the end of chemistry, we should have a science useful, indeed, but incapable of making any constructive advances, inasmuch as it would merely report upon the material at hand. There are so many different kinds of things in the world that this reporting and tabulation of properties would become a very monotonous and dreary occupation. In actual fact, however, the examination of what different substances will do is far from being dull routine, for the chemist finds that very often, in the course of an investigation, his original materials disappear and he has entirely new substances in their place ! This at once opens up a much more attractive field of work, because these new substances possess, of course, altogether new properties ; and since the usefulness of a substance depends upon its properties and its cost, the production of new substances with desirable properties at a reasonable expense from relatively cheap and useless natural materials is frequently a very profitable undertaking. The various ways in which such changes can be made to occur will be discussed in our next chapter.

Before we leave the present one, a word of warning is in order. Not all of the changes that take place when substances are given the third degree by the chemist are of equal importance. For example, when we take water and cool it until it is completely frozen into ice, or heat it until it has completely evaporated into steam, we have obtained in each case a substance with properties notably different from those of water ; but the chemist does not choose to regard a mere change from the liquid to the solid or to the vapour state as a *chemical* change. He does not consider water, ice, and steam to be three distinct substances, but rather three different physical states of the same substance, and he calls the change from one state to another a *physical* change. This is justifiable as a matter of convenience, since it is so easy to get back from ice or steam to water again merely by melting the ice or condensing the steam ; but it does leave an opening for occasional controversy, as we shall see later.

The temperature at which water changes into ice is called the *freezing-point*. On the thermometer scale commonly employed throughout the British Empire—a scale which was devised by a German named Fahrenheit—this point comes at 32 degrees above zero. The temperature at which water boils is called the *boiling-point*; and on the Fahrenheit scale the boiling-point is at 212°. These two temperatures, however, are so important to the chemist that he prefers to use a simpler scale—the Centigrade or "hundred degree" scale—according to which water freezes at 0° and boils at 100°. All of the temperatures quoted in scientific papers refer to this Centigrade scale. We might give you an algebraic formula by means of which you could convert temperatures from one scale to another, but the diagram on the next page (*Fig. 7*), we believe, will serve your purpose just as well and with the expenditure of less time and trouble. The mercury thread in the thermometer in this diagram stands, as you will observe, at 0° C. (Centigrade) on the vertical scale and at 32° F. (Fahrenheit) on the horizontal scale. Read along the two scales from any other point on the thermometer and you will have equivalent temperatures. Thus 18° C. = 64° F. (room temperature); 98° F. (blood temperature) = 37° C., and so on.

The boiling-point and the freezing-point of a substance are two of its most important properties. Not only do they enable us to recognize a substance, but also to test its purity. Thus sea-water does not freeze until well below 0° C., nor boil until well above 100° C. In accurate work on boiling-points we find it necessary to state the exact pressure under which the measurement is made, for although it requires an enormous change in pressure to alter the freezing-point of a substance one degree, the boiling-point is very susceptible to small variations in pressure. Thus it takes a much longer time to cook eggs or potatoes at the top of a high mountain, since water boils at a much lower temperature under the reduced atmospheric pressure there existent. We measure the atmospheric pressure by means of a barometer (*Fig. 8*), and while the *average* height at which the mercury column stands in this instrument is about 30 inches at sea-level, it varies quite appreciably from day to day. The chemist,

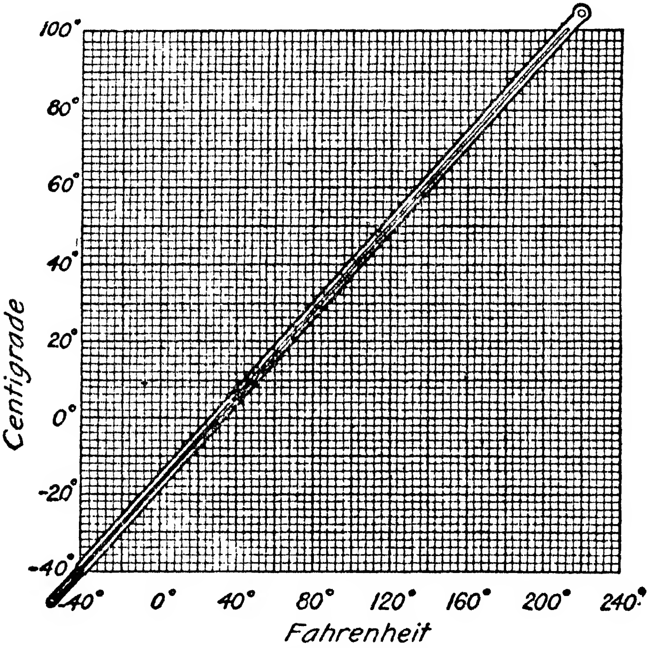

FIG. 7.—Temperatures Centigrade and Fahrenheit.

therefore, when he states that water boils at 100° C., finds it necessary to add "at standard pressure." The pressure chosen, for convenience, is that corresponding to a barometer reading of 760 millimetres.

You will notice that a metric unit, the millimetre, has been selected in preference to the more familiar inch. Why do we use the metric system in chemistry? There are several good and sufficient reasons. In the first place, metric units are international, whereas in our common weights and measures many chances for confusion exist. Thus the British quart differs from the United States quart, while the avoirdupois ounce, the fluid ounce, and the troy ounce are all at variance. Science speaks a universal language, and it is essential that the same statement should convey the same meaning to everybody. In the second place, calculations carried out with metric units are ridiculously simple, since decimal relationships hold throughout. How much easier it is to convert millilitres to litres than cubic inches to cubic feet, or grammes to kilogrammes than ounces to tons! Thirdly, metric units are interrelated. The unit of weight, the gramme, is identical with the weight of the unit volume, one millilitre, of water at the temperature of its maximum density, 4° C.

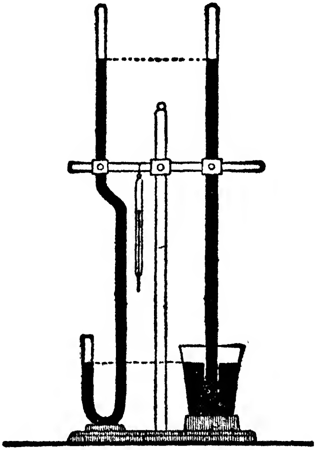

FIG. 8.—Barometer (two forms).

If you are not already familiar with metric units, you had better turn to the Appendix at the end of this book, where the exact conversion factors to our ordinary standards of length, volume, and weight are given. Some day the English-speaking peoples will wake up to the enormity of the handicap they are imposing upon school-children by retaining the casual and cumbersome factors of their antediluvian ancestors, and the employment of decimal units (already used in the United States for coinage) will then be universal. Wireless and sports are steadily doing their bit to familiarize the public at large with the significance of the metre, and even the despised cross-word puzzle has helped a little!

Chapter III

WHAT CAN BE DONE WITH THINGS

> Experiment is the interpreter of nature. Experiments never deceive. It is our judgment which sometimes deceives itself because it expects results which experiment refuses.
>
> —LEONARDO DA VINCI.

BEFORE we describe some typical experiments which illustrate how the chemist proceeds when he wants to produce new substances from old, perhaps we may begin with an example which carries an almost universal appeal, since practically every one has tried his, or her, hand at making toffee at some time or another.

Let us take some sugar in an old porcelain dish—it is not worth while wasting a new one—and heat it in the oven, or better still, so that we may watch it more closely, over a low flame on the kitchen range. We shall observe several changes to occur. First of all, the sugar will melt to a clear, colourless liquid. This is not particularly exciting, nor of any special importance; it is merely a change of state like the melting of ice, and if we are careful we can get back to substantially our original crystalline sugar by immediate cooling. If we continue the heating, however, the liquid begins to darken; it turns first yellow, then brown, and finally black, while vapours billow off which for a time consist almost entirely of steam, and later carry a very disagreeable and persistent odour of "burning." The skilful caramel-maker knows that he must stop before this last stage is reached, but if by accident matters are allowed to go too far, nothing finally remains in the dish but a black mass of charcoal. Leaving out the unpleasant vapours, which result from minor side changes, the main fact to be noted is that, starting with *sugar*, we finish with charcoal or (as the chemist calls it) *carbon*, and with steam, which readily condenses to *water*. In order to avoid wasting space in making statements of this kind, we may write our results *in condensed form*, thus:

$$\text{Sugar} \longrightarrow \text{Carbon} + \text{Water}.$$

The arrow denotes the direction in which the reaction has proceeded. Later on we shall learn how to abbreviate such a statement still further and at the same time add to its significance.

Having broken the ice and acquired a little confidence in our powers of observation, although the rightful occupants of the kitchen may not be so enthusiastic on the subject, we can now proceed to carry out a rather more technical experiment. First of all, let us take some freshly prepared iron filings. We all know the properties of iron in bulk—it is a grey metal, it is attracted by a magnet, and it quickly rusts when left in a damp place. These properties are, of course, not destroyed or changed by its division into very tiny fragments. Secondly, let us take some sulphur, which is ordinarily sold in the chemist's shop as yellow sticks of brimstone or " roll sulphur." Our older readers will recall this substance as forming, in fellowship with treacle, one of the first signs of spring in their youthful days, and also as one of the three ingredients of the mixture called " gunpowder." If we look at the broken surface of a piece of sulphur, we notice that it is composed of shining crystals, and if we examine these crystals very carefully we find that they are of a definite rhombic pattern (*Fig.* 9). Let us now crush some of these crystals into a fine powder and mix them with the iron filings, shaking the two together until they are thoroughly intermingled. From casual inspection it might seem as if we had a new substance, but, just as in the case of granite, it can easily be shown that we have merely a *mixture* of iron and sulphur, and that in this mixture the properties of the two original substances remain unchanged. Thus we can draw a magnet repeatedly over the powder and it will pick up the particles of iron, leaving those of sulphur. This is a very simple way of separating the two ingredients of this particular mixture. Other methods exist; for instance, we might shake up the powder with some liquid, such as carbon disulphide (a substance with a very disagreeable smell, used for killing insects), which dissolves the sulphur and leaves the iron. If now we pour off the liquid into a shallow dish and let it stand for a time, the carbon disulphide quickly evaporates, just as water would under similar conditions, and restores to us our crystals of sulphur in their original form.

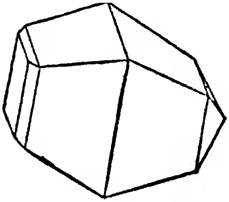

FIG. 9.—Rhombic Sulphur.

The substances which make up any mixture, indeed, are readily separated from one another by taking advantage of suitable differences in properties. But now suppose you go farther in your treatment of this particular mixture. Dump a small heap of it on an iron plate and heat it very strongly over the gas burner, just as you did in the case of the sugar. You will find that the mixture soon begins to glow and to throw off sparks, and when, after it has cooled, you examine the residue, you will discover that it has undergone a very thorough transformation. It is now black and porous; it does not dissolve in carbon disulphide as sulphur does; and it is not attracted by a magnet as is iron. It consists, in fact, of an entirely new substance, which the chemist calls "ferrous sulphide" (Latin, *ferrum*, iron). We write the condensed statement, or equation, for this reaction as follows:

$$\text{Iron} + \text{Sulphur} \longrightarrow \text{Ferrous sulphide.}$$

This is a very important type of chemical change. It is called *combination* or *union*. We start with *two* substances and we finish with *one* substance. Since we cannot destroy or create matter, this new substance must contain both of our original substances, but their original properties are wiped out completely. Because it is the result of a combination, we call the new substance, ferrous sulphide, a *compound*.

This compound, it must be added, differs from a mixture of iron and sulphur in a very significant way, apart from the total change in properties which we have already observed. When we make a mixture of iron and of sulphur, we can use any proportions that we like. We can take a lot of iron and a little sulphur, a lot of sulphur and a little iron, or equal amounts of the two. In forming ferrous sulphide, however, we find that we always use up fixed proportions of iron and of sulphur. In other words, the compound has a perfectly definite composition, which is expressed by the statement that 56 parts by weight of iron combine with 32 parts by weight of sulphur to give 88 parts by weight of ferrous sulphide.[1] If the mixture which you have just massacred happens to contain more of sulphur than is necessary to take care of all the iron present, the extra amount is simply burned away owing to the heat produced in the course of the combination, and you will probably find yourself forbidden to carry out any more experiments at home, since the rest of the family may not appreciate their good fortune in "inhaling the grateful incense arising from your services to the cause of humanity." On the other hand, if your mixture contains an excess of iron beyond the required proportions, this excess simply remains unchanged and can be separated from the ferrous sulphide, after it has been cooled and ground into a powder, by means of a magnet as before.

The fact that has been noted in the preceding paragraph is true not only for ferrous sulphide, but for countless other chemical compounds. Each of such compounds has a definite "composition"; it contains fixed proportions by weight of the substances from which it is formed. When the chemist comes across a general behaviour of this character he gives it proper dignity by calling it a *law*. In this case the law is called the "law of definite proportions," and this law is one of the most important principles in the whole of chemistry. No one chemist can be given the sole credit for its

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[1] "Parts by weight" is not an attractive phrase to repeat continuously, and we promise not to employ it often. If you wish, you may substitute any unit of weight you prefer, such as ounces, or pounds, or tons. Remember, however, that the chemist always uses metric units of weight in his work, and that we shall not refer to any other unit except grammes in this book in the future, unless there is some particular reason for doing so.

discovery, but Joseph Black, professor of chemistry at the University of Edinburgh in the latter half of the eighteenth century, was the first to turn attention to the study of the exact quantities of substances that are concerned in chemical reactions, and it was by his introduction of the use of the balance into experimental work that the development of chemistry into a quantitative science became possible. At the end of the same century a long and bitter battle was waged between Berthollet of Paris, and Proust of Madrid, on the validity of this law of definite proportions. Berthollet said it was not true, Proust said it was, and most rigorous experimental tests were finally held to decide the argument in favour of Proust. Conceding him the victory for the present, we shall reopen hostilities on an entirely new front in a much later chapter.

It is time now to take up a second example of a chemical change, and we shall find it instructive to select the case of the rusting of iron in the air. This is an action which proceeds very slowly under normal conditions, except in the presence of water, but like almost every chemical change it can be made to occur more quickly by means of heat. Iron articles are more subject to corrosion in summer than in winter. The life of an iron boiler is very short unless it is properly protected, and a piece of iron heated in a fire, or on a blacksmith's forge, rusts more rapidly still. The red rust that is formed on the surface of iron at ordinary temperatures, or when it is heated only to a dull red heat, is a very brittle substance; it is not attracted by a magnet; it is much lighter, volume for volume, than iron itself. You may find it difficult to believe, but the red powder, or rouge, with which the modern flapper is accustomed to adorn herself is nothing more or less than very finely divided "iron rust." At higher temperatures, as in the white heat of the blacksmith's furnace, the rust that is formed on iron is of a totally different nature. It is black and scaly, hence it sometimes goes under the name of "hammer-scale"; it is attracted by a magnet; it is identical, indeed, with a mineral found in nature that is called "lodestone" because of its magnetic properties. It also, like the red rust, is much lighter, volume for volume, than iron. Whatever the conditions under which the rusting of iron occurs, however, the resulting increase in volume is quite large, and the actual weight of a completely rusted article considerably exceeds that of the original metal. Evidently, then, the rusting of iron is another instance of *combination*. Where does the additional material come from and what is it?

These are two questions which racked the brains of early chemists most persistently, and it is remarkable what a long time passed before they were answered correctly. For more than a century, indeed, an explanation was accepted which is recognized nowadays as eminently ridiculous. It was thought that when a metal like iron was changed to a rust by heating, something was driven out of it, and this invisible something was christened "phlogiston." When the advocates of the phlogiston theory were challenged to explain why the rust was heavier than the metal, instead of lighter, they put forward the beautiful idea that phlogiston had a negative weight, that is, it weighed less than nothing and actually made a body that contained it tend to float! Chemists are evidently just as prone to error as their brothers in other professions, but they are more fortunate than most men in that their mistakes are always, sooner or later, detected by the eagle eye of the experimenter, whereupon, after due discussion, new opinions are formulated which harmonize more closely with all of the known facts. Without going into further detail regarding the hallucinations of our forefathers (since, no doubt, our delusions will appear equally quaint to our descendants), we shall here describe as simply as possible how the conclusion now regarded as valid—namely, that the extra material comes from the air—has been arrived at. Fortunately, this experiment, unlike our previous efforts, will not offend the nostrils of the community.

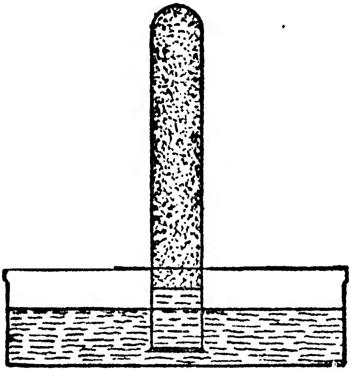

Fig. 10.—Rusting of Iron.

Let us take a glass tube which is closed at one end, moisten its interior surface, and sprinkle into it some iron filings in such a way that the small particles of metal stick to and cover the wet surface of the glass. Then let us invert the tube and set it, mouth downward, in a vessel of water as in the diagram (*Fig. 10*). We have now a definite volume of air confined in the tube together with the finely divided iron. After a time we shall notice that the water is beginning to creep slowly up into the tube, showing that some of the air is being removed as the iron gradually rusts. The water stops, however, when it has gone only about one-fifth of the way up the tube, even although there is plenty of iron which still remains unrusted. However long we wait, nothing more happens. This is rather perplexing, until the notion strikes us that perhaps only a part of the air may be capable of combining with iron to form rust, the residue being quite inactive. If now, acting on this idea, we examine the gas that remains in the tube, we actually do find that it is not the same as ordinary air, since it at once extinguishes a light.

In the process of rusting, then, iron has taken something out of the air, and it is to this something that one valuable property which air possesses—that of allowing things to burn in it—is presumably due, since when it has been removed from air, this property vanishes. It would be interesting if we could isolate this active substance in a pure state and examine it separately from the inactive remainder.

This we can do as follows. Other metals rust in the same way as iron, some more and others less easily. Mercury, the liquid which is commonly used in thermometers, is a metal which does not rust at all at ordinary temperatures, but when it is heated its surface gradually becomes covered with a dark scum. This scum has properties quite different from those of mercury. When collected in quantity, it appears as a bright red powder. If now we put a little of this rust of mercury in a glass tube and heat it strongly over a gas flame, we find that we get a very noteworthy change. The upper or cooler part of the tube becomes covered with a mirror of metallic mercury, and if we take a glowing match and hold it at the mouth of the tube we find that it bursts of its own accord into a bright flame. Evidently the substance which combined with the mercury to form rust is now escaping in the free state in the form of a gas. Later on we shall devote a whole separate chapter to the discussion of this substance and its properties. At this point we shall merely mention that it is, without any shadow of a doubt, the most important of all the substances that exist in the world in which we live, and that its name is "oxygen." Incidentally, it may also be mentioned here that the inactive gas which remains after all the oxygen has been removed from a sample of air by the rusting of a metal in it is a substance called "nitrogen," although small traces of other substances, still more inactive than nitrogen, are also present therein. The properties of nitrogen are clearly quite different from those of oxygen. For example, it does not combine at all under ordinary conditions with metals like iron, and it immediately extinguishes a glowing match that is held in it. Air is substantially a *mixture* of oxygen and nitrogen in the proportions by volume of about one to four, and its properties, like those of all gaseous mixtures, are intermediate between those of its components.

We may now resume our examination of the rusting of iron. This is evidently a reaction which may be expressed in chemical language as follows:

$$\text{Iron} + \text{Oxygen} \longrightarrow \text{Iron rust (or Iron oxide).}$$

There are, as we have already noted, two common kinds of iron rust—red and black. A third kind, even, can be made under special circumstances, which is also black and magnetic, but possesses the peculiar property of catching fire spontaneously when exposed to the air, giving the more familiar red rust as a result of its combustion. Each of these three substances contains iron and oxygen only. The properties of the three

John Dalton

kinds of iron rust, however, are essentially different, and the proportions by weight in which the iron and the oxygen are present in the three kinds are also different. Later on we shall see that this is a very interesting point for further study.

In the same way, the rusting of mercury in chemical terms is written thus :

$$\text{Mercury} + \text{Oxygen} \longrightarrow \text{Mercuric oxide.}$$

In both cases we start with two substances and finish with one substance which is a compound of the original two, so that both of the above equations represent examples of combination.

However, when we heat mercuric oxide and obtain mercury and oxygen from it, we have a reaction of an entirely different type :

$$\text{Mercuric oxide} \longrightarrow \text{Mercury} + \text{Oxygen.}$$

Here we start with *one* substance and finish up with *two* substances. In other words, we have broken down or decomposed a compound into the simple substances which lie concealed within it. This type of change is called *decomposition*. Since the composition of each compound is fixed, it will be clear that we must obtain the products of its decomposition in definite proportions. In this particular example the proportions are 200 grammes of mercury to 16 grammes of oxygen.

Combination and decomposition are the two most useful weapons of the chemist in his experimental work. Other types of chemical change are known, however, and we must be on our guard against making the mistake of assuming that *all* reactions are instances of combination, or decomposition. Before we close the chapter, let us add a few more experiments to our répertoire which will both furnish us with some new and useful information and also serve to illustrate this point.

We give you fair warning not to perform these experiments personally, for motives of economy, since each involves the destruction of a diamond. Unless, therefore, you can find a friend who is willing to loan you a few diamonds on the understanding that he is not going to get them back, you will be well advised to accept our word as to the accuracy of the results obtained. Let us, then, carry out these experiments together merely in imagination, pretending that we are staging an act similar to that presented in 1695 by the court alchemists of the Grand Duke Cosmo III of Tuscany, who really did have " diamonds to burn." Let us also, however, in order to get the most out of the experiments that we can, assume that we have at our disposal modern apparatus such as was not available to his Highness the Grand Duke, in spite of his wealth, and let us also interpret our results in accordance with modern ideas, rather than on the old phlogiston basis.

We first take a diamond and heat it very strongly in the presence of air. The clear crystal soon becomes red hot and then begins to glow and throw off sparks. It gets smaller and smaller in size, and finally disappears entirely. What is the explanation of this mystery? Since the diamond " burned " in the air, a good detective would at once suspect that it had reacted with the active component of the air, oxygen, just as iron does when it burns or rusts. Furthermore, since the product of the reaction was not visible to any of the witnesses, the inference would follow that this product is a gas.

Unfortunately, this gas has been allowed to escape, so that we, in our rôle of Sherlock Holmes, find ourselves temporarily foiled. Let us, however, reconstruct the crime in such a way that we can imprison the reacting substances and the products of the reaction securely, in order to examine the latter at our leisure. This can be done by " electrocuting " a diamond in a closed flask filled with air—or, better still, with oxygen—as shown in the diagram (*Fig. 11*). The passage of a high-voltage current between the two electrodes raises the temperature of the diamond suddenly to the point where it burns brilliantly. When, now, we examine the gas left in the flask, we find that it possesses some properties which are quite distinct from those of oxygen or of nitrogen. For example, it dissolves readily in water, and if we venture to taste the solution, we find that it is faintly and pleasantly acid. The gas that is formed when a diamond burns in oxygen is, indeed, the very same substance which is used in the manufacture of effervescent drinks. It used to be called carbonic acid gas, but chemists now prefer to name it carbon dioxide.

FIG. 11.—Electrocuting a Diamond.

" Why," you will immediately ask, " should we obtain carbon dioxide from a diamond ? Surely there is no connection between a beautiful crystalline diamond and the dirty black material known as carbon or charcoal ! " And yet, if we repeat our last test with a small lump of high-grade charcoal (for example, sugar charcoal, from our very first experiment) instead of a diamond, we find that the products of the two combustions are identical. A piece of graphite—the soft grey material which is used in lead-pencils—might also be substituted for charcoal or diamond, and precisely the same results would follow. Still further, if a chemist were to make a painstaking and careful investigation of the exact quantities of the different substances used and produced in the three parallel experiments, he would report that the same proportions held in each case—namely, that 12 grammes of diamond, graphite, or charcoal combine with 32 grammes of oxygen to form 44 grammes of carbon dioxide.

This second mystery can be cleared up by electrocuting our last diamond. Let us make our final experiment in a closed flask, as before, but let us remove all of the air from the flask by means of a vacuum pump before we throw the switch. We are now heating a diamond in the total absence of oxygen, and we may note in advance that there is no possibility of the formation of carbon dioxide. The diamond is, however, changed ; it is puffed up into a dark grey mass of graphite, the weight of which is exactly the same as that of the original diamond.

This is not a case of combination, neither is it a case of decomposition. We start with *one* substance, diamond, and we finish with *one* substance, graphite. The change that we have observed is similar in this respect to the familiar change that occurs when water freezes to ice. Are diamond, graphite, and charcoal, then, merely different physical states of the same substance, carbon ? That is, indeed, the point of view that the chemist, *as a chemist*, is forced to hold. It is true that the physical characteristics of the three materials are extraördinarily different; when we describe what they are and how they behave alone, the three stories are quite distinct. But as soon as we go over to the study of chemical behaviour, as soon as we analyse the reactions which they undergo with other substances and the products of these reactions, we find that they are identical.

It appears, consequently, that there are two sorts of properties, *physical* and *chemical*, and that our original idea that each " kind of thing," or substance, must possess perfectly definite distinguishing features, or properties, already needs to be recast if we are to be consistent. We can make the necessary apology, for the present, in the following way. If two specimens of matter, each alike in all its parts, are identical in both physical and chemical properties, then they are the same substance. If two specimens differ merely in physical properties, they represent different states of the same substance. Only if they differ in their chemical properties as well are they *fundamentally* different.

We are still not quite out of the wood, however. Although the conversion of diamond to graphite bears certain resemblances to the change of water to ice or steam, yet there are points of dissimilarity which must not be left unmentioned. In the first place, the freezing of water into ice, and its evaporation into steam, are reactions which, according to our everyday experience, are readily reversible. But the change of diamond to graphite or charcoal is not, as the high price of diamonds testifies, a reaction which can be made to occur in the opposite direction on a commercial scale. It is true that a French chemist, Henri Moissan, did discover about thirty years ago that small hard crystals which seemed to exhibit the physical properties of diamonds could be obtained from charcoal under special circumstances. Moissan heated pure sugar charcoal and pure iron together to as high a temperature as possible in an electric furnace and found that the molten iron dissolved the charcoal just as readily as water dissolves sugar or salt. The crucible containing the white-hot liquid was then plunged suddenly into cold water. The outer layers of the molten mass immediately solidified and brought an enormous pressure to bear upon the contents of the inner core. The carbon was thus forced to come out from its solution in iron under a terrific stress, and when the metal was later dissolved away by means of acid, microscopic crystals remained which were claimed to be artificial or " synthetic " diamonds. Recently, however, some doubt has been thrown upon the question whether these crystals, partly transparent and partly black, obtained by Moissan, are true diamonds after all. Whatever the final decision on the matter may be, it is certain that synthetic diamonds cannot be expected to compete in the near future with the natural product. Under what conditions the diamond is formed in nature still remains nature's secret.

A second point of dissimilarity between the change from diamond to graphite and the change from water to ice is that, in the former case, both materials are solid, while the freezing of water involves an obvious " change of state." This, however, is not a serious difficulty to the chemist. He considers a difference in crystalline form, such as we have in the cases of diamond and graphite, just as important as the difference between the solid and the liquid states. He calls the two crystalline varieties *allotropic modifications*, which is only a long-winded way of expressing the fact that they *are* essentially different in form. When we come to look into the matter closely, we find that a great many substances exist in several distinct allotropic modifications. You can easily prove it for yourself in the case of sulphur. Take some of the rhombic crystals of "roll sulphur" which you used in an earlier experiment in a small porcelain dish and melt them carefully by heating the dish over a very low gas flame. Be careful not to let the sulphur catch fire, and as soon as the crystals have all melted put out the gas and let the limpid yellow liquid cool. Crystals soon begin to form again on the surface. When these have collected into a fairly thick coherent crust, puncture this crust in two or three places with an old knife, pour out the sulphur from the still fluid interior, and then cut the crust away. You will find that the walls of the dish and the under surface of the crust are covered with beautiful long needle-shaped crystals (*Fig. 12*), quite different in their physical characteristics from the original rhombic ones. This kind of sulphur is called *monoclinic sulphur*. In a day or two's time, if you look at the dish again, you will notice that the crystals have become opaque and brittle; they crumble readily into powder; and examination under the microscope will show that they have reverted into tiny specks of rhombic sulphur. A more critical examination of the two forms of sulphur crystals at different temperatures would disclose the interesting fact that *above* 96° C. rhombic sulphur is unstable and gradually changes to monoclinic, while *below* 96° C. monoclinic

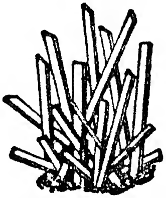

FIG. 12.—Monoclinic Sulphur.

sulphur is unstable and slowly goes over to rhombic. The temperature at which the order of stability is reversed is called the *transition-point.* Its resemblance in character to the freezing-point or boiling-point of a liquid serves to confirm us in our opinion that the change from one allotropic form to another represents merely a physical change.

Ice, even, exists in several "allotropic modifications," although we encounter only one crystalline form under ordinary conditions. When water is frozen under a tremendous pressure, however, no fewer than four other forms of ice may be isolated. All of these four forms differ from ordinary ice in one remarkable respect—instead of being lighter than water, they are heavier. It is certainly fortunate for skaters as well as for fish that the conditions under which they are stable do not exist in nature!

We are now ready to return, after this necessary digression, to the study of things in general.

Chapter IV

GETTING DOWN TO THE REAL NATURE OF THINGS

> Methinks the Chymists, in their searches after truth, are not unlike the navigators of *Solomon's Tarshish* fleet, who brought home from their long and perilous voyages, not only gold, and silver, and ivory, but apes and peacocks too : for so the writings of several (for I say not all) of your hermetick philosophers present us, together with divers substantial and noble experiments, theories, which either like peacock's feathers make a great show, but are neither solid nor useful, or else like apes, if they have some appearance of being rational, are blemished with some absurdity or other, that, when they are attentively considered, make them appear ridiculous.
>
> —ROBERT BOYLE, *The Sceptical Chymist.*

EXPERIMENTS such as we have described in the last chapter have been carried out by chemists for centuries. Countless mixtures of substances have been made and submitted to all possible changes of conditions in the attempt to make the original materials combine and furnish new substances with novel and useful properties. Many of these experiments, of course, have not given the desired results, but still the number of distinct substances which the chemist recognizes increases rapidly year by year. At present the chemical dictionaries list in all approximately 300,000 different " things." It would seem as if our science was destined to become hopelessly congested and confusing.

Fortunately, chemists have also been working just as strenuously in the reverse direction. They have performed innumerable experiments in which single substances have been taken and maltreated in all possible ways in order to induce them to decompose. In the majority of cases these attempts have, sooner or later, been crowned with success, but there are some substances which no chemist has ever managed to disintegrate even under the most intense heat or by the aid of the most powerful electric current. Gradually the idea has developed, in consequence, that these substances are *incapable of decomposition.*

Mercuric oxide, as we have seen, is split up by heat into mercury and oxygen. Nobody, however, has ever found it possible to decompose mercury or oxygen into simpler substances. Nobody has ever yet succeeded in getting anything else out of iron, or sulphur, or gold; neither has it been found possible to build up, or synthesize, any of these by the combination of other substances. There are, in all, only about ninety different cases of this type, and the chemist distinguishes these ninety substances by calling them simple substances or *elements*. All compound substances are known to be built up out of these ninety, and can be decomposed into them again.

The belief that the fundamental forms of matter are limited in number was prevalent long before any experimental facts in support of the idea were obtained, and a short retrospect into history may help us to understand and appreciate the present situation. The ancient Greek and Indian philosophers cherished the opinion that the earth and all that it contained were composed of one basic material which assumed different forms in different substances. From this elementary material everything was derived, and if we but knew how to go about it we could convert everything back into this same " element." Unfortunately, the philosophers did not trouble to make any experiments in this direction —that was beneath their dignity—and they could not agree in their speculations as to what the basic element was. Some said that everything started with *water* and ultimately returned to water ; some insisted that everything was formed from *air*. Others favoured *fire* as the initial and final material ; still others considered that *earth* was the beginning and end of all things.

Since they were unable to settle their quarrels by argument, a compromise was finally reached whereby it was postulated that there were *four* elements—fire, water, earth, and air—and that all matter was built up out of these four elements, assembled together in different proportions. This theory was favoured by Aristotle and survived until the middle of the seventeenth century. A typical argument by a zealous advocate of the school of Aristotle may be quoted :

> If you will but consider a piece of green wood burning in a chimney, you will readily discern in the disbanded parts of it the four elements. . . . The fire discovers itself in the flame by its own light ; the smoke by ascending to the top of the chimney, and then readily vanishing into air, like a river losing itself in the sea, sufficiently manifests to what element it belongs and gladly returns. The water in its own form boiling and hissing at the ends of the burning wood betrays itself to more than one of our senses ; and the ashes by their weight, their fineness, and their dryness, put it past doubt that they belong to the element of earth.

It was not often that the philosophers would condescend to carry out even so simple an experiment as that which has just been described. A class of men arose, however, during the Middle Ages, who took more interest in laboratory work. These " alchemists " were stimulated by the belief that, if the same elements were indeed present in all substances, then it ought to be possible to change one substance into another by rearranging the elements in the right proportions. The change in which they were most interested, of course, was the transmutation of the baser metals into gold. In their attempts to achieve this transmutation they failed to get anywhere with the fire-water-earth-and-air idea, and accordingly abandoned it in favour of a theory that all metals were composed of three " principles "—mercury, sulphur, and salt. These three principles they tried to reassemble in their furnaces in the way necessary to create gold.

Again they were unsuccessful in their quest, and yet some of them did good service, inasmuch as they substituted experiment for philosophy. Their superiority over the supporters of the theory of Aristotle may be indicated by an extract from the works of Paracelsus, the most notorious of all alchemists. He says of his disciples :

> They are not given to idleness, nor go in a proud habit, or plush and velvet garments, . . . but diligently follow their labours, sweating whole days and nights by their furnaces. . . . They put their fingers among coals, into clay and filth, not into gold rings. They are sooty and black, like smiths and miners, and do not pride themselves upon clean and beautiful faces.

As time went on, however, and still no gold rewarded their efforts, the alchemists also sank back into speculation. They began to have visions of their three " principles " as representing the essence or soul of things rather than anything material. They became obsessed with the fantastic delusion that they might succeed in side-stepping all their experimental difficulties by the intervention of a magical "philosopher's stone," a trace of which would transmute a large mass of baser metal into gold, or through the discovery of an "elixir of life" or a "universal solvent." Apparently none of them ever thought of the fact that a universal solvent would prove to be a white elephant if it did chance to be discovered, since it would clearly be impossible to devise a vessel in which to keep it. Encouraged by the avarice of kings and nobles who supported their quests, the alchemists were often, it must be confessed, guilty of flagrant dishonesty and fraud, always holding out the hope that the next experiment must surely be successful if only the necessary funds for its performance were provided. When powerful people were imposed upon too badly in this way, naturally they became seriously annoyed, and it is on record that the Bishop of Würzburg used to maintain a special gallows in his diocese upon which to hang chemists. There are some people, no doubt, who would be glad to see this pleasant old custom revived.

It was at this stage that a reformer arrived on the scene in the person of the Honourable Robert Boyle. Robert Boyle—"the father of chemistry and the brother of the Earl of Cork"—was born in Ireland in 1627. He spent most of his life in the experimental study of various branches of science, but also indulged in meditations in theological fields.[1] He was one of the founders

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[1] One of Dean Swift's most savage satires bears the title *Pious Meditations on a Broomstick in the style of the Honourable Mr. Boyle.*

of the Royal Society, and afterward its president. In his famous treatise *The Sceptical Chymist*, published in 1661, Boyle accused both the " hermetick philosophers," who put all their trust in the four elements of Aristotle, and the " vulgar spagyrists," who based their faith on the three principles, of " playing with names at pleasure," sagaciously adding : " They write darkly, not because they think their notions too precious to be explained, but because they fear that if they were explained men would discern that they are far from being precious." After criticizing the loose methods of inquiry then in vogue, Boyle proceeded to examine carefully certain simple experiments in order to show that the results to which they led were not in accordance with the standard beliefs, and then went on to give his own definition of the term element. This may be reproduced in condensed form, leaving out certain repetitious matter, as follows :

> I must not look upon any body as a true principle or element, which is resoluble into any number of distinct substances, how small soever. I mean by elements certain primitive and simple, or perfectly unmingled bodies ; which not being made of any other bodies, or of one another, are the ingredients of which all those called perfectly mixt bodies are immediately compounded, and into which they are ultimately resolved.

While this is all very clear and logical as far as it goes, Boyle was unable to develop any satisfactory experimental method of testing his definition. In fact, Boyle himself was not entirely exempt from some of the errors of his period. It is true that Boyle knew that the rust of a metal was heavier than the original metal from which it had been obtained by the action of heat, but he thought that the increase in weight was due to the fact that particles of heat had permeated the metal. He even wrote a book to prove that heat had weight! Until the phlogiston theory, which reigned supreme for nearly a century after Boyle, was finally overthrown by the famous French chemist Lavoisier, who showed with the help of the balance the true nature of combustion and the rôle which oxygen played in rusting, no real progress was possible. After Lavoisier had studied in detail, however, a number of simple examples of combination and of decomposition, such as we have already discussed in Chapter III, he was in a position to propose an *experimental* definition of the term element, and this he put in practically the same words as those given at the beginning of this chapter.

There is one obvious point of weakness in this definition of Lavoisier. It postulates that a substance is an element if it has "*not yet*" been decomposed into, or formed from, other substances. Now chemists are always making improvements in their experiments and are steadily becoming more artful in their methods of attack. Consequently, it soon became evident that certain substances which were regarded by Lavoisier as elements were, after all, only compounds. For example, ordinary quicklime was thought to be an element until 1808, since it could not be broken up even by the most intense heat, but its compound nature was clearly demonstrated in that year by Sir Humphry Davy. Davy passed an electric current through lime, and through substances derived from lime, and thereby obtained a new metal, which he called calcium. This metal tarnishes very readily when exposed to air, and Davy was soon able to show that quicklime was the rust, or oxide, of calcium. Conversely, the yellowish-green gas chlorine, which was first made by other methods but is now most conveniently manufactured by passing an electric current through a solution of common salt, was for a long time erroneously thought to be a compound, whereas we now recognize it to be a true element. Mistakes of this kind, however, can no longer recur, since, as will appear in later chapters, our command over the situation has recently become so complete that we can now line up the elements like a company of soldiers on parade, number them off from right to left, and be certain from their responses that all places in the ranks are properly filled.

Such an absolute mastery of mind over matter, of course, has been established only gradually, and represents the net result of many successive and partial advances. We shall make it our duty to outline the essential features of each of these in turn as opportunity arises. For the present moment we shall be satisfied with the definition of an element as given by Lavoisier. Later on we shall be in a better position to demonstrate how subsequent discoveries have made it necessary for us to modify and reword this definition; but it will suffice for our immediate purposes.

The ninety known elements are by no means all equally abundant. Indeed, only a very few of them are ever found in nature in the simple, or free, state. The majority occur entirely in the form of compounds with other elements and must be released therefrom by the discovery of appropriate reactions. The total number of compounds that can be built up out of the ninety elements, and decomposed back into them again, is, as we have already pointed out, practically unlimited. Dozens of different substances are known which contain gold, hundreds which contain iron, thousands which contain sulphur, and tens of thousands which contain oxygen and carbon.

The total plentifulness (free and combined) of the various elements in this world, at least as far as can be ascertained by a study of the atmosphere, all lakes and oceans, and the available part of the earth's crust, has recently been estimated by F. W. Clarke and H. S. Washington of the United States Geological Survey. Their results show that the first twelve elements account for more than 99 per cent, the order being as follows :

Oxygen	49·2	Potassium	2·4
Silicon	25·7	Magnesium. . . .	1·9
Aluminium	7·4	Hydrogen	0·9
Iron	4·7	Titanium	0·6
Calcium.	3·4	Chlorine	0·2
Sodium	2·6	Phosphorus . . .	0·1

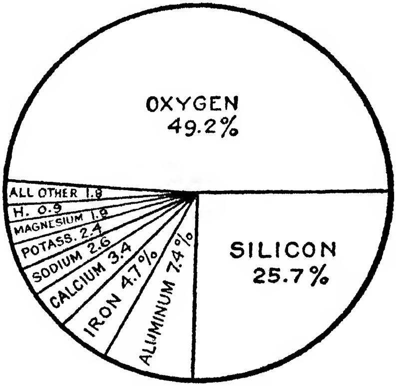


FIG. 13.—Composition of the Earth's Crust.

The accompanying diagram (*Fig. 13*) will convey the significance of this table to you in a more striking manner. It will be noted that *oxygen* is by far the most abundant of all the ninety elements in this world, constituting, in point of fact, nearly one-half of its whole obtainable mass. The air, as we have seen, is approximately one-fifth free oxygen; water is a compound which is eight-ninths oxygen; and oxygen is also a constituent of most common rocks.

The second most plentiful element is *silicon*, which does not occur free in nature, but which in its combinations makes up more than one-quarter of the available material on this earth. The compound which oxygen forms with silicon is known in a pure state as quartz, and in a less pure state as sand. More complex compounds, such as feldspar and clay, are very numerous. Feldspar and clay also contain *aluminium*, the third element on our list, a metal which is in common household use nowadays, although less than fifty years ago it was practically unknown. The "true story" of aluminium is a romance that will make you hold your breath in a later chapter.

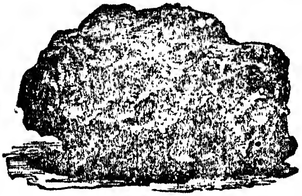


FIG. 14.—Meteorite.

Iron, of course, is familiar to everybody, although pure iron is too soft for most purposes, and all common "iron" and "steel" articles contain small amounts of other elements. Iron occurs free in meteorites (*Fig. 14*), but its chief ores are the rusts, or oxides, of the metal. *Calcium* we have already mentioned as an element discovered by Davy in quicklime. Chalk, limestone, marble, and gypsum are common substances containing calcium.

Sodium, *potassium*, and *magnesium* are three other metals which were first isolated by the genius of Davy. They are all far too reactive with other substances to be found free in nature, but compounds of each are quite abundant. Ordinary table salt is a compound of sodium and chlorine, hence its chemical name of sodium chloride. Potassium is a constituent of feldspar and of mica. Magnesium compounds occur in sea-water and in many common rocks.

Hydrogen, the gas which is used in filling balloons because of its extreme lightness, owes its position of ninth on the list to the fact that one-ninth of water is hydrogen. The completeness of the transformation which elements undergo when they unite together to form compounds is wonderfully illustrated when we compare the properties of water with those of hydrogen and oxygen, or the properties of common salt with those of sodium (a shining metal) and chlorine (a poisonous gas).

Titanium is a minor constituent of sand, clay, and granite rocks. *Chlorine* is the second (and better) half of sodium chloride, and *phosphorus* is found in all fertile soils in the form of phosphates.

The remaining elements, according to the figures of Clarke and Washington, are *relatively* rare. In actual fact, however, many of them are much better known to the man in the street than are the majority of those on our select list. In case you should happen to feel at all disappointed over the omission of any of your own favourites, such as copper, silver, and gold, from the roll of honour, it might be well for us to mention that even nitrogen (which makes up almost four-fifths of the atmosphere) and carbon (which enters into our daily life more intimately than any other element, being a constituent of every conceivable article of food, clothing, and fuel) likewise fail to qualify in the first twelve.

A complete list of the elements which have been discovered up to 1928 will be found in the Appendix at the end of this book. A great many of the substances on this list are so scarce, or the uses to which they have been put are so trivial, that we shall not need to consider them in any detail; but it is necessary for you to have a complete "telephone directory" of the elements at hand for reference purposes, as you will see when you begin to get better acquainted with the more important and interesting members of the "Ninety Club."

This brings us back to the analogy which we introduced into the very first chapter of this volume. We have now succeeded, as you will admit, in reducing the innumerable material objects which surround us to ninety "roots of things" or elements, and the number of these fundamentally different elements is comparable with the number of different makes of motor-cars. Furthermore, we have found that some of the ninety elements are much more plentiful than others, just as certain makes of cars (no names will be mentioned!) are more popular than others.

We might very profitably continue this analogy through subsequent chapters. For example, we shall find in a little while that it is possible to classify our ninety elements into distinct groups, just as we classify automobiles into four-cylinder, six-cylinder, and eight-cylinder cars. All analogies are imperfect when pressed too far, however, and we believe that we can find a better way of securing your interest in our chemical elements than by means of motor-cars. After all, a motor-car is a mere machine, and what we need in chemistry is a more personal and human touch. We shall therefore make the attempt to correlate each element that we examine in special detail with a different set of individuals, striving our utmost to keep our comments good-natured and to avoid rousing any ill feelings. Please be lenient with us in our efforts, even if we do seem sometimes to be lacking in proper respect. Chemistry can be made a very arduous and uninteresting subject, and we are merely doing our best to help you over the rougher and less attractive parts of the road that you have to travel, without tiring you too sorely.

For the present, we shall restrict ourselves to calling your attention to the most obvious division of our elements into two kinds, corresponding with the most obvious division of human beings into two kinds—male and female. In the case of the elements the two kinds are named *metals* and *non-metals*. While the reader might find it difficult to give a satisfactory scientific definition of the word " metal," yet the term has a plain practical meaning to everybody. Iron, copper, and aluminium are unmistakable members of the class of metals—strong, hard-working men, all of them. Carbon, sulphur, and oxygen are just as clearly non-metals—fragile and ethereal ladies. We shall find later that although the chemist employs other means for drawing a distinction between metals and non-metals than mere physical appearance, or the possession of " metallic lustre," yet, after all, his decision coincides almost invariably with common usage.

Before you can expect to get on a friendly footing, however, with the various men and women to whom you will be presented as you pass around among the chemical elements, it is necessary for you to be initiated into a method of " sizing them up " properly.

Chapter V

A FEW FIGURES

"What's twice eleven?" I said to Pooh.
("Twice what?" said Pooh to Me.)
"I *think* it ought to be twenty-two."
"Just what I think myself," said Pooh.
"It wasn't an easy sum to do,
But that's what it is," said Pooh, said he.
"That's what it is," said Pooh.
—A. A. Milne, *Now We Are Six.*

We have seen that chemistry made no real progress so long as it was confined to speculations and to haphazard experiments. It was not until Black and Lavoisier had succeeded in convincing their colleagues that it was essential to look very carefully into the *quantities* of substances that are concerned in chemical reactions in order to interpret the results of these reactions correctly that things began to move forward. Chemistry now counts as one of the exact sciences; it is on a 100% quantitative basis.

This change in view-point, it must be granted, has had one unfortunate result—namely, it has made chemistry less appealing to the outsider. There is a widespread impression, not confined to the class-room, that chemistry is a very difficult subject indeed and that it is necessary to absorb a considerable amount of mathematics before daring to approach it with any degree of confidence. Now, most people have a healthy instinctive dislike for advanced mathematics. It is with pleasure, therefore, that we make the assurance here that nothing more will be required of the reader of this volume in the way of figures than a few applications of simple arithmetic. It is true that if one wishes to wander into the unexplored areas of chemistry and to extend its boundaries, then such luxuries as the integral calculus and differential equations must first of all be laboriously acquired. In this book, however, we shall not need any equipment more elaborate than that demanded in solving the problem of " What's twice eleven ? " at the beginning of the chapter, and, if you will promise to give us the same cordial co-operation which *Pooh* gave to *Christopher Robin* in the solution of that problem, I am sure that we shall have no difficulty in arriving at correct answers.

Now let us go back to our elements and to the compounds which they form. The question arises : " In what fashion do the several elements which make up any particular compound exist therein ? " The elements are certainly there, since the weight of the compound is exactly equal to the weight of the elements from which it was made. The properties of these elements, however, are entirely eradicated in favour of the new properties of the compound. *What happens when elements combine ?*

The key to this riddle begins to appear when we remember the fact that the proportions by weight of the elements in each compound are fixed. For example, all ferrous sulphide contains iron and sulphur in the proportions of 56 grammes of iron to 32 grammes of sulphur. Every sample of water contains hydrogen and oxygen in the ratio by weight of one to eight. This shows that there must be some definite plan into which the iron and the sulphur, or the hydrogen and the oxygen, enter when combination takes place between them. Evidently the elements are not at liberty to choose any casual proportions ; the amount of one which can combine with a fixed amount of another is determined in advance.

Let us watch the chemist while he carries out a typical combination in his laboratory in order that we may obtain a clearer idea of what he does and what his results signify. The experiment which we shall select is one with which you are already familiar—namely, the burning of carbon in air to form carbon dioxide.

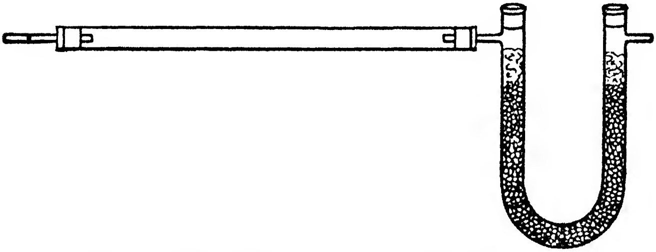


FIG. 15.—Measurement of Combining Proportions.

The product of the reaction in this case, as you will remember, is a gas, and the chemist, therefore, has to exercise some ingenuity in devising a method for capturing this gas without loss in a weighable form. The apparatus which he uses is shown in part in the accompanying diagram (*Fig. 15*). A weighed amount of pure, dry sugar charcoal is taken in a small porcelain boat and set in the middle of a long tube made of hard glass. A slow current of air, which has been carefully dried and purified by passing through suitable chemicals, enters this tube from the left. On the right there is attached a U-shaped tube which contains small lumps of soda-lime, a mixture of which the active component is quicklime. This U-tube is also carefully weighed before the experiment is started.

The chemist proceeds to heat the charcoal in the boat by means of a gas burner. The charcoal soon begins to glow and finally burns up completely. The carbon dioxide formed by the combustion passes through the soda-lime and is entirely absorbed thereby. The weight of this carbon dioxide, therefore, can readily be obtained at the end of the experiment by weighing the U-tube again and subtracting the original weight from the final one.

Now we are ready to start our calculations. Let us suppose for simplicity that the chemist began with exactly three grammes of carbon in his boat. If he has been careful in his technique, he will find that the increase in weight of the U-tube is very close to eleven grammes. To avoid trouble let us agree to call it exactly eleven grammes. We do not want to be worried with decimals in our first trial.

The situation, then, may be summed up as follows: Three grammes of carbon have combined with an unknown quantity of oxygen to form eleven grammes of carbon dioxide. Now, in all the chemical reactions that have ever been carried out, nobody has ever been able to establish any real difference in weight between the substances existent at the beginning of the experiment and those existent at the end of the experiment. This general fact is called the "law of the conservation of mass." Employing this law in the present case, we do not require a profound intellect to deduce that the unknown weight of oxygen is eight grammes.

So far so good. But another chemist who did this same experiment would probably not happen to start with exactly three grammes of charcoal ; he might use a larger or a smaller quantity. How are we going to express the results of our experiment in a way which will show at a glance whether they agree with those which he obtains ? We can do this very easily by converting our figures into *percentages*.

Percentages simply show the number of parts in one hundred. In our experiment we found that three grammes in eleven of carbon dioxide come from the carbon and that the remaining eight grammes come from the oxygen. The percentage of carbon in carbon dioxide is therefore three-elevenths of one hundred, or 27·3. The percentage of oxygen is eight-elevenths of one hundred, or 72·7. Whatever amount of carbon we start with, we shall obtain these percentages, within the limits of our experimental error.

There is, however, a second compound of carbon and oxygen which has entirely different properties and an entirely different composition. This substance is known as carbon monoxide. It is formed whenever carbon, or a substance containing carbon, burns at a high temperature in the presence of a limited supply of oxygen. For example, it is present in considerable quantity in the gas that issues from the exhaust of a motor-car, and, since one of the properties of carbon monoxide is that it is intensely poisonous, fatal accidents frequently occur through careless people leaving an engine running in a closed garage. We cannot go through all the details of the preparation and analysis of pure carbon monoxide here. It will suffice to state that its percentage composition has been shown to be : carbon 42·9 ; oxygen, 57·1.

When you compare these numbers with those obtained for carbon dioxide, you will be excused for not seeing any connection between the two. All that the figures appear to convey is the dismal fact that for some obscure reason carbon insists on combining with oxygen in two unrelated proportions to form two different compounds. But are the proportions unrelated ? If we were to start with the same amount of carbon in each case, what would be the amounts of oxygen required ? In order to answer this question, we must first of all decide what fixed weight of carbon we will choose. To save you from hesitating in making a choice, we shall suppose that you have picked upon twelve grammes of carbon. The reason for selecting this particular figure will reveal itself later.

Now let us see how it works out. In our own experiment on carbon dioxide, 3 grammes of carbon combined with 8 grammes of oxygen, therefore 12 grammes of carbon would require how many grammes of oxygen ? We do not expect to wait long before hearing your confident answer: 32 grammes.

In the case of carbon monoxide, the percentage composition is all that we have at our disposal at the present moment. This tells us that 42·9 parts by weight of carbon have combined with 57·1 parts by weight of oxygen, and if you are energetic enough to work out the proportion you will find that this means that 12 grammes of carbon would demand 57·1 × 12 ÷ 42·9 grammes of oxygen, and the final result of this calculation is 16 grammes.

Here is, then, a simple and quite unexpected relation between the two compounds of carbon and oxygen. The weight of oxygen combined with a fixed quantity of carbon in one compound is *exactly double* the weight of oxygen required by the same quantity of carbon in the other compound.

This is not an isolated instance. It represents the general behaviour of all cases of a similar type. The law which the chemist draws up to express this fact is called the " law of multiple proportions." It states that whenever two elements, A and B, combine in more proportions than one, the quantities of A which combine with a fixed quantity of B are connected by a simple multiple. In the case that we have looked into, this multiple was two ($2 \times 16 = 32$). The multiple, of course, is not always two, *but it is always simple.*

This law of multiple proportions was first enunciated by John Dalton in 1804. Dalton was born of very humble Quaker parentage in the county of Cumberland in 1766. He began to teach in the village school at the age of twelve, he continued to teach at the neighbouring town of Kendal as he grew older, and he earned his living as a schoolmaster at Manchester up to the very close of his active career. There is a story which relates that a well-known Parisian savant, M. Pelletier, visited Manchester in 1826 with the express purpose of calling upon the famous Dalton. He found his lodgings only after very diligent search, and discovered the illustrious chemist in a tiny room on a back street, looking over the shoulders of a small boy who was working out a problem on a slate. " Est-ce que j'ai l'honneur de m'addresser à Monsieur Dalton ? " asked the visitor. Dalton was not perceptibly embarrassed at being spoken to in French, but the stranger *was* astonished at his reply. " Aye ! " said the stolid Quaker. " Wilt thou sit down whilst I put this lad right about his arithmetic ? " It is not improbable that this prolonged contact with the simple facts of addition, subtraction, multiplication, and division empowered Dalton to detect the simple idea expressed by the law of multiple proportions which so strangely escaped the notice of his more prosperous contemporaries. Proust worked for years on percentage compositions in a magnificent laboratory where almost every vessel was made of platinum. Dalton performed only a few rough experiments, mainly with home-made apparatus, but Dalton achieved the discovery which Proust had at his finger-tips yet failed to grasp.

A few additional examples of the law of multiple proportions may now be considered. There are two compounds of oxygen and hydrogen. One of these compounds is very well known to us indeed: it is water. The other is a substance familiar to synthetic blondes in times of trouble: it is named hydrogen peroxide. We shall not go into the methods by which the chemist arrives at the exact composition by weight of each of these substances, since the determination is not so easily performed as in the case of carbon dioxide, and accurate results demand extreme care and skill. However, the percentage compositions obtained when the most scrupulous precautions are taken run thus:

	Oxygen.	Hydrogen.
Water	88·81	11·19
Hydrogen peroxide . .	94·07	5·93

Once again these percentages, on the surface, seem quite unrelated, but if we calculate the weight of hydrogen required to combine with 16 grammes of oxygen in the two compounds, we find that the result comes as 2·016 grammes in the case of water and 1·008 grammes in the case of hydrogen peroxide. The ratio of these weights is, as before, exactly two to one.

Two compounds of iron and sulphur have been previously mentioned. One of them, ferrous sulphide, obtained by the direct combination of iron and sulphur, contains the two elements in the proportions of 56 grammes of iron to 32 of sulphur. The second, iron pyrites or "fools' gold," gives on analysis a percentage composition which is equivalent to 56 grammes of iron to 64 grammes of sulphur. Once more we note the simple ratio of two to one.

The three rusts or oxides of iron represent a rather more complicated case. The uncommon black oxide which catches fire of its own accord when exposed to air (or is "pyrophoric," as the chemist calls it) contains 56 grammes of iron to 16 of oxygen. The red rust, obtained when iron is heated in air at a fairly low temperature, contains 37·3 grammes of iron to 16 grammes of oxygen. The black rust obtained at higher temperatures shows on analysis 42 grammes of iron to 16 grammes of oxygen. A little thought will be necessary before it becomes apparent that these ratios may be more logically expressed as follows:

	Iron.	Oxygen.
Pyrophoric oxide . .	56	16
Red oxide of iron . .	2 × 56	3 × 16
Black oxide of iron . .	3 × 56	4 × 16

We can thus express the composition of all of the oxides of iron by the use of two key numbers—56 for iron and 16 for oxygen—multiplying each by small whole numbers where necessary.

The most important fact of all, however, still remains to be exposed. We have seen that the two compounds of carbon and of oxygen contain these elements in the ratios of 12 to 32 and 12 to 16 respectively. We have seen that the two compounds of oxygen and of hydrogen contain these elements in the ratios of 16 to 2·016 and 16 to 1·008 respectively. This information does not, in itself, give us any indication as to what the composition of compounds between carbon and hydrogen will be. There are a great many compounds formed between these two particular elements, and we shall consider here only the three simplest of them—methane, ethylene, and acetylene. All three of these are combustible gases. Methane (or marsh gas) is formed when vegetation decays under water, and its spontaneous ignition as bubbles rise to the surface is the cause of will-o'-the-wisps. Methane is also one of the main components of ordinary illuminating gas. The brightness of the gas flame when used for lighting purposes, however, is due chiefly to ethylene. Acetylene burns with a still more luminous flame, and used to be extensively employed for bicycle and motor-car headlights. When these three gases are analysed, the surprising nature of the final results obtained is summarized in the following table:

	Carbon.	Hydrogen.
Methane	12	4 × 1·008
Ethylene	12	2 × 1·008
Acetylene	12	1 × 1·008

Not only do we have the law of multiple proportions once more illustrated, but we also have the opportunity of observing that it is possible to employ *the same key numbers* which were used for carbon and for hydrogen in their compounds with oxygen in depicting the proportions of these elements present in their compounds with one another.

This is found by experiment to be true, not only for this particular pair of elements, but *for all elements*. The general fact thus brought to light may again be expressed in the form of a law—the " law of combining proportions " —which runs thus : In every compound substance the proportion by weight of each element may be represented by a fixed number, one for each element, or by this number multiplied by two, three. four, five, or some other small whole number. To put the matter in still plainer terms, we might imagine that each of our elements has a particular telephone number. When we call up any element, we *must* use this number—no other number will do. It may happen, however, that the operator is not very responsive, and then it will be necessary for us to repeat the proper number several times before it finally penetrates.

There must be some important reason for this wonderfully simple scheme into which the compositions of all of the hundreds of thousands of compounds known to chemists fit so perfectly. How its complete mechanism has been made manifest to us will be disclosed in the following chapter.

Chapter VI

THE MIGHTY ATOMS

> Chymists! pursue your experiments with indefatigable ardour and perseverance. Give us the best possible Bread, Butter, and Cheese, Wine, Beer and Cider, Houses, Ships and Steamboats, Gardens, Orchards, Fields, not to mention Clothiers or Cooks. If your investigations lead accidentally to any deep discovery, rejoice and cry "Eureka!" But never institute any experiment with a view or a hope of discovering the first and smallest particles of Matter.
>
> —JOHN ADAMS.

As a chymist, John Adams was evidently a good President of the United States. The quotation given above is from a letter addressed to Professor Gorham of Harvard in 1817, and certainly shows no indication of pussy-footing on the subject of prohibition, although the writer does conclude with the significant words: "I pray you consider this as confidential. If it should get abroad, I should be thought a candidate for the new Hospital."

Now, in spite of the ban which President Adams would have imposed upon experiments on "the first and smallest particles of matter," chemists have continued to be actively interested in this impractical topic, and their labours have been rewarded with truly remarkable success. Indeed it is obviously impossible to prevent people from sharpening their wits upon problems of this character; the less likely the prospects of obtaining a solution, the more attractive the attempt appears. Thus the argument "Which came first, the hen or the egg?" has furnished entertainment for amateur philosophers for untold centuries. The discussion as to the effect of an irresistible force acting upon an immovable body is equally ancient.

From the very earliest times when man began to wonder what the material universe was made of, and whence it arose, one of the questions which have provoked most intense speculation is this : " Can we continue to divide things indefinitely or not ? " In other words, if we take a material object, such as a gold coin or a drop of water, and split it into two and then into two again, and then into two again, and so on as long as we care to keep up the good work, shall we ever reach a point where we *have* to stop ? Only the first stages of the division, of course, are matters which can easily be put to a practical test, but the human brain is more difficult to tire than the human hand, and the Greek philosophers became very deeply involved, indeed, in the purely abstract dispute as to whether or not it was still logical, long after our piece of gold had diminished to an invisible speck or our drop of water had vanished into imperceptible spray, to persist in visualizing it in our imagination as capable of further division for ever and ever.

A party finally arose, headed by Democritus, which put a plank in its platform asserting that there was a limit beyond which we could not go. All matter, according to Democritus, was granular in its structure ; it was made up of very, very minute particles or " atoms," and when we had (in fancy or otherwise) subdivided any particular sample of matter until we had descended to the individual atoms, then it was impossible to persevere any longer, since the atoms were indestructible, indivisible, unchangeable, and eternal. Democritus had no specific grounds for supporting this point of view ; he just felt that way. Possibly he had a personal grudge against some philosopher of the rival party of Anaxagoras, who favoured the opposite opinion, that all matter was continuous and infinitely divisible. At any rate, Democritus did nothing more with the conception of atoms except to bequeath it to one of the Roman poets, Lucretius, who put it into eloquent verse in a six-volume treatise entitled *Concerning the Nature of Things*. Lucretius, however, was popularly supposed to have gone mad through the administration of a love-philtre, and the universal prestige enjoyed by the principles of Aristotle (who was a rabid " continuist ") through the Middle Ages kept the atoms completely in eclipse. Attempts were made, it is true, by Boyle and by Newton, as well as by many men of lesser fame, to reopen the subject, but their arguments that matter was made up of minute particles, or " corpuscles," or " molecules," found little favour and were quickly forgotten. The general feeling regarding such lines of thought, as late as the beginning of the nineteenth century, may be well illustrated by another quotation from John Adams :

> In former times, when I looked a little into Classicks, and a very little indeed it was, while I was fascinated with the Numbers of Lucretius, I could not comprehend his Atoms. In after times, when I was delighted with the eloquence of Buffon, I could not help laughing at his Molecules.

However, at the very same period that President Adams was poking fun at atoms and molecules, our old friend Dalton was busily engaged at Manchester in investigations that were destined to upset the whole apple-cart. Dalton added two simple suggestions to the classical concept of atoms. Indeed, it is surprising how simple all great discoveries in science are, after somebody else has made them. The first idea introduced by Dalton was that, while the atoms of the same element are all alike in weight, the atoms of different elements differ in weight. The second assumption was that, when elements combine, the individual atoms come together in pairs or groups (or "molecules," as Dalton called them), or arrange themselves in a regular pattern, as wholes, and are never broken. In other words, the structure of a compound made up from two elements A and B is such that we have present either equal numbers of atoms of A and B, or two atoms of A to every one of B, or some other simple whole-number ratio. We cannot possibly have varying or complicated intermediate proportions of the two types of atoms in the compound, since chemical union consists, according to Dalton, in an elaborate regrouping of the ultimate particles of the elements into a fixed design, and if we could isolate and examine single "molecules" of the compound, we should find that they were all uniform in containing atoms of A and atoms of B in the same simple whole-number ratio.

Let us now refer back to the last chapter and see if we can apply Dalton's ideas to the numerical results which were there described. After a little contemplation a sudden light will burst upon us. It will be evident that, if we assume that the relative weights of the atoms of carbon, of oxygen, of hydrogen, of iron, and of sulphur are 12, 16, 1·008, 56, and 32 respectively, then we have obtained at one and the same time a simple explanation of three of the most important laws of chemistry—namely, the law of definite proportions, the law of multiple proportions, and the law of combining proportions. When carbon combines with oxygen in the ratio of 12 to 32 to form carbon dioxide, then we have one atom of carbon to two atoms of oxygen. When carbon combines with oxygen in the ratio of 12 to 16 to form carbon monoxide, then we have one atom of carbon to one of oxygen. Each of the two compounds has a definite composition, and there is a simple relation between the proportions of the two elements present in both. We cannot possibly have varying or intermediate compositions, since the two elements must combine in simple atomic ratios.

In the same way the three oxides of iron are built up as follows:

Pyrophoric oxide: 1 atom of iron to 1 of oxygen.
Red oxide of iron: 2 atoms of iron to 3 of oxygen.
Black oxide of iron: 3 atoms of iron to 4 of oxygen.

As a final example, the three compounds of carbon and hydrogen may be cited:

Methane: 1 atom of carbon to 4 of hydrogen.
Ethylene: 1 atom of carbon to 2 of hydrogen.
Acetylene: 1 atom of carbon to 1 of hydrogen.

We see now why each element possesses its own "telephone number" which we must ring up, whatever compound of the element we are interested in. If there is only one atom of this particular element in the molecule, or simplest possible unit of the compound, then we get our connection right away. If there are two atoms, then we have to repeat our number again, and so on. It is no use trying to get through by using any other number, or any fraction of the right number, since we cannot have varying or fractional proportions of atoms present. This distinctive telephone number possessed by each element is so important a property that chemists have given it a special name. Since it expresses the relative weights of the atoms of each species of matter, it is called the *atomic weight.*

The exact atomic weights of all of the elements will be found in a table in the Appendix. Frequently, for reasons of convenience, we have employed, and shall continue to employ, in the course of our discussions throughout this volume, approximate values. Thus, instead of writing the clumsy figure 55·84 for the atomic weight of iron, we can round it off to the nearest whole number 56, while for most purposes the atomic weight of hydrogen may just as well be regarded as 1 instead of as 1·008. Occasionally, however, it will be found that we *must* emphasize the exact values in order to drive home some important theoretical point. Do not let the decimals annoy you too seriously, at any rate just at present.

The whole scheme of things, as outlined above, appears so simple and self-evident that you will be amazed to learn that Dalton's atomic theory was accepted by the chemists of his time only very gradually and with the greatest reluctance. There were two reasons for the hesitation shown in accepting the new ideas which Dalton put forward. The first reason was that Dalton himself complicated the problem by his very efforts to keep it as clear as possible. Dalton assumed that the most natural way for elements to combine was in *equal* atomic ratios, and he accordingly supposed that the most common compound of each element contained equal numbers of atoms of the two elements. Thus, since water was the only compound of hydrogen and oxygen known at the time, Dalton assumed that a molecule of water contained one atom of hydrogen to one atom of oxygen. This made the relative weights of the atoms of hydrogen and oxygen one to eight, instead of one to sixteen. In the same way, Dalton thought that a molecule of ethylene contained equal numbers of atoms of carbon and of hydrogen, and that in a molecule of methane the ratio was one to two. This entailed assigning to carbon the telephone number of six, instead of the number twelve which we now use. Chemists who followed Dalton thus dug themselves into a hole from which they had a great deal of trouble in emerging, and it was many years before agreement was reached as to which values were really correct. Without entering into further discussion of the question here, and without even stopping to justify our own choice of figures, we shall simply state that this difficulty is now completely settled and that the numbers which we are using now are known to be right by several independent lines of reasoning. You will receive convincing proof of this in later chapters.

A second difficulty arose owing to the fact that many of the original analyses made by Dalton of typical compounds were very inexact, and consequently the combining proportions which he derived therefrom did not lead up to the true relative weights of the atoms. Dalton was a very crude experimenter, and it must be remembered that he had few of our modern facilities for making accurate measurements at his command. His first published ratio for the weights of oxygen and hydrogen in water was 5·5 to 1. This was soon amended to 5·66 to 1, later to 7 to 1, and finally (much later) to 8 to 1. Similar errors were made in the analyses of other substances, and the correct combining proportions were only slowly recognized.

At the present time this difficulty also no longer exists.

We are able to take the utmost precautions against error in our experiments, and by the use of the most refined apparatus we can establish the relative weights of atoms in most cases to several places of decimals. It must be carefully borne in mind, however, that the "telephone numbers" which we use for the different elements do *not* represent the weights of single atoms. A single atom is far too minute an object to be isolated and weighed directly. All that we can do is to compare the weights of equal numbers of atoms and see how much heavier those of one element are than those of another. The basis employed in fixing the numbers on our switchboard is oxygen = 16, and when we state that the atomic weight of hydrogen is 1·008 we mean that the weights of the hydrogen atom and the oxygen atom are in the ratio of 1·008 to 16. The question might very well be asked: "Why pick on oxygen = 16 as a standard? Why not rather use hydrogen = 1?" The hydrogen atom is the lightest of all atoms, and it would seem more natural to select this as our unit rather than to give it the awkward value of 1·008. As a matter of fact, hydrogen = 1 was employed as a standard by chemists for many years, and it led to an awful lot of confusion. Hydrogen combines directly with very few other elements, whereas oxygen combines directly with almost all. We can obtain the exact combining proportions of most elements with respect to oxygen very easily, but the accurate determination of the hydrogen-oxygen ratio in water is, as you will notice from Dalton's early results, a matter of much greater difficulty. As soon as chemists awoke to the fact that it was necessary for them, under the hydrogen standard, to change the atomic weight of almost every element on the list every time some ingenious person redetermined the hydrogen-oxygen ratio a little more accurately, they very sensibly decided to go over to the more rigid oxygen basis. Even although we have succeeded more recently in fixing the hydrogen-oxygen ratio very exactly, new reasons have also developed, as you will see later, which support our decision not to shift back.

Now that we have cleared our mind on these points, we may proceed to make some practical use of Dalton's ideas. The first way in which the atomic theory can be put to work is in assisting us to express the compositions of compounds in what may be called chemical shorthand. To give you a hint as to how this is done, a short extract from one of Dalton's note-books may be presented.

○	Hydrogen	⦶	Azote
⊙	Oxygen	●	Carbone, pure charcoal
		⊕	Sulphur

⦶⊙⦶	Nitrous Oxide	○⊙	Water
⦶⊙	Nitrous Gas	⦶○	Ammonia
⊙⦶⊙	Nitric Acid	⊙●	Gaseous Oxide of Carbon
⊙⦶⊙ ⦶⊙	Nitrous Acid	⊙●⊙	Carbonic Acid
		⊙⊕⊙	Sulphuric Acid
⊕⊙	Sulphurous Acid		

When you recollect, however, that ninety elements in all are now known, a scheme like this evidently becomes too complicated for general use, and a much simpler system was soon introduced by Berzelius, a Swedish chemist. Berzelius gave each element a symbol; for example, oxygen was denoted by O, hydrogen by H, and carbon by C. Let us clash a few of the most important symbols for you in the following table:

Element.	Symbol.	Approximate Atomic Weight.	Element.	Symbol.	Approximate Atomic Weight.
Aluminium	Al	27	Mercury	Hg	200
Calcium	Ca	40	Nitrogen	N	14
Carbon	C	12	Oxygen	O	16
Chlorine	Cl	35·5	Phosphorus	P	31
Copper	Cu	63·5	Potassium	K	39
Gold	Au	197	Silicon	Si	28
Hydrogen	H	1	Silver	Ag	108
Iron	Fe	56	Sodium	Na	23
Lead	Pb	207	Sulphur	S	32
Magnesium	Mg	24	Tin	Sn	119

These symbols are international and no general rule has been observed in the christening ceremonies. Sometimes the symbol is derived from the Latin name of the element; for example, Au = *aurum*, gold; Ag = *argentum*, silver. Sometimes the German name has been called into service; for instance, K = *kali*, potassium (compare our word alkali). In many cases the first letter of the English name of the element is used, but since the twenty-six letters of the alphabet are not sufficient to go around, a second letter is often necessary; thus C = carbon, Ca = calcium, Cl = chlorine.

Compounds may be represented by writing down the symbols for the different elements of which they are composed, placing a small number as a subscript after each symbol when we wish to denote that more than one atom of that particular element is present in each molecule of the compound. For example, carbon dioxide is written CO_2, carbon monoxide CO. The three oxides of iron appear as FeO, Fe_2O_3, and Fe_3O_4 respectively. Such a collection of letters and numbers is called a formula. If you are an addict of the cinema, you will remember that every chemist who appears on the silver screen spends all his time looking very intently into a test-tube trying to discover a "formula." As soon as he discovers it, the villain steals it from him, and then the plot begins to thicken. Well, now you know what a formula is and you can judge how valuable it might be.

It is very important to bear in mind that the formula tells us not only what elements are present in a compound but also the proportions by weight in which they are present. Thus CO_2 is chemical shorthand which may be expanded not only into the statement that carbon dioxide contains one atom of carbon to two atoms of oxygen, but also into the statement that it contains 12 grammes of carbon to $2 \times 16 = 32$ grammes of oxygen. In order to keep this dual meaning in mind, the chemist generally reads a symbol as denoting not one atom of an element but *one gramme atomic weight* of an element. Thus O in a formula equals 16 grammes of oxygen; C denotes 12 grammes of carbon, and so on. On the same principle CO_2 is understood to signify $12 + 32 = 44$ grammes of carbon dioxide, and this quantity is called the *gramme molecular weight*.

When we put several symbols together we get a formula; when we put several formulas together we get an equation. An equation is a useful and compact way of writing down the essential facts of a chemical reaction. Thus when we made one of our first experiments and persuaded iron and sulphur to combine to form ferrous sulphide, we wrote the reaction like this:

$$\text{Iron} + \text{Sulphur} \rightarrow \text{Ferrous sulphide.}$$

We may now express the same reaction much more neatly as follows:

$$Fe + S \rightarrow FeS.$$

This tells us not only the qualitative fact that iron and sulphur combine to form ferrous sulphide, but also gives us the quantitative information that one gramme atomic weight (56 grammes) of iron combines with one gramme atomic weight (32 grammes) of sulphur to form one gramme molecular weight (88 grammes) of ferrous sulphide.

If we could make the other compound of iron and sulphur, "fools' gold," by direct combination from the elements, the equation would run :

$$Fe + 2S \rightarrow FeS_2.$$

A numeral placed before a symbol or a formula signifies that more than one atomic or molecular weight of the substance must be taken in order to make the equation "balance." We cannot create or destroy matter, and we must have the same number of atoms of each element present on the two sides of every equation.

The equation for the combustion of carbon in oxygen to give carbon dioxide (watch this carefully !) is written as follows :

$$C + O_2 \rightarrow CO_2.$$

We are prepared to hear your immediate protest : "Why in the world do you write sulphur as plain S, iron as plain Fe, carbon as plain C, and then suddenly change your plan and write oxygen as O_2 ?" There is a reason, of course, for this apparent inconsistency. If you had not already sustained a great deal of heavy reading in this chapter, we should go straight ahead and develop this reason for you in complete detail. To do this, however, would require you to follow us through another long theory which parallels the atomic theory in its importance in chemistry and is called the "kinetic-molecular hypothesis." Perhaps, now that you have heard its name, you will be agreeable to its postponement for the present, since this volume is already somewhat overburdened with theoretical discussions.

All that you really need to know immediately is this. Free oxygen does *not* consist of single oxygen atoms. In many gaseous elements, such as oxygen, hydrogen, and chlorine, the atoms are always, under ordinary conditions, associated in pairs, and each group of two atoms behaves as a distinct and separate unit. Such a unit containing two atoms in very close companionship conforms evidently to Dalton's idea of a molecule, the only novel feature which we have to recognize being that the atoms coming together here are similar in nature instead of being atoms of different elements. When we study the behaviour of gases critically, it is the molecule and not the individual atom that we find to be of primary importance. We therefore always write the formula of a gaseous substance, whether elementary or compound, in such a way as to indicate the magnitude of the molecule. Thus we write oxygen as O_2, hydrogen as H_2, nitrogen as N_2, and so on. For the same reason, we write ethylene as C_2H_4 instead of CH_2, acetylene as C_2H_2 instead of CH, and hydrogen peroxide as H_2O_2 instead of HO. These formulas are called *molecular formulas*. They may not mean much to you at present, but you will find them indispensable later.

We have now brought you to a point where you are prepared to meet the elements on really intimate terms. The first individual to whom we shall introduce you is . . .

Chapter VII

OXYGEN: THE WORKING-GIRL

> Heaven will protect the working-girl.
>
> —*Old Popular Song*

Do not be afraid. Heaven has already protected her.

If we had kept up the motor-car analogy with which this book began, it would have been necessary for us, since oxygen is by far the most abundant of all the elements, to call her Lizzie. As matters now stand, however, she becomes entitled to the much more appropriate and attractive name of Cinderella. For a long time she toiled in seclusion among the ashes, and nobody had any notion how beautiful she really was. Her fairy godmother, however, finally appeared and she is now, after passing through various misadventures, recognized by everybody as the belle of the ball and is privileged to take her pick of eligible princes.

The difficulties which oxygen went through in emerging out of her former obscurity into the realms of romance were mainly due to the fact that she suffered from an excess of fairy godmothers—or, more accurately, godfathers. Once her existence began to be suspected they clustered around her so closely that they got in one another's way and quarrelled over their respective claims to lead her out into the limelight. It is a long story, and we do not intend to imitate the tabloid press by baring all the lurid details. We shall, indeed, restrict our account of the discovery of oxygen to the part played therein by two celebrated chemists, Priestley and Lavoisier.

Joseph Priestley was born near Leeds, in Yorkshire, in 1733. His father was not well-to-do, being, like Dalton's father, a weaver and dresser of cloth, carrying on his work at home with a hand-loom, as was the custom in those times. His mother died when he was only seven years old, and he was placed in the care of an aunt, who had strong Calvinistic convictions. It is no wonder, then, that when Priestley grew up and entered the ministry he exhibited an independent spirit which drew upon him the displeasure of his congregations. After several attempts to eke out a living as a schoolmaster, he accepted an invitation to take charge of Mill Hill Chapel, at Leeds, and in that city he began his chemical work. He was first attracted to science through meeting Benjamin Franklin—an interesting fact in view of his later career. His studies were almost entirely on different gases, and he relates in characteristically candid fashion how he first became interested in this field:

> It was in consequence of living for some time in the neighbourhood of a public brewery that I was induced to make experiments on fixed air, of which there is a large body, ready formed, upon the surface of the fermenting liquor.

By " fixed air " Priestley means what we now call carbon dioxide, CO_2. It may interest you to know that Priestley was the first man to observe the agreeable taste of a solution of this gas in water and that he thereby became the inventor of soda-water—" a service," said Huxley, " to naturally, and still more to artificially, thirsty souls, which those whose parched throats and hot heads are cooled by morning draughts of that beverage cannot too gratefully acknowledge."

Priestley's chief claim to fame, however, rests on his work with oxygen. The offhand way in which he went about his experiments may best be appreciated by giving an account of this in his own words, with the addition of a few explanatory notes in brackets. The events described below happened on the first of August, 1774.

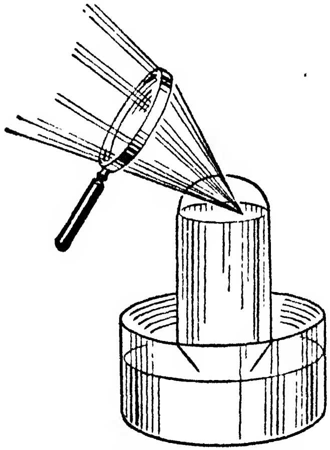


FIG. 16.—Priestley's Apparatus.

Having procured a lens [burning-glass], I proceeded with great alacrity to examine, by the help of it, what kind of air [gas] a great variety of substances would yield, putting them into vessels filled with quicksilver, and kept inverted in a basin of the same (*Fig. 16*). After a variety of other experiments, I endeavoured to extract air [gas] from *mercurius calcinatus per se* [mercury rust], and I presently found that, by means of this lens, air [gas] was expelled from it very readily. Having got about three or four times as much as the bulk of my materials, I admitted water to it, and found that it was not imbibed by it [the gas did not dissolve]. But what surprised me more than I can well express, was that a candle burned in this air with a remarkably vigorous flame. I was utterly at a loss to account for it.

In October of the same year Priestley took a trip to Paris, and remarks in his diary, "I frequently mentioned my surprise at the kind of air which I had got from this preparation to Mr. Lavoisier, Mr. Le Roy and several other philosophers, who honoured me with their notice in that city, and who, I dare say, cannot fail to recollect the circumstance." Another extract from Priestley's journal may now be given:

> In ignorance of the real nature of this kind of air, I continued from this time [November] to the first of March following. Till this first of March, 1775, I had little suspicion of the air from *mercurius calcinatus* being wholesome. I particularly remember telling Dr. Price that I was myself perfectly satisfied of its being common air; though for the satisfaction of others, I wanted a mouse to make the proof quite complete.
>
> On the eighth of this month I procured a mouse, and put it into a glass vessel, containing two ounce measures of the air from *mercurius calcinatus*. Had it been common air, a full-grown mouse, as it was, would have lived in it about a quarter of an hour. In this air, however, my mouse lived a full half-hour; and though it was taken out seemingly dead, it appeared to have been only exceedingly chilled; for, upon being held to the fire, it presently revived, and appeared not to have received any harm from the treatment.

About this time Priestley, though still not certain regarding the nature of his new product, decided to call it "dephlogisticated air," or air deprived of phlogiston. Priestley was a devoted believer in the phlogiston theory, which—as we have already had occasion to mention—insisted that the rust of a metal was formed by driving phlogiston out of it. Under this theory it is hard to see exactly how the gas given off when a rust was reconverted into metal could be explained at all; if it was anything, it was *negative phlogiston*! However, Priestley did not appreciate this difficulty clearly. Having given his gas a name, he went on to do some more experiments with it, as follows:

> From the greater strength and vivacity of the flame of a candle in this pure air, it may be conjectured that it might be peculiarly salutary to the lungs in certain morbid cases, when the common air would not be sufficient to carry off the phlogistic putrid effluvium fast enough. But, perhaps, we may also infer from these experiments, that though pure dephlogisticated air might be very useful as a *medicine*, it might not be so proper for us in the usual healthy state of the body: for, as a candle burns out much faster in dephlogisticated than in common air, so we might, as may be said, *live out too fast*, and the animal powers be too soon exhausted in this pure kind of air. A moralist, at least, may say that the air which nature has provided for us is as good as we deserve.
>
> My reader will not wonder, that, after having ascertained the superior goodness of dephlogisticated air by mice living in it, and the other tests above mentioned, I should have the curiosity to taste it myself. I have gratified that curiosity, by breathing it, drawing it through a glass siphon, and, by this means, I reduced a large jar full of it to the standard of common air. The feeling of it to my lungs was not sensibly different from that of common air; but I fancied that my breast felt peculiarly light and easy for some time afterwards. Who can tell but that, in time, this pure air may become a fashionable article in luxury. Hitherto only two mice and myself have had the privilege of breathing it.

So much for Priestley's work on oxygen. His later fortunes may be outlined in brief. After the outbreak of the French Revolution, Priestley made himself very obnoxious to his fellow-countrymen in England by his open sympathy with its leaders, and in 1791 a riot occurred in the course of which his house and laboratory at

Birmingham were attacked by the mob, set on fire, and destroyed. Priestley himself, fearing personal violence, decided to leave his native country, and in 1794 he followed his three sons to America. He was enthusiastically received at first and was offered the professorship of chemistry in the University of Pennsylvania. Declining this, he settled down to continue his scientific and theological studies at Northumberland, Pennsylvania, where he remained until his death in 1804. His last years were occupied with cantankerous controversies with his American colleagues regarding the phlogiston theory. Priestley refused to be convinced that the gas which he had himself discovered had sounded the death-knell of this chimera, and he went to his grave ignorant of the real significance of the wonderful contribution which he had made to chemistry. With all his genius, Priestley was essentially a very stubborn and a rather stupid man. One chemist has said that he made so many mistakes in his work that it was necessary for him to invent the use of india-rubber as an eraser in order to put them right!

Let us now turn to a man of a totally different type, Lavoisier. Born in Paris in 1743, the son of a wealthy lawyer, Antoine Laurent Lavoisier turned to science at an early age, and in 1766 was awarded a gold medal by the Academy for an essay on the best method of lighting the streets of a large city. While Priestley held that "more is owing to what we call chance than to any proper design or perceived theory in this business," Lavoisier in his scientific work followed the example of Joseph Black in performing all of his experiments according to a definite plan and in a quantitative way. He never tired of weighing and measuring. Just before Priestley's visit to Paris in 1774, Lavoisier had found that when tin was heated with air in a closed vessel the tin gained in weight and part of the air disappeared. Lavoisier recognized that these facts could not be explained under the phlogiston theory, and as soon as he was informed of Priestley's observations he proceeded to deduce from them the true facts which Priestley himself had missed. In 1777 Lavoisier heated the metal mercury in a retort (*Fig. 17*), the neck of which projected into a bell-jar

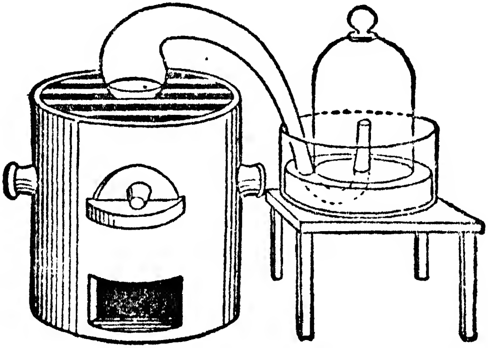


FIG. 17.—Lavoisier's Apparatus.

standing in a larger dish of mercury. After twelve days, the enclosed air had diminished in volume by one-fifth and the surface of the mercury in the retort was covered with red particles of mercury rust. The gas which remained was found to support neither life nor combustion. The mercury rust, when heated more strongly by itself, broke up into mercury and a gas, the volume of which corresponded exactly with the shrinkage that the air in the first experiment had undergone, and this gas supported life and combustion to an exaggerated degree. Lavoisier was the first to call this gas oxygen, and the experiment which he made was the first decisive proof that this gas was the active component of the air. You will remember that we carried out some similar experiments in Chapter III, but without the prior experience of Lavoisier we should probably not have found it feasible to interpret our results correctly.

It would lead us too far astray to enter here into the debate as to whether Priestley or Lavoisier deserves to be placed first as fairy godfather to our Cinderella. Priestley discovered oxygen; Lavoisier recognized her true character. We shall leave the matter there.

To complete our detour into history it remains only for us to relate the tragic and untimely close of Lavoisier's life. Lavoisier was for many years a servant of the French Government. In that capacity he made investigations which led to a great improvement in the quality of the gunpowder manufactured in the factories of his country. He was also, unfortunately, a farmer-general of the revenue; that is, he was a member of a syndicate of financiers to which the State farmed, or leased, the right to collect internal taxes, such as the duty on salt, tobacco, and other goods. A system like this was evidently open to grave abuses, and the agents of the Ferme-Générale were only too prone to enrich themselves by unjust and oppressive extortions of money from the general public. When the French Revolution broke out, therefore, the members of the Ferme did not escape the hatred of the mob; Lavoisier, with his colleagues, was brought to trial in Paris during the Reign of Terror and sentenced to die on the guillotine within twenty-four hours. When the presiding judge was petitioned to delay the execution in order that Lavoisier might complete some of his experiments, the retort was made: "The

Republic has no need of chemists; let justice take its course!" Yet only a few years later the French Republic, ringed around with enemies, was almost at the point of collapse owing to the lack of chemicals necessary for carrying on the conflict. Common substances like soda and saltpetre were produced in sufficient quantity with the greatest difficulty, and a Lavoisier would have been of even greater service to his country at that time than a Napoleon.

It is, perhaps, a foolish speculation, but there is a certain lure in wondering what would have happened to chemistry in America if the English and the French nations had interchanged their methods of dealing with undesirable chemists. If Priestley's head had fallen under the guillotine and Lavoisier had been exiled to America, it is not unreasonable to assume that chemistry would have begun to flourish actively in the United States long before it did in fact. Lavoisier was undoubtedly a genius on a much higher plane than Priestley. When he was murdered, the great French mathematician Lagrange remarked, "A moment was all that was necessary in which to strike off this head, and probably a hundred years will not be sufficient to produce another like it." The famous German chemist Liebig later said, "Lavoisier discovered no new body, no new property, no natural phenomenon previously unknown; all the facts established by him were the necessary consequences of the labours of those who preceded him. His merit, his immortal glory, consists in this, that he infused into the body of the science a new spirit." Instead of wasting his time in futile and despairing efforts to disinter the corpse of the phlogiston theory, as Priestley did in America, Lavoisier would probably have added new and equally noteworthy contributions to those which he had already made to chemistry. It may be, however, that we are all wrong in our regrets. Priestley, as we have observed, was somewhat stupid, but he was always honest. Lavoisier was more ingenious, but he occasionally betrayed his legal ancestry by indulging in rather sharp practice. His memory has, it is true, been cleared of the charge of deliberately stealing Priestley's results; he himself publicly acknowledges his indebtedness to Priestley for the suggestions which led up to his epoch-making experiments. But on other points his reputation is not altogether spotless, and when we recall that the main indictment on which he was found guilty and put to death in Paris was that of watering the soldiers' tobacco, we may be excused for harbouring the suspicion that if Lavoisier had gone to America he might not have applied india-rubber to its legitimate use as an eraser but might instead have been the first to propose its employment as an ingredient in chewing-gum.

Enough now of fairy godfathers; we are neglecting Cinderella most shamefully. And this neglect is all the more inexcusable since oxygen, instead of letting popularity spoil her, continues to function in the interests of mankind in the same hard-working fashion that she did before her début, leaving the dance promptly at the stroke of twelve and never neglecting her allotted tasks.

The services which oxygen renders to us all, indeed, are incomparably more valuable than those given by any other element. We could manage to exist quite well without gold, but we need oxygen every moment of our lives. Ordinarily, the amount contained in the air is quite sufficient for us, but Priestley's surmise that pure oxygen might be valuable as a medicine is now justified by its common use to relieve patients suffering from pneumonia or suffocation. The pure gas is also employed to vivify the atmosphere inside submarines and to assist the breathing of aviators and mountain climbers at very great altitudes.

The oxygen used for these purposes is not obtained, of course, by heating mercuric oxide. There is a much cheaper supply available in unlimited quantities in the air. When air is cooled to an exceedingly low temperature (almost 200° below zero on the Centigrade scale) it condenses to a liquid in just the same way that steam condenses to water when it is cooled beyond 100° C. If this "liquid air" is allowed to evaporate slowly, the nitrogen in the mixture boils off much more freely than the oxygen, just as when we boil a mixture of alcohol and water the alcohol distils more quickly than the water. After a time the residual liquid is almost pure oxygen, and the gas coming off from this is compressed by pumps into the steel cylinders (*Fig. 18*) in which it is sold.

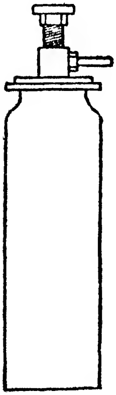


Fig. 18.—Oxygen Cylinder.

Another important use of oxygen was also foreshadowed by Priestley. A final extract from his journal follows:

> The dipping of a lighted candle into a jar filled with dephlogisticated air is alone a very beautiful experiment. The strength and vitality of the flame is striking, and the heat produced by the flame, in these circumstances, is also remarkably great.
>
> Nothing, however, would be easier than to augment the force of fire to a prodigious degree, by blowing it with de-

phlogisticated air instead of common air. This I have tried, in the presence of my friend, Mr. Magellan, by filling a bladder with it, and puffing it, through a small glass tube, upon a piece of lighted wood.

Possibly much greater things might be effected by chymists, in a variety of respects, with the prodigious heat which this air may be the means of affording them. I had no sooner mentioned the discovery of this kind of air to my friend, Mr. Mitchell, than this use of it occurred to him. He observed that possibly platina might be melted by means of it.

Every city dweller who has seen workmen on a new building cutting pieces of steel apart by means of an " oxygen torch," in which intense heat is furnished by the combustion of hydrogen or acetylene in pure oxygen, will recognize that Priestley's prophecy has been more than fulfilled. When the flame from this torch is allowed to play upon a piece of quicklime, the solid becomes white-hot and glows brilliantly, giving what we all know as *calcium light or limelight.*

Our dependence upon oxygen in combustion is, however, not limited to these particular cases. Power and heat for commercial purposes, as well as in our own homes, are almost always obtained by burning coal or oil, and in this burning the oxygen from the air plays an essential part. If we had to buy the oxygen as well as the coal, it would be necessary for us to purchase at least three tons of oxygen for every ton of coal. Because we get the oxygen free, we forget our debt to Cinderella. Even now, you see, her devotion to duty is not properly appreciated.

This applies also to another form of fuel universally employed—namely, the food we eat. Every woman studies her calories nowadays, but few understand exactly what the term means. We obtain all the energy to carry on our bodily activities by the combustion of the carbon and hydrogen in our food with the oxygen that we breathe. The amount of heat (or the number of calories) obtained by burning a given amount of food is therefore just as much due to the oxygen as to the food itself. But, again, we have to pay for the food, while we get the oxygen free, and so we remain quite oblivious of our obligations to Cinderella in this connection also.

Now for some straight chemistry. The most important chemical property which oxygen possesses is its power as a very active element to combine with other elements. Elements like carbon, sulphur, and phosphorus, which burn in the air, all burn much more vigorously and brightly in pure oxygen. Carbon, as we have already seen, burns to give a gas called carbon dioxide. Sulphur similarly gives a gas with a stifling odour called sulphur dioxide. Phosphorus gives a white powdery solid, phosphorus pentoxide. The equations for the three reactions are given below:

$$\begin{aligned} C + O_2 &\rightarrow CO_2 \\ S + O_2 &\rightarrow SO_2 \\ 4P + 5O_2 &\rightarrow 2P_2O_5. \end{aligned}$$

The names for the three oxides denote the number of atoms of oxygen present in the molecule in each case, the prefix *di* meaning two and the prefix *penta* meaning five. Note that we must have the same number of atoms present on both sides of each equation, so that it will balance correctly.

Metals also burn in oxygen much more rapidly than in air. A strip of magnesium (the metal which is employed in flash-light powder), ignited by a match and plunged into a jar of oxygen, burns with a very bright flame to give magnesium oxide. In the same way iron wire, which oxidizes in the air only comparatively slowly even at high temperatures, burns with surprising brilliancy in pure oxygen, throwing off a shower of molten globules of the black magnetic oxide The two equations read:

$$2Mg + O_2 \rightarrow 2MgO$$
$$3Fe + 2O_2 \rightarrow Fe_3O_4.$$

You will recall that we have already mentioned the convenience of dividing our ninety elements into two groups, metals and non-metals. Now the chemist has discovered a very interesting fact which enables him to make this division more rigid and systematic. If we take the oxide of any one of Davy's active metals and dissolve it in water, we find that a vigorous reaction occurs and that the resulting solution has certain standard properties. For example, it has a soapy feeling, it turns red litmus (a vegetable extract present in pickled cabbage) blue, and it possesses some other important characteristics which are peculiar to a class of substances called *alkalis* or *bases*. The most common alkalis, caustic soda and caustic potash (or, as the chemist prefers to name them, sodium hydroxide and potassium hydroxide), result from the solution of the oxides of sodium and potassium in water. The equations showing the reactions that occur are written thus:

$$Na_2O + H_2O \rightarrow 2NaOH$$
$$K_2O + H_2O \rightarrow 2KOH.$$

The group OH (or hydroxyl) is an indispensable part of the formula of all bases.

When we add the oxide of calcium, quicklime, to water we obtain another typical base, calcium hydroxide or slaked lime. A great deal of heat is given off in the process, as you have undoubtedly seen for yourself while watching the clouds of steam arising when workmen add water to quicklime and sand in making mortar. The reaction may be written:

$$CaO + H_2O \rightarrow Ca(OH)_2.$$

We write the formula $Ca(OH)_2$ rather than CaO_2H_2 in order to emphasize the presence of the characteristic hydroxyl group OH. Magnesium oxide does not react so energetically with water, but a finely divided suspension of the typical base magnesium hydroxide $Mg(OH)_2$ is familiar to us all under the name of "milk of magnesia."

The oxides of non-metallic elements, on the other hand, react with water to give solutions of an entirely different nature. These solutions have a sour taste, they change the colour of blue litmus back to red, and they possess other important properties which are peculiar to a class of substances called *acids*. The acid formed by dissolving carbon dioxide in water is called carbonic acid. The acid formed by dissolving sulphur dioxide in water is called sulphurous acid. More important than either of these in chemistry is sulphuric acid, which is obtained by dissolving a higher oxide of sulphur, sulphur trioxide SO_3, in water. Phosphorus pentoxide, finally, gives phosphoric acid.

$$\begin{aligned} CO_2 + H_2O &\rightarrow H_2CO_3 \\ SO_2 + H_2O &\rightarrow H_2SO_3 \\ SO_3 + H_2O &\rightarrow H_2SO_4 \\ P_2O_5 + 3H_2O &\rightarrow 2H_3PO_4. \end{aligned}$$

The chemical properties of acids and of bases are so different that chemists find it most profitable to use these two classes of substances as a basis of distinction for the elements from which they are derived. Those elements whose oxides react basic in water solution are classed as metals; those elements whose oxides react acidic in water solution are classed as non-metals. It is true that there are certain elements whose oxides are not sufficiently soluble in water to give us the peculiar soapiness of an alkali or the sour taste of an acid. We can *generally* put such an element in its proper place, however, by noting the action of a typical acid or of a typical base upon its oxide. The plan of attack depends upon the fact that any base will "neutralize" any acid so that, when we mix the two in the right proportions, the properties of both disappear. An instance of this is the use of alkaline tooth-pastes for counteracting an acid condition in the mouth. Another example is the use of milk of magnesia for correcting an acid condition of the stomach. Consequently, when we find that an oxide such as copper oxide, which is quite insoluble in water, dissolves rapidly in an acid like sulphuric acid but is practically unaffected by a base like sodium hydroxide, we conclude that our common habit of calling copper a metal is in complete accordance with the chemical classification.

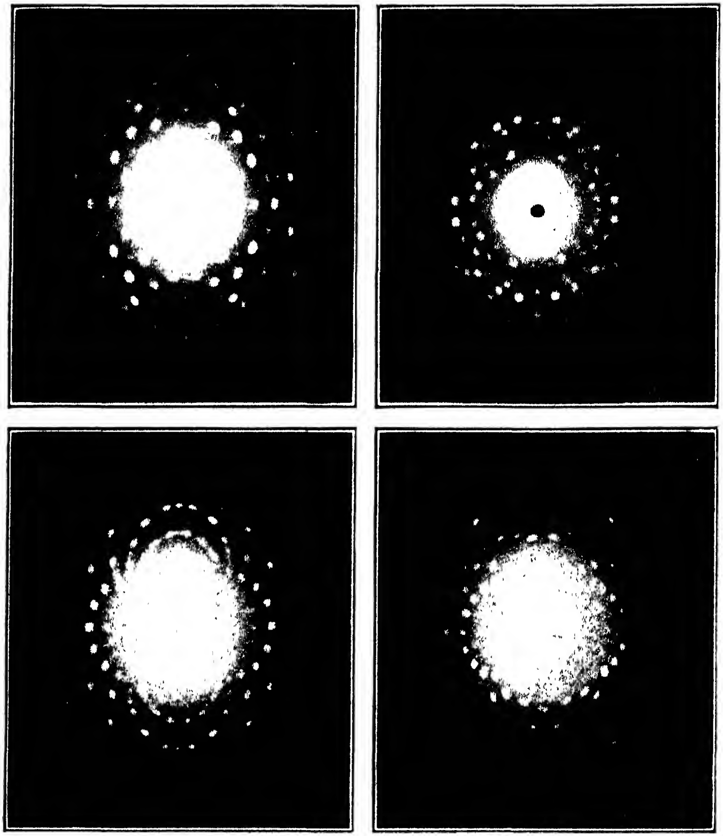
Some time back we noted that the hydroxyl group OH was an indispensable part of the formula of a base. Is there any indispensable part of the formula of an acid? Lavoisier thought that he had the answer; he believed that *oxygen* was necessary in the make-up of every acid. As soon as he waved his wand over Cinderella, therefore, he christened her by this name of oxygen, which he derived from two Greek words meaning *acid-producer*. This is where Lavoisier made a very bad blunder. It was soon suspected that certain typical acids such as hydrochloric acid (the formula for which was finally established to be HCl) contained no oxygen, but it was only after long debate that chemists recognized the real truth—namely, that the essential constituent of an acid is not oxygen but *hydrogen*. Nevertheless, Lavoisier's name has stuck to this day.

The activities of oxygen are not always, it must be confessed, employed in a manner which advances the interests of the human race. Sometimes the element is a serious nuisance. The slow oxidation or rusting of metals in air is a case in point. It has been calculated that more than twenty million tons of iron are rusted away in the world every year, and this rusting, of course, entails a prodigious financial loss. The guilt, however, should not be placed upon the shoulders of Cinderella. It is the metals which are at fault in not being strong enough to resist the fascination of oxygen, and the solution of our corrosion problems is gradually being found in the invention of alloys, or mixtures of metals, which are less susceptible to her charms.

One last topic in connection with oxygen and we are finished. The point has already been noted that under ordinary conditions oxygen in the free state consists of molecules formed by the association of two atoms. Under special circumstances, however, the association of oxygen atoms may even be into groups of three. The substance thus resulting is called *ozone* O_3.

Ozone is obtained in the laboratory by passing a silent electric discharge through oxygen gas, and was first discovered through the penetrating odour which is observed near an electrical machine in operation. It is a blue gas which is very unstable and very rapidly reverts to ordinary oxygen, O_2. It is much more reactive than oxygen and can oxidize many substances which oxygen cannot. In view of the fact that it will oxidize and destroy the bacteria in water much more rapidly than oxygen does, ozone has been used for sterilizing drinking-water in Leningrad, Lille, and other cities. Certain seashore resorts are wont to brag about the ozone present in their bracing breezes, and use its hygienic reputation as a bait to attract visitors. The tests which were formerly employed, however, to show the presence of ozone in the air are now known to be misleading, since other substances give similar reactions, and it is to be feared that the only place where ozone persists for any significant time after its formation is in the advertisements. A party consisting of three is, as everybody knows, not a very stable combination; one member of the trio is always likely to disagree with the others and to be ejected from the group.

Ozone is often called an allotropic form of oxygen, just as diamond and graphite are allotropic forms of carbon. If you ask the chemist, however, whether ozone and oxygen are the same substance or different substances, do not expect to receive an immediate reply. A very pretty argument, indeed, can be put up on both sides of the question. Ozone and oxygen have the same composition, it is true, but even the chemical properties differ with the molecular complexity. Seeing that the experts are still at odds, we advise you frankly not to bother too much about the matter. Difficulties like this are quite common in chemistry. In a sense, they are not to be deplored, since they stimulate investigation and make for future progress. Perhaps on the same principle that a dog is afflicted with fleas in order to prevent him from worrying over the fact that he is a dog, so it is good for the chemist to have problems of this type to occupy his spare time and to prevent him from worrying over the fact that he is a chemist.



TYPICAL LAUE PHOTOGRAPHS

The two upper plates and lower left plate show the photographs obtained by the Laue method of analyzing the arrangement of atoms in crystals, using a crystal of anhydrite (calcium sulphate) in three different positions, so that the X-rays fall perpendicular to three different axes. From a critical examination of these three diagrams the complete lattice structure may be deduced.

The plate on the lower right represents one photograph obtained in the same way with aragonite (calcium carbonate). Both anhydrite and aragonite crystallize in the rhombic system, and a general resemblance between the two lower plates, which represent the crystals in similar positions, may be traced. It is important to note, however, that the resemblance is far from being an identity. A crystal may be recognized from its photograph just as surely as a criminal from his fingerprint.

Chapter VIII

COUNTING THE UNCOUNTABLE AND DIVIDING THE INDIVISIBLE

CAROLINE: To confess the truth, Mrs. B., I am not disposed to form a very favourable idea of chemistry, nor do I expect to derive much entertainment from it. I prefer those sciences that exhibit nature on a grand scale, to those which are confined to the minutiæ of petty details. . . .

MRS. B.: I rather imagine that your want of taste for chemistry proceeds from the very limited idea you entertain of its object. You confine the chemist's laboratory to the narrow precincts of the apothecary's shop, whilst it is subservient to an immense variety of other useful purposes. Besides, my dear, chemistry is by no means confined to works of art. Nature also has her laboratory, which is the universe, and there she is incessantly employed in chemical operations. You are surprised, Caroline; but I assure you that the most wonderful and the most interesting phenomena of nature are almost all of them produced by chemical powers. . . .

CAROLINE: If this is the case, I have certainly been much mistaken in the notion I had formed of chemistry. I own that I thought it was chiefly confined to the knowledge and preparation of medicines.

—JANE MARCET, *Conversations on Chemistry*, 1807.

YOU also will be surprised, dear reader (it is a pity that we cannot call you Caroline, just for this single chapter), at the most wonderful and most interesting consequences that will follow from our examination of "the minutiæ of petty details." Not only will you be astonished to discover how exceedingly minute those petty details, the atoms, are, but you will be called upon to realize the phenomenal fact that these infinitesimal atoms are, contrary to the ideas of Democritus and Dalton, themselves capable of further subdivision.

We do not anticipate that you will be able to gain a complete insight into the experimental methods which have been employed in obtaining the results that we are going to describe. These results may quite well seem to you to be so supernatural that there is a danger that you may consider our efforts to make you understand them somewhat similar to those of a conjurer pretending to initiate his audience into the technique of his favourite tricks. We shall try our utmost, however, to keep our presentation as simple as possible, and we trust that the quickness of the hand will not deceive the eye too often.

In discussing the work of Dalton on the subject of atoms we were careful to call the conclusions which he drew a *theory* rather than a *law*. The distinction between the two words as used by chemists requires a brief investigation. Progress is made in science, most frequently, in the following manner. First of all, some curious person notes a fact, such as the fact that iron rusts in air, and feels uncomfortable because he cannot explain it. He attempts to soothe his discomfort by performing a few experiments, and succeeds in discovering that other metals rust also. This reconciles him in some degree to the fact that iron rusts, and he experiences a certain amount of mental satisfaction in knowing that the case is not an isolated one. His next step is to try to get some inside information about the nature of rusting. Thus, when he discovers that the rust is heavier than the metal and that the increase in weight is due to combination with the oxygen of the air, he feels still happier about the matter and is likely to say that rusting has now been "explained." Notice, however, that all that has actually been demonstrated is the new fact that metals rust by taking up oxygen—which tells rather *how* they rust than *why* they rust. To answer the question why metals rust it is necessary to delve into the unknown much more deeply. The standard method employed in such an emergency is to imagine a state of affairs which, *if true*, would give us a clearer picture of the problem, accounting for all known facts and bringing them into harmony with facts in other, perhaps apparently remote, fields. Such a state of affairs is called a "hypothesis." The final and most difficult step still remains to be made—namely, to verify this hypothesis. Additional evidence is sought in as many independent directions as possible, and when the hypothesis seems to be fairly well established it is promoted to the title of "theory." When at last the proof is regarded as positive, then the theory becomes a "law."

Now, the atoms of Democritus were absolutely hypothetical. Dalton, however, correlated the concept of atoms so beautifully with the law of definite proportions, the law of multiple proportions, and the law of combining proportions, that the step to a theory was clearly justified. Nevertheless, neither Dalton nor any of his generation ever had any direct proof of the existence of atoms. In the same way Darwin's evolution theory constitutes a rational attempt to explain the gradual and progressive changes which have occurred in different species of living things throughout the ages ; it is supported by a mass of convincing evidence, but it involves a number of assumptions which have not even yet been directly confirmed.

Since the weak points of any theory are always exposed to suspicion, the atomic theory has not escaped attack. Just as some eminent public characters have objected to the theory of evolution because they regarded it as merely a hypothesis, or guess, so certain prominent chemists, even as late as the beginning of the present century, openly argued that chemistry, being an experimental science, ought not to be based upon a notion which was incapable of straightforward experimental proof, and suggested that we should relegate the atom to a much minor rôle in the development of the subject, stressing more prominently in its stead the experimental law of definite proportions.

The last few years have destroyed all excuse for such a procedure. Indeed, it is most interesting to note that recent advances in our knowledge, which have made atoms unquestionable realities, have at the same time demonstrated to us that our so-called law of definite proportions is not a fundamental law at all in its original form, but is open to exceptions and must be modified. Not only are we sure now that atoms exist, but we know exactly how big and how heavy (or, rather, how small and how light) the individual atoms of each element are.

The fact has already been mentioned in a previous chapter that one of the physical properties which the chemist uses for distinguishing substances is crystalline form. A perfect crystal of table salt (*Fig. 19*) differs altogether from a perfect crystal of cane-sugar (*Fig. 20*); quartz, mica, and feldspar all show definite and characteristic crystal habits. A perfect crystal is most readily obtained by gradual growth, from a liquid which is freezing or from a solution which is being evaporated.

The slow addition of material occurs throughout as if some unseen hand was constructing a carefully premeditated design, and the most fascinating and remarkable patterns may be thereby produced. Every one has seen the beautiful fern-like tracings which appear upon the window-pane in cold weather due to the crystallization of a thin layer of ice. When water freezes in larger masses we obtain more regular patterns, such as are shown in snow crystals (*Fig. 21*).

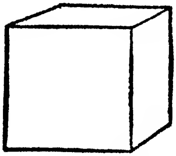


FIG. 19.—Crystal of Sodium Chloride (Rock Salt).

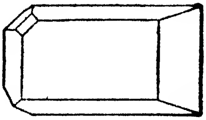


FIG. 20.—Crystal of Sucrose (Rock Sugar).

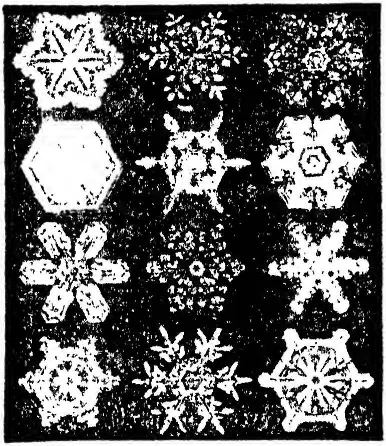


FIG. 21.—Snow Crystals.

Until recently chemists admired only the *exterior* of crystals, making accurate measurements of the angles at which different faces met so as to obtain exact figures for identification purposes. A whole science—crystallography—was built up as the result of this study. Now, however, it is recognized that a knowledge of the *interior* arrangement of the particles of which the crystal is composed is of still deeper import. In order to explain how this knowledge may be arrived at, we must go off on another tangent for a time.

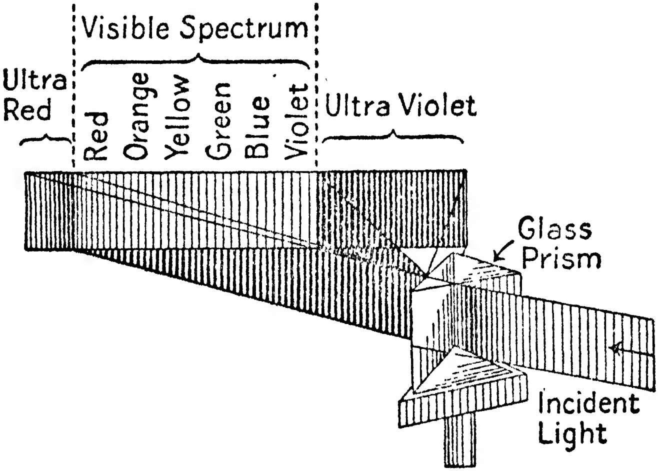


Fig. 22.—Analysis of White Light into a Spectrum.

Ordinary sunlight is considered to consist of vibrations of very small wave-lengths. When we break up a narrow beam of white light into a spectrum by means of a triangular prism (*Fig. 22*), it is found that the wave-length varies with the colour, the actual figures being about thirty thousand waves to an inch at the red end of the visible spectrum, and about sixty thousand waves to an inch at the violet end. Now waves of such very small lengths should, according to mathematicians, show a slight tendency to bend around corners, and such effects, indeed, may be easily observed. They are called, in general, diffraction phenomena.

If you look at a distant lamp through a fine linen handkerchief or through the fabric of a silk umbrella, you will notice a number of coloured images of the flame arranged in two directions at right angles to the weave of the material. These colours are due to the bending of the light around the narrow threads through which you are looking, the amount of bending varying with the colour. A method has been worked out, indeed, by which we can measure the exact wavelength of light from a definite source by the study of such diffraction spectra. Perhaps it will help if we describe a simple experiment which may be within your own powers of performance.

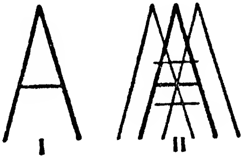


FIG. 23.—Diffraction Spectra.

Take a small pocket mirror and, with a sharp knife, scratch a long capital A through the coating on the back so that light can pass through your design as in the diagram (I, *Fig. 23*). Set the mirror up with the cross-bar of the A horizontal in front of a yellow flame, which you may conveniently obtain by powdering the wick of your cigarette-lighter, temporarily filled with alcohol, with common table salt. This is necessary in order to get what the physicist calls a " sodium flame," the light from which is almost all of the same wave-length (*Fig. 24*). Now fix a piece of fine wire gauze (a square from a mosquito screen will serve) some distance in front of, and parallel to, your mirror. Lastly, get your opera-glasses, see that they are in proper focus, and set them up so that you are looking through them and the gauze directly toward the centre of the cross-bar of the A and the flame behind it. You will see a direct image of the letter A, somewhat out of focus, and also several well-focused diffraction spectra. Fix the gauze so that one set of wires is vertical and, keeping everything else in place, move the gauze backward and forward until you get the first diffraction spectra intersecting each other exactly on the cross-bar of the A (II, *Fig. 23*). If you are clever enough to succeed in doing all this, then it will interest you to know that you can calculate the wave-length of sodium light from the simple formula :

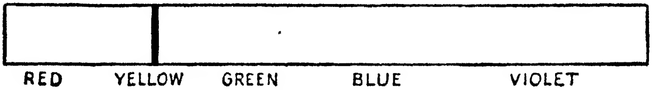


FIG. 24.—Sodium Flame Spectrum.

$$ld/2D.$$

l is the length of the cross-bar ; d is the distance between the wires of the gauze ; and D is the distance between the mirror and the gauze.

Of course, the experts make their measurements in a much more elaborate way, and in order to get very accurate values they use instead of a wire screen what are known as *diffraction gratings*, which consist of parallel lines ruled as closely together as possible on a glass plate. By most ingenious devices, gratings have been made containing as many as fifty thousand lines to the inch. Such gratings are admirable for the exact analysis of ordinary visible light, since the distance between the lines on the grating is about the same as the wave-length of the light, but they are quite useless in the study of X-rays. Later on you will learn more about X-rays; here we shall interrupt our story only for a moment to tell you that they are fundamentally similar to light rays in their nature. The wave-lengths of X-rays, however, are only about one-thousandth part of those of ordinary light, and, of course, we cannot rule lines —and no machinery has ever been devised which will rule lines for us—so close together as would be necessary for their examination.

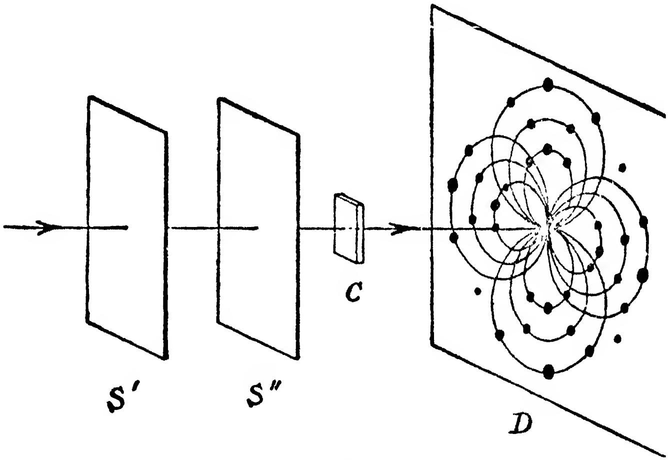


Fig. 25.—Laue Method for the Determination of Crystal Structure.

The beam of X-rays, after passing through pin-holes in the lead sheets S′ and S″, is broken up by the crystal C into a diffraction pattern, registered on the photographic plate at D.

This is where we come back to the crystal. It was discovered by Professor Laue of the University of Zurich, in 1912, that crystals supply us with *natural gratings* for the study of X-rays, the particles which make up the crystal being arranged in rows and the distance between these rows being approximately equal to the wave-length of X-rays. Employing X-rays of known wave-length, therefore, it is possible for us to work out the complete internal structure of a crystal.

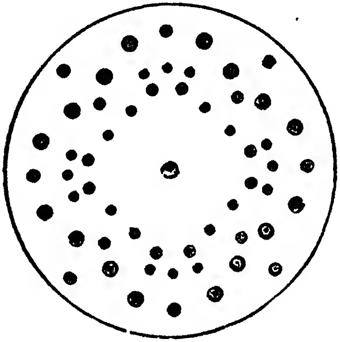


FIG. 26.—A Laue Photograph of Rock Salt.

Laue's method for doing this consists in directing a narrow beam of X-rays through a thin film of the crystal, set up as shown in the diagram (*Fig. 25*), and then allowing the rays to fall on a photographic plate, where a diffraction pattern of the crystal lattice is obtained. A typical Laue photograph given in this way by a crystal of rock salt (sodium chloride) is shown in *Fig. 26*. Each spot in this picture represents the reflection of the X-rays by a certain plane in the crystal lattice, and, while the patterns are sometimes very intricate (since the grating is in three dimensions), we can deduce from the positions of the spots the complete crystal structure.

Other methods have since been devised which lead to the same result more easily, but it will not be necessary for us to describe them here, since, no doubt, you will be impatient to know what the inside of a crystal actually looks like when an expert crystal-gazer like Professor Laue gets to work. A preliminary picture of the interior of a crystal of ordinary salt is shown in *Fig. 27A.* You will see that the particles which make up the crystal are arranged in such a way as to build up tiny cubes. So tiny are these cubes, indeed, that the side of each measures only about one 100-millionth of an inch. But what is still more astonishing is this: *There are no molecules of NaCl in a crystal of sodium chloride!* By setting the crystal in different positions, the professor discovers that diagonal planes are alternately all sodium and all chlorine, and our second picture of the crystal (*Fig. 27B*) makes it apparent that the ultimate particles which build up the unit cubes are alternately sodium *atoms* and chlorine *atoms.*

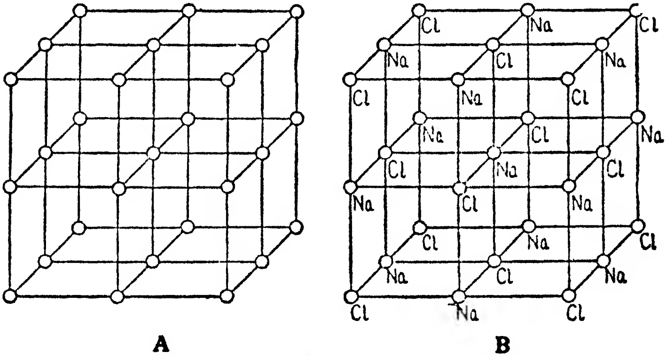


Fig. 27.—Interior Structure of a Crystal of Common Salt (Preliminary Stages).

In crystals of metallic copper the atoms are arranged in a space lattice which is a little more complex, but still cubical in type, as you may see by referring to the model opposite. This model, you must understand, does *not* represent an actual photograph. It is scarcely necessary to state that the wires joining the atoms are not present in the crystal itself, but you do need to be cautioned not to consider the atoms as really little hard spheres such as are shown in the picture. The actual " structure " of the atom itself will be considered in a later chapter.

At the present moment we have something very important to deduce from our crystal of copper—namely, the *weight of a single atom*. We can obtain this by a very simple calculation. From the way in which the atoms are arranged and from the size of the unit cube, we discover that there are no fewer than 85,000,000,000,000,000,000,000 atoms of copper in one cubic centimetre. Now one cubic centimetre of copper weighs a little less than nine grammes. Therefore, there are nearly ten thousand trillion (10,000,000,000,000,000,000,000) atoms of copper in one gramme ! More significant still is the number of atoms in the *gramme atomic weight* of copper, since this number should be the same for all elements. The atomic weight of copper is 63·5, so that one gramme atomic weight represents six hundred thousand trillion atoms. We obtain this same number by analysing the crystal lattices of other elements, and it agrees also within error limits with values that had been deduced by chemists much earlier, but indirectly, with the assistance of the kinetic-molecular hypothesis. Consequently, we feel quite certain that all the steps in our argument are valid, and that we have at last tracked the wily atom to its lair.

You do not need to tell us that you are surprised, dear reader ; we have no doubt, indeed, that you are overwhelmed. Such numbers as the above are beyond all ordinary experience. However, let us try to make their magnitude more comprehensible to you by one or two illustrations. It is possible that these illustrations may not be of much more help than a statement like "If all of the seats in the Wembley Stadium were placed end to end, a large number of the spectators would not be able to see anything of the game," but they will be better than nothing at all.

Let us assume, first, that you have become sufficiently suspicious regarding the accuracy of the figure which we have given above for the number of atoms in 1 gramme of copper to undertake the task of counting them for yourself, one by one. The services of some magician would be necessary to expand the metal to a point where you would be able to distinguish the atoms individually, and if he should enlarge each atom to the size of a grain of sand (about 1 cubic millimetre) your original gramme of copper would cover the British Isles to a depth of about 40 metres (over 130 feet). Working alone, and counting at the rate of one atom per second, you would not make very rapid progress upon this enormous mass, so let us suppose that you could call upon the total population of the British Isles (say 50 million people) to assist you. How long do you think it would take to complete the atomic census under these conditions? The answer is approximately six million years, Sundays and holidays included.

Our second illustration is, perhaps, even more striking. There are fifty thousand million atoms of gold in a drop of sea-water! Do not get inordinately excited over this piece of news; it is really not necessary for you to dash off to the seashore at once with a bucket in order to be ahead of the rush. When you make the required calculations, you will find that the stupendous figure given above actually corresponds to only 0·0005 grammes of gold in a ton of sea-water, and you would therefore have to extract all of the gold from about sixty tons of the ocean to obtain one pennyworth of the precious metal. The cost of such extraction would considerably exceed one penny, which explains why the gold continues to stay where it is. The rejected wash-water from the gold mines in South Africa contains much more gold than sea-water does. Consequently, you will understand why, although companies are floated at intervals for the purpose of extracting gold from sea-water, the only gold ever obtained in appreciable quantities from such ventures is extracted from the innocents who buy the stock.

Now you may come back to crystal-gazing again. The atoms in crystals are not arranged in a simple cubic lattice for all substances; frequently the pattern is much more complex. In magnesium the atoms fall in a hexagonal lattice, as indicated in the lower picture facing page 108. Probably you are not terribly interested in magnesium, however, and you would find it much more alluring to gaze into the interior of a diamond. Concentrate your attention upon a diamond and keep on looking at it very intently. If you are of a romantic disposition, of course, your mind may wander away from the subject of atoms completely, but if you could continue to concentrate seriously until you saw the individual atoms the result would look something like *Fig. 28*.

There is still one point to be discussed before we conclude this chapter, and in order to approach it easily we must return to our crystal of common salt once more. It is evident that, when elements combine, although the original properties of the elements disappear completely, the *atoms* do *not* lose their identities. They do not even undergo actual union, if we understand that word in the sense of merging together or coalescing; the combination consists merely in an orderly atomic rearrangement which is, of course, something quite different from the crude intermingling which we obtain in a mixture. The question does arise, however, for a substance such as common salt, " What keeps the individual atoms of sodium and of chlorine in a definite and rigid pattern ? "

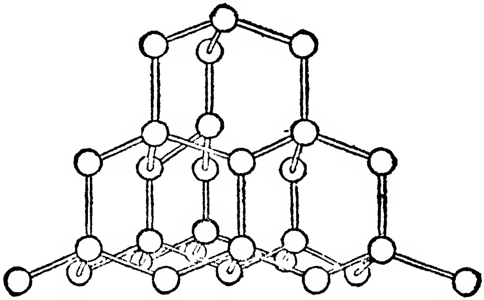


FIG. 28.—Ultimate Structure of a Diamond.

The answer to this question may be partially presented at this point, leaving a more detailed discussion until later. It has been found that the atoms in the crystal lattice of sodium chloride are *electrically charged*, the sodium atoms being charged positively and the chlorine atoms being charged negatively. Atoms, or groups of atoms, that are thus charged are called in chemistry " ions," so that, strictly speaking, the sodium chloride crystal lattice is really built up not of atoms of sodium and of chlorine, but of *sodium ions* Na^+ and *chlorine ions* Cl^- (*Fig.* 29). Now oppositely charged particles attract one another very strongly, and if we fix our attention upon one particular sodium ion (for example, the one in the middle of *Fig.* 29), we see that, although it is not *specifically* combined with any particular chlorine ion, yet its closest neighbours are six chlorine ions which surround it on every side and imprison it in its central position by their attractive forces.

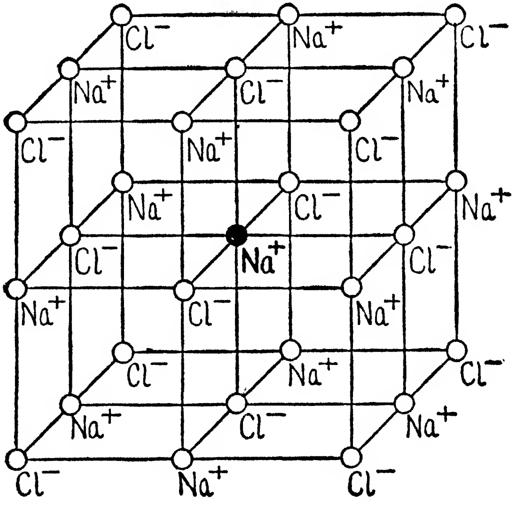
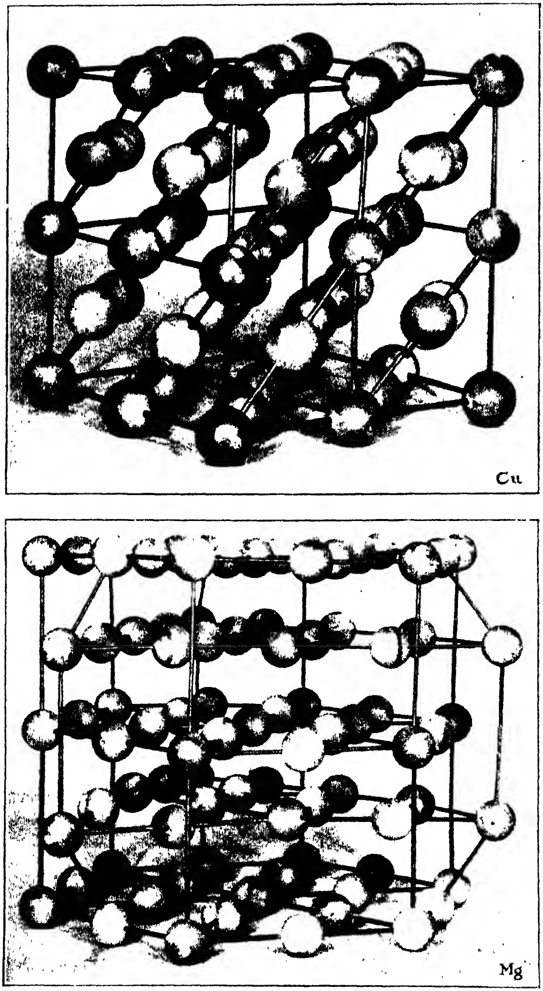


FIG. 29.—Ultimate Structure of a Crystal of Common Salt.

How do the atoms obtain these electrical charges? This is where we begin to improve upon Dalton's original theory. As a result of recent progress in our science, we do not regard the atom any longer as the ultimate limit of divisibility; *the atom is itself divisible.* For



ARRANGEMENT OF ATOMS IN COPPER AND MAGNESIUM

each element, in fact, a definite atomic structure has been worked out which increases in complexity with the atomic weight. In general it is a very difficult matter indeed to pick an atom to pieces, but almost every species of atom contains, as its outermost shell, one or more *electrons*, or units of negative electricity, which are relatively loosely held. Only these loose electrons are affected by ordinary chemical reactions, the rest of the atom stays fixed. Atoms of metallic elements, as a rule, tend to give away their loose electrons ; such atoms evidently have the characteristics of spendthrift males. Atoms of non-metals, on the other hand, generally prefer to acquire additional electrons ; they represent the receptive females. Consequently, when sodium combines with chlorine to form sodium chloride, all that the marriage rite actually involves is the transfer of one electron from each sodium atom to each chlorine atom. Each sodium atom, being originally *electrically neutral* (that is, containing just as many positive as negative units of electricity), and having *lost a negative charge*, is thereby converted into a sodium ion Na^+ with an extra positive charge, while each chlorine atom becomes a chlorine ion Cl^- with an extra negative charge.

We should like to describe the whole wedding service to you in full detail now, but you will be in a better position to appreciate the ceremony after the next few chapters have made you more intimately acquainted with the principal actors in it.

Chapter IX

HYDROGEN: "NICE BABY"

> "What's the use of a baby? Some day it will grow up!"
> —MICHAEL FARADAY.

THE greatest of all British chemists was very fond of repeating the above retort of Benjamin Franklin to those of his friends who reproached him for wasting so much time on foolish experiments with "lightning." Faraday's chemical babies certainly have grown up, as may be appreciated by a second extract from the address given by the Prince of Wales at Oxford in 1926 in his capacity as president of the British Association for the Advancement of Science:

> Faraday's labours provide one of the most wonderful examples of scientific research leading to enormous industrial development. Upon his discovery of benzene and its structure the great chemical industries of to-day are largely based, including, in particular, the dyeing industries. Still wider applications have followed upon his discovery of the laws of electrolysis and of the mechanical generation of electricity. It has been said, with reason, that the two million workers in Great Britain only who are dependent upon electrical industries are living on the brain of Faraday; but to his discoveries in the first instance many millions more owe the uses of electricity in lighting, traction, communication, and industrial power.

In this chapter we are going to describe the birth and early life of the infant among the chemical elements, hydrogen. We shall carry it safely through its childhood days and show you that it is now sufficiently grown up to be put to work.

The father of hydrogen was a most eccentric member of our English aristocracy, the Honourable Henry Cavendish. It is true that, way back in the days of alchemy, Paracelsus had noticed that an inflammable gas was given off when a metal like iron was attacked by an acid like vinegar. For centuries, however, nobody distinguished this particular gas from other varieties of "inflammable air." Cavendish did so in 1766, and observed that not only did the gas burn but that a mixture of it with air exploded violently when a light was applied. Priestley heard of Cavendish's experiments and used to amuse his friends by exploding such mixtures for them. He might easily, indeed, have gained the glory of being the "discoverer" of the element, since on one such occasion his attention was drawn by a Mr. John Waltire to the fact that the inside of the glass globe in which the explosion occurred was covered with dew at the end of the experiment, although it was quite dry at the beginning. Priestley, unfortunately, did not follow up this suggestion, but in the year 1781 Cavendish, who had also noted this "condensed dew," repeatedly exploded mixtures of "inflammable air" (hydrogen) and "dephlogisticated air" (oxygen) in the same vessel until he had a sufficient quantity of liquid to test its properties. He found that it looked like water; it smelled like water; it tasted like water; it boiled like water; it froze like water; in fact, it was water!

This discovery put Cavendish in an awful predicament. Like Priestley, he was an earnest apostle of the phlogiston theory. Since he had obtained his very light "inflammable air" by the action of an acid on a metal, he had become convinced that this buoyant gas was the long-sought mythical *phlogiston*, which had been expelled from the metal by the acid! The only logical conclusion, according to this idea, to be drawn from his latest experiment was, of course:

Phlogiston + Air deprived of phlogiston → Air!

Cavendish now found himself compelled, consequently, inasmuch as water and *not* air was the product of his explosion, to change his mind with regard to the nature of one or both of his original gases. He could not, however, find it in his heart to abandon the phlogiston theory; and even after the correct explanation of his results had been given by Lavoisier, Cavendish refused to accept it. Like Priestley again, he went to his grave without being converted to the new beliefs which his own work had inspired. As we have stated, Cavendish was a *most* eccentric person. Apart from a few scientific friends, he had no associations with his fellow-men. He wore the costume of his grandfathers, and the only portrait of him which exists was drawn by stealth. He lived like a hermit, and was annoyed when he had to speak with people, or even pass them. There is a story that, having met one of his maids by chance while descending the stairs of his home, he had a back staircase specially constructed to avoid such encounters in the future. He actually quarrelled with his bankers for worrying him by asking what they should do with his funds. Needless to say, he was never married. When he died it was discovered that his fortune amounted to £1,500,000, and one of his biographers has aptly styled him "the wealthiest of all the wise and probably also the wisest of all the wealthy."

Now let us see what Lavoisier did for his foster-child. First of all, repeating the work of Cavendish, Lavoisier immediately realized its true significance. In 1783 he proved the constitution of water, not only by building it up directly from its elements as Cavendish had unconsciously done, but also by the reverse process of decomposing it back again into those same elements. Lavoisier did this by passing a current of steam through a glass tube containing red-hot iron wire. The iron was converted into the black magnetic oxide, and a gas was collected from the end of the tube which proved to be identical with Cavendish's "inflammable air." In view of the fact that the most important property of this "inflammable air" was its combustion with oxygen to form water, Lavoisier took the liberty of christening his adopted baby *hydrogen*, from two Greek words meaning *water-producer* (compare our word hydrant).

It will repay us at this point to study Lavoisier's last experiment a little more closely. Although water was decomposed in the course of this experiment, yet the reaction is not a simple decomposition like those we have discussed in earlier chapters. Lavoisier started with *two* substances, iron and water; he finished with *two* substances, iron oxide Fe_3O_4 and hydrogen. The balanced equation runs as follows:

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2.$$

Looking at this equation, we see that what has really happened is that the iron has *displaced* the hydrogen from its combination with oxygen, turning it loose and taking the oxygen under its own protection. This type of reaction is accordingly called *displacement*.

When a metal is attacked by an acid and hydrogen is set free, the reaction is also a displacement. For example, if we take some pieces of scrap zinc in a flask and pour hydrochloric acid on them, a rapid stream of bubbles soon begins to escape. If we collect some of the gas carefully in a dry tube and put a match to it, it burns with a very light blue flame, and the appearance of a faint mist on the inside of the tube shows that water has been formed. When the reaction has ended, if we take some of the clear liquid in the flask and evaporate it to dryness, we find that we have a residue which looks something like salt and which analysis shows to consist of zinc chloride $ZnCl_2$. The equation for the reaction may therefore be written:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2.$$

In the same way we might use iron with sulphuric acid, in which case the solid product remaining would be light-green ferrous sulphate $FeSO_4$:

$$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2.$$

With aluminium and hydrochloric acid the solid product is aluminium chloride $AlCl_3$:

$$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2.$$

Metals like zinc, iron, and aluminium react with acids in this way very rapidly, releasing hydrogen and forming substances which, because of their general similarity to common salt, are called "salts." With other metals such as tin and lead, the action proceeds much more slowly, and some metals, such as silver and gold, will not displace hydrogen from cold dilute acids at all.

On the other hand, a few metals are so exceedingly active that they will displace hydrogen even from cold water. The metals discovered by Davy—potassium, sodium, and calcium—fall into this class. In the case of potassium, which floats on water, the reaction is so violent that the hydrogen catches fire as soon as it is released, and with sodium also the risk of an explosion is so great that the experiment is too dangerous for the novice to attempt. Calcium, however, sinks to the bottom of the water and acts much less vigorously, so that the hydrogen coming off may easily be collected (*Fig. 30*). The solid products obtained here are the hydroxides of the metals, compounds which we have encountered before under the name of alkalis or bases. The equations are:

$$2K + 2H_2O \rightarrow 2KOH + H_2$$
$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$
$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2.$$

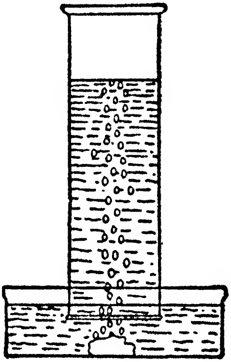


FIG. 30.—Action of Calcium on Water.

Aluminium and magnesium, like iron, decompose water rapidly only when it is supplied to them at a high temperature in the form of steam. The reason for all these differences in "activity" on the part of the metals will be discussed later.

We can obtain hydrogen from water in another way—namely, by electrolysis. This is the technical term employed to signify a decomposition carried out by means of the electric current. Water itself does not conduct the electric current sufficiently well, but if we add to water a few drops of an acid, like sulphuric acid, or a few crystals of a salt, like sodium chloride, then we find that the current passes through the solution very much more readily. You can try this out for yourself, either using an ordinary storage or automobile battery as a source of current, or plugging in on the electric light circuit, as shown in *Fig. 31*. You will find that bubbles of hydrogen immediately begin to appear on the wire connected to the negative pole of the battery,

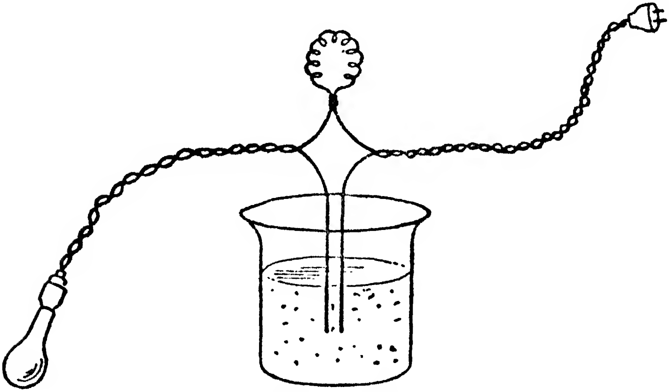


FIG. 31.—Electrolysis of Water.

and with patience you will be able to collect sufficient gas in a tube to burn it and recognize that it is hydrogen.

The hydrogen used for commercial purposes is manufactured in several ways. Electrolysis, in general, is too expensive, but a certain amount of hydrogen is procured as a by-product in other electrolytic preparations. A great deal is made by the method of passing steam over heated iron. The largest quantity, however, is obtained by starting with what is known as the water-gas reaction. Steam is blown over white-hot coke or anthracite with the following result:

$$C + H_2O \rightarrow CO + H_2.$$

The mixture of carbon monoxide and hydrogen (water-gas) which comes off is combustible, and is often used as a cheap source of heat and power and in the manufacture of illuminating gas. The separation of pure hydrogen from water-gas, however, is a matter of some difficulty, and we shall not attempt to describe the various steps to you in detail in this volume.

What is industrial hydrogen used for? Until recently, apart from small amounts employed for filling balloons and dirigibles, the bulk went (in its mixtures with other combustible gases, such as water-gas) into light, heat, and power plants, but other important applications have lately been devised. Enormous quantities of hydrogen are now consumed yearly in the conversion of cottonseed oil and other vegetable and marine oils, which formerly went to waste, into appetizing solid fats and into products useful in the soap and candle industries. The reaction is made to take place by forcing hydrogen through the heated oils in the presence of a finely divided metal such as nickel, and represents a case of direct combination. Another very extensive use of hydrogen which has recently been developed is in the manufacture of synthetic ammonia, NH_3. This compound is a source of many valuable fertilizers in time of peace and of many valuable explosives in time of war. Both of these new processes have now progressed far beyond the point where they can continue to be regarded as "infant industries." Our baby element is growing up!

It is growing up so rapidly, indeed, that during the last year or two it has for the first time been given the responsibility of going out to work quite alone. Under ordinary conditions, hydrogen gas consists of molecules with the formula H_2, but Dr. Irving Langmuir of the General Electric Company has shown that molecular hydrogen at an exceedingly high temperature is largely decomposed into atomic hydrogen H. When these single atoms recombine, an enormous amount of heat is liberated, and this fact has now been put to use in the hydrogen arc welding torch. In this torch a jet of hydrogen gas passes through an electric arc and is converted into atomic hydrogen. Immediately beyond the arc the atoms recombine, and the flame which they form in so doing is very much hotter even than that of the oxyhydrogen or oxyacetylene torch. The atomic hydrogen torch may thus be employed very advantageously in making welds of particularly refractory metals.

Still, in comparison with oxygen, hydrogen is rather a lazy element. It combines with only a few elements directly, such as oxygen and chlorine among the nonmetals and sodium and potassium among the metals. Compounds of hydrogen which are formed indirectly. however, are most numerous and useful. Hundreds of different *hydrocarbons* (as its compounds with carbon are called) may be isolated from natural gases and oils, increasing steadily in complexity from marsh gas or methane CH_4 through gasolene, naphtha, kerosene, and lubricating oils to solid paraffin and asphalt. All of these, as well as many other synthetic hydrocarbons of still greater importance, must be left for treatment in a later volume.

Chapter X

VALENCIA

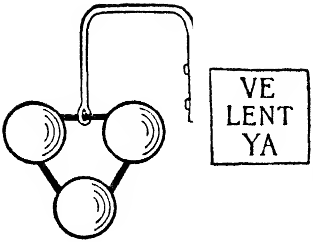


FIG. 32.—Ozone, O_3, according to Crum Brown (modified).

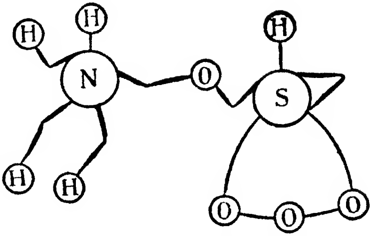


FIG. 33.—Ammonium Hydrogen Sulphate, NH_4HSO_4, according to Frankland.

I've tried hard, but vainly, to realize plainly
    Those bonds of atomic connection
Which Crum Brown's clear vision discerns with precision
    Projecting in every direction ;
Though Frankland's notation commands admiration,
    As something exceedingly clever,
And Mr. Kay Shuttleworth praises its subtle worth,
    I give it up sadly for ever.

—J. C. BROUGH, *The B-Hive*, 1868.

EVERY textbook of chemistry contains, just about at this point, a dreadfully stodgy chapter entitled " Valence." Nobody appears to know exactly what to do to improve the presentation ; students have to get valence, but they never like it. Last year, however (if I may be pardoned a personal digression), I chanced to look through a copy of a translation of one of my own textbooks into Spanish and my eyes fell on a chapter headed "Valencia." It was astonishing what a different attitude of approach the mere change of title induced. I tried humming the chapter through to the tune of "Valencia," and it did not seem nearly so stiff as before. Our antipathy toward the topic of valence, then, is mainly psychological, and can be overcome if proper measures are taken. Fortunately, in fact, recent developments in our ideas regarding the atom have now brought the theme of valence more harmoniously into the general concert of modern chemistry than was the case in the old days, when it functioned merely as an isolated classical number.

You will probably have been badly puzzled by some of the formulas that we inflicted upon you in the last chapter. "Why," you might well object, "should sodium chloride be $NaCl$, zinc chloride $ZnCl_2$, and aluminium chloride $AlCl_3$? Is there any rhyme or reason in this variation? Suppose I were suddenly called upon to state the formula for the chloride of another metal, such as magnesium, how could I tell how many chlorines to take?"

The answers to all of these questions will emerge in due time. The first and most obvious point for you to note is this. One atom of certain elements, such as sodium, combines[1] with only one atom of chlorine. Referring back to the displacement reactions in Chapter IX, you will find that one atom of any of these same

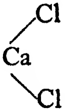
---

[1] The word *combines* is used here in a statistical sense only. As we saw in an earlier chapter, no sodium atom in a crystal of sodium chloride is actually united with any single specific chlorine atom. Even a crystal of microscopic size, however, contains myriads of sodium ions and chlorine ions in equal numbers, and, for convenience, we shall continue to express this fact by the empirical formula $NaCl$.

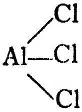
elements also displaces only one atom of hydrogen from combination. One atom of certain other elements, however, such as zinc, combines with two atoms of chlorine and displaces two atoms of hydrogen. For still other elements, such as aluminium, the number is three in each case.

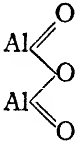
In order to start with a standard, the chemist calls the number of atoms of hydrogen or of chlorine which one atom of a given element combines with or displaces the *valence* of the element. Thus the valence of sodium is one; of zinc, two; and of aluminium, three. Frequently, in writing formulas, valence is marked by means of lines, the number of lines pointing toward a symbol indicating the valence of the atom for which it stands. Thus:

Na—Cl



Ca = O





This procedure, though useful in a way, has certain dangers, since the tendency naturally arises to regard such a "graphic formula" as giving an actual picture of the arrangement of the atoms in the molecule. This, of course, no representation on paper can possibly do, since the molecule is a three-dimensional affair, and you have already seen in your course on crystal-gazing that the picture of sodium chloride is much more complex than plain Na–Cl. But it is only human nature to try to connect an abstract idea with something tangible, and, although chemists have known quite well all along that the lines were not there, they often used to speak —and sometimes do so still—of "bonds" or "hooks" connecting the various atoms in a molecule. Another favourite analogy in this connection was that of "holding hands." On this basis we might imagine the water molecule H_2O (where one working-girl is in sole charge of two small children) to be something as follows (*Fig. 34*):

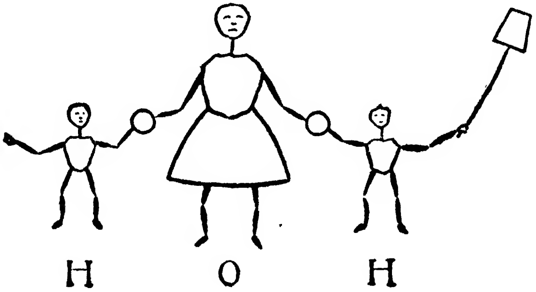


Fig. 34.—Water, HOH.

Or an atom of sodium getting "hooked up" with an atom of chlorine might be pictured thus (*Fig. 35*):

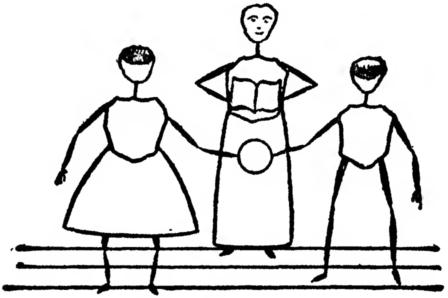


Fig. 35.—Sodium Chloride, NaCl.

Perhaps you feel that we are becoming too frivolous, but let us refer you to the pictures and verses at the beginning of the chapter, and you will see that we are sinning in good company. Evidently even the leaders among chemists in mid-Victorian times felt the need of occasional relaxation. Still, we admit your argument that we have got to approach the subject solemnly sooner or later. This hooking-up of sodium and of chlorine may tickle our elementary sense of humour temporarily, but it is a serious step for the atoms concerned in making up the combination.

That it is a very serious step indeed, at least in the three most important classes of compounds with which we have to deal in this volume—namely, acids, bases, and salts—is evident from the fact that the character of the individual atoms is entirely changed after combination has occurred. In the free state all atoms are electrically neutral, but in compounds of the three classes mentioned above we can always divide the formula into a positive part and a negative part. The justification for such a division may be very simply shown in an experimental way by dissolving an acid, base, or salt in water and passing an electric current through the solution. The part of the compound which is positively charged, such as hydrogen in H_2SO_4 or sodium in NaCl, is attracted to the negative pole (or cathode), while the part of the compound which is negatively charged migrates in the opposite direction toward the positive pole (or anode). Because of this behaviour, such compounds are frequently styled *polar*. When each part reaches its goal it may be that it loses its charge and is released in the free state, as happened in the case of hydrogen in the last chapter, or it may be that some other reaction occurs which complicates matters. Cases of this second type will come up later.

The positive part (or, as it is often called, the positive *radical*) of a compound is almost always a metal. Hydrogen, it will be noticed, is considered to fall into the class of metallic elements according to this scheme. Other elements commonly regarded as non-metallic may enter into a positive radical, but only in groups which imitate metals in their properties, such as the ammonium group NH_4. The negative radical of a compound is sometimes a single non-metal and sometimes a group of non-metals, such as the sulphate group SO_4. The valence of the positive part must necessarily be equal to the valence of the negative part, and you will find this simple rule of great service to you in writing down formulas correctly. Thus, for example, you can work out the formula of aluminium sulphate (a salt used extensively in baking-powders, in paper-making, in water purification, and in dyeing) from what you know already in the following way. The valence of aluminium is three, since one atom of aluminium combines with three atoms of chlorine and displaces three atoms of hydrogen. The valence of the sulphate radical is two, since it is combined with two atoms of hydrogen in sulphuric acid H_2SO_4. The simplest possible formula for aluminium sulphate is, consequently, $Al_2(SO_4)_3$. We write $(SO_4)_3$ in this formula rather than S_3O_{12} in order to emphasize the continued presence of the SO_4 group, just as in the formula for slaked lime $Ca(OH)_2$ we preserved the identity of the hydroxyl group.

Now let us retrace our steps a little and consider how the two radicals acquire their positive and negative charges. We stated in an earlier chapter that the combination of sodium with chlorine to form sodium chloride involved merely the transfer of one electron from each sodium atom to each chlorine atom. We did not, however, look into the origin of this mysterious electron in any detail.

Not so very long ago it used to be regarded as rather a breach of etiquette to pry too closely into the nature of electricity. When pressed to give an opinion on the subject, the expert would concede to the curious inquirer that there was an elusive " something " which flowed along a wire from the positive pole to the negative pole of a battery when the two poles were connected, but that was about as far as he would go. Perhaps he was wise in refusing to incriminate himself further, because we know now that he was all wrong. The flow is from the negative pole to the positive pole. Furthermore, we have found out exactly what it is that flows—it is a stream of electrons. These electrons may be most conveniently isolated and studied by means of the *vacuum tube*, devised by Sir William Crookes. If we take a tube fitted up as shown in *Fig. 36*, pump all of the gas out of it that we can, and then discharge a high-voltage current through the tube, we find that " rays " shoot off from the negative electrode, or

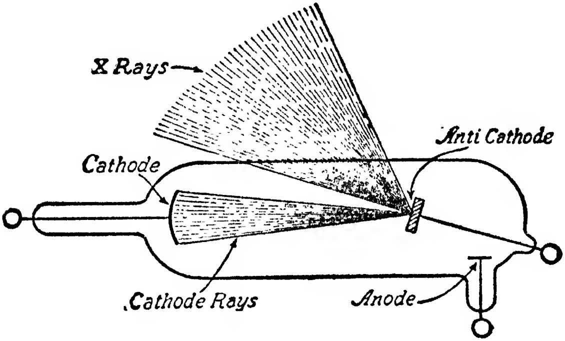


FIG. 36.—Crookes' Vacuum Tube.

cathode, which possess extraordinary properties. They cause the small amount of residual gas to glow very brightly, the colour depending upon the nature of the gas with which the tube was originally filled. They are so powerful that their impact upon the glass would soon melt the tube if a heavy stream of the rays was allowed to strike upon it. In the vacuum tube the cathode rays are, therefore, focused so as to fall upon an anti-cathode of suitably resistant material. Their impact upon this anti-cathode induces another type of ray which causes the glass of the tube to emit a greenish-yellow fluorescent light. Röntgen in 1895 discovered accidentally that these reflected rays could pass through paper, flesh, and other materials composed of elements of low atomic weight, and we now are all familiar with them under the name of X-rays.

X-rays, as was stated in a previous chapter, are of the same nature as ordinary light, but possess a much shorter wave-length. The cathode rays, however (which are all that directly concern us at the present moment, the information about X-rays being merely incidental), are more tangible. They have been shown to consist of electrons, which possess a mass about one part in 1,850 that of the hydrogen atom. Looking back to Chapter VIII, you will discover that it takes six hundred thousand trillion hydrogen atoms to equal one gramme. Consequently, multiplying this number by 1,850, you may convince yourself that it would require more than one thousand quadrillion electrons to weigh one gramme. We hestitate to write this number out in full, since it would be very monotonous for you to wade through all the zeros.

Now it is these same almost imponderable electrons, or units of negative electricity, which pass between atoms when they combine. Why they should want to pass is a question which we shall leave for later discussion. At this point, however, we can proceed to consider how the new ideas which we have introduced help to explain the facts of electrolysis.

When we immerse two wires, each connected with one of the poles of a battery, in a water solution of a polar compound, such as sodium chloride, what happens is as follows. Because of chemical reactions occurring inside the battery, the negative wire or cathode is teeming with an excess of electrons, while the positive wire or anode has lost a lot of electrons and seeks to make up its deficiency. Of course, if the wire continued all through the solution, a steady stream of electrons would pass along it in order to equalize matters. In the absence of such a connecting wire, the positive and negative radicals of the salt do their best to assist. The positive radicals here are sodium atoms, each of which has lost an electron and has thereby become a sodium ion Na^+. The negative radicals are chlorine atoms, each of which has acquired an electron and has thereby become a chlorine ion Cl^-. In the water solution, even the companionate or group union between Na^+ and Cl^- existent in the crystal state is dissolved, and the two kinds of ions act quite independently. The Na^+ ions are attracted to the negative pole in the expectation of recovering there the electrons which they have lost, while the Cl^- ions migrate to the positive pole with the charitable purpose of unloading their extra electrons at the spot where they are urgently needed. Actually only the Cl^- ions succeed in fulfilling their mission, and chlorine gas Cl_2, formed by the union of the discharged atoms into pairs, bubbles off from the anode. At the cathode a different type of reaction occurs, but the failure of the Na^+ ions to achieve their original aim does not affect the truth of the story we have told in any essential way, as you will see later.

We can distinguish, then, two types of valence—*positive valence*, which represents the number of electrons that an atom (or group of atoms) loses, and *negative valence*, which represents the number of electrons that an atom (or group of atoms) gains. As we have stated once before, loss of electrons is characteristic of spendthrift males, or metallic elements. Gain of electrons, on the other hand, is characteristic of acquisitive females, or non-metallic elements. There is, however, a third type of valence, called *co-valence*, which we must not omit to examine.

You have already encountered a number of common compounds which do not fall into any of the three classes, acids, bases, or salts. A few examples may be mentioned to make this clear to you—for instance, methane CH_4, carbon monoxide CO, and carbon disulphide CS_2. These compounds do not conduct an electric current under any circumstances (hence they are styled *non-polar*), and there is no justification for splitting their formulas up into positive and negative parts. Why should these compounds differ so decidedly from those we have hitherto examined? If you look at their formulas you will notice that they are composed entirely of non-metallic elements. Now when a party of girls gets together there is bound to be some discussion as to who is going to be responsible for incidental expenses. A gentleman entertaining a lady friend almost always insists upon paying the bill, but in the absence of a gentleman each lady present is more likely to contribute her due share. All of this, translated into scientific language, means that in compounds like methane CH_4 there is no *transfer* of electrons from one atom to another; combination into a molecule consists instead in the *sharing* of electrons between atoms. The electrons are shared in pairs in every case, each of the atoms concerned generally supplying one of the pair. The co-valence of an element represents, then, the number of pairs of electrons which an atom of that element can share with atoms of other elements. In a later chapter we shall scrutinize the three kinds of valence—positive, negative, and co-valence—more intently, and we shall discover that the valence number in each case does not vary in a casual way from element to element, but follows a simple and definite scheme.

In the course of the above discussion, we have grouped compounds into two distinct classes, *polar compounds* which conduct the electric current in consequence of an electron transfer between their positive and negative radicals, and *non-polar compounds* which do not conduct the electric current, since electrons are merely shared. Strictly speaking, however, the distinction between polar and non-polar compounds is not fundamental, but one of degree only. No substance is a perfect non-conductor of electricity and no substance exhibits, under any practical conditions, its maximum possible conducting capacity. In fact, a list of compounds could be drawn up, starting with methane and ending with sodium chloride, in which there is a gradual and continuous transition from very poor conductance to almost perfect conductance. Even the same compound behaves differently according to the environment in which it finds itself.

Thus pure sulphuric acid H_2SO_4 is a liquid which is an extremely poor conductor. Pure water H_2O is also a liquid which is practically a non-conductor. A mixture of water and sulphuric acid, however, conducts the electric current excellently.

We can understand all this by recognizing the fact that an electron is not *forced* to pass over completely from one atom to another or else be shared equally by the two; it may choose to follow an intermediate course. In no non-polar compound, probably, are pairs of electrons held in common between two atoms so impartially that they will not tend, to some (even though very small) extent, to associate themselves with one atom rather than with the other. In the same way, in no polar compound does an electron pass from one atom to another irrevocably and irretrievably; the attractive forces induced between neighbouring and oppositely charged atoms must tend, to some (even though very small) extent, to pull it back occasionally into a straddling position.

After all, we might expect this also from our everyday experience. Even although a lady is sure that her companion on a projected outing is a perfect gentleman, yet she does not forget to take her purse with her, since she knows that she may occasionally find some small change useful. In the same way, even although two girls may plan to share the expenses of an excursion equally, yet it would be strange indeed if one should actually spend in the course of the trip *exactly* the same amount as the other.

This is one of many instances in chemistry where we find it convenient to classify substances or reactions into various groups, without being able to draw any definite boundaries between the groups. Our classification of elements into metals and non-metals is similarly imperfect, one group merging gradually into the other. The fact that the boundary is not strictly defined does not impair the usefulness of the classification. An illustration from outside the domain of chemistry may help to make this point clear. If we were asked to describe the Continent of Europe, we should certainly begin by naming various countries, such as France and Germany. Everybody has a sufficiently definite idea as to what these names denote to make them useful (in fact, practically necessary) in any detailed description, whatever standpoint we adopt. Nevertheless, as soon as we attempt to establish a logical boundary between France and Germany, we find ourselves confronted with a difficulty which has baffled mankind for centuries! There is no indisputable line where French predominance logically ends and German predominance logically begins; the inhabitants of the border zone merge gradually and almost indistinguishably from one race and language to the other. It would be foolish, merely because of this fact, to deny the existence of France and of Germany as separate countries with distinct characteristics. In the same way, no one would dream of abandoning the convenient division of the year into the four seasons just because we get some winter days in spring, or vice versa. Bearing all this in mind, we see that the chemist is perfectly justified in using terms, such as metals and non-metals, polar compounds and non-polar compounds, as an aid in classifying the science, even although he is unable to assign every single case definitely into one class or the other.

A final difficulty in connection with valence may be more easily disposed of. The source of this difficulty lies in the fact that two elements may form more than one compound, as we saw when we discussed the law of multiple proportions. Now, if the valences of hydrogen, oxygen, and carbon are one, two, and four respectively, as they are generally considered to be in accordance with the formulas of their most common compounds, water H_2O, methane CH_4, and carbon dioxide CO_2, what course are we to take in dealing with other compounds of these same elements, such as hydrogen peroxide H_2O_2, ethylene C_2H_4, acetylene C_2H_2, and carbon monoxide CO? In some cases we find it possible to write the formula for such a substance consistently without any change in valence numbers. The properties of hydrogen peroxide, for example, are such as to justify the accompanying picture (*Fig. 37*):

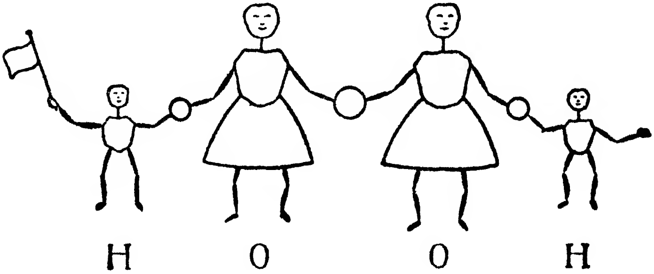
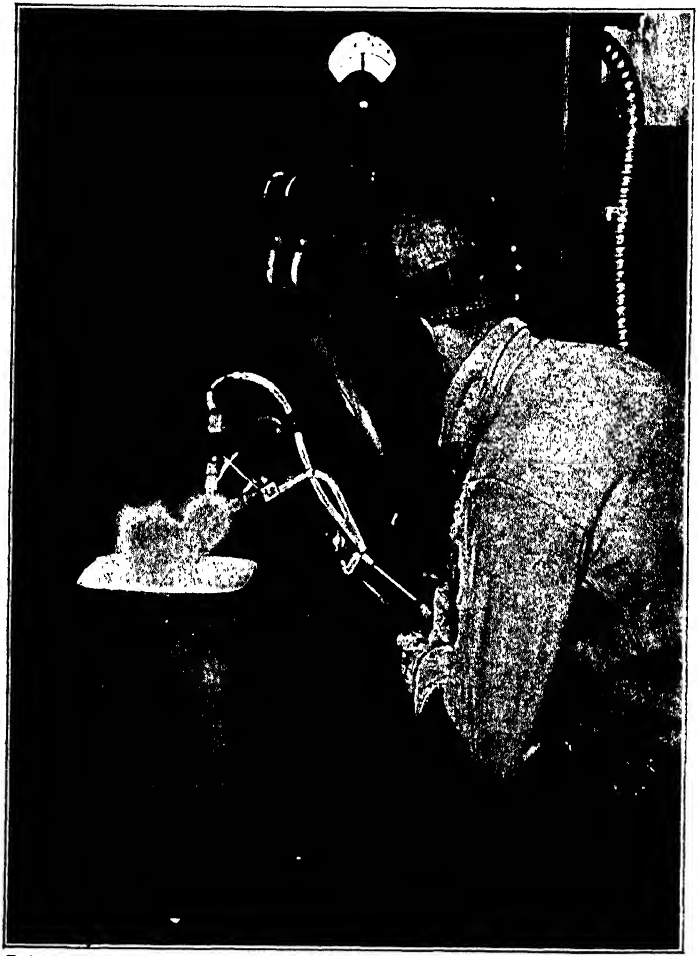


FIG. 37.—Hydrogen Peroxide, H_2O_2.

In the same way, acetylene is usually written in graphic form as follows: $H — C \equiv C — H$.

Frequently, however, such an obvious way out of our embarrassment cannot be discovered. Thus phosphorus is generally regarded as possessing a valence of five, since it gives compounds like PCl_5 and P_2O_5, but phosphorus



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ATOMIC-HYDROGEN ARC WELDING

Workman making a Refrigerator Evaporator (see p. 122)

© *General Electric Co.*

MODERN CATHODE RAY TUBE (COOLIDGE TUBE)

can also form compounds in which it exhibits a valence of three, like PH_3 and PCl_3. In the same way, tin forms two compounds with chlorine, a liquid $SnCl_4$ and a solid $SnCl_2$.

Do not let this confuse you. After all, even although it is *possible* for an atom of phosphorus to combine with five atoms of chlorine, this fact does not *compel* phosphorus to exercise its maximum combining capacity in every single one of its compounds. To continue with the analogy that we have used before, while an atom of phosphorus may, under certain circumstances, be willing to pool all of its worldly wealth with five atoms of chlorine, yet it may feel much more dubious over trusting its last coppers with five atoms of hydrogen. Children have a trick of losing coins, so the phosphorus atom prudently decides to keep back part of its cash for use in case of emergency. Similarly we can explain the case of tin and chlorine by remembering that a man with four pounds in his pocket does not have to hand all of it over to his wife when she asks him for money. If he is in a warm-hearted mood or if she uses very concentrated arguments, he may do so, but ordinarily he might quite properly prefer to keep two pounds for himself. The course which he actually takes depends on conditions, and we find, indeed, that it is this matter of *conditions* (temperature, concentration, etc.) which decides whether tin and chlorine will form $SnCl_4$ or $SnCl_2$, and whether phosphorus and chlorine will form PCl_5 or PCl_3. When carbon burns in excess oxygen, it gives CO_2; when the oxygen supply is deficient, it gives CO. Iron rusts at a low temperature to form Fe_2O_3, and at a high temperature to form Fe_3O_4.

In order to distinguish clearly between different compounds of the same elements, some systematic plan of naming such compounds is obviously needed. Unfortunately, while efforts have been made in this direction, they cannot be regarded as entirely satisfactory. Sometimes a numerical prefix is employed, as in carbon *mon*oxide CO, carbon *di*oxide CO_2. Sometimes the terminations *ous* and *ic* are used to indicate lower and higher states of combination respectively, such as ferr*ous* oxide FeO, ferr*ic* oxide Fe_2O_3. This latter plan is most commonly followed in the case of compounds containing a metal, while the former procedure is generally adopted with compounds containing only non-metals. There is no real consistency observed, however, and the terminations *ous* and *ic* are also very conveniently applied to denote acids formed from lower and higher oxides of non-metals. Thus sulphurous acid H_2SO_3 is derived from sulphur dioxide SO_2; sulphuric acid H_2SO_4 from sulphur trioxide SO_3. It must be remembered that names are usually given to compounds immediately after their discovery, and often the name sticks in spite of a change in our views regarding the nature of the substance. If we could wipe our chemical slate clean and start all over again, we could be much more logical in our nomenclature than we are at present. Tin and lead are similar metals which give similar compounds, yet $SnCl_2$ is called stann*ous* chloride, while $PbCl_2$ is called plumb*ic* chloride! Were it only possible to name everything anew, much confusion would be avoided in the long run, even although we might get bewildered for a time if we began by amending Lavoisier's blunder and calling hydrogen "oxygen."

It was this transgression of Monsieur Lavoisier that started all the trouble in the next chapter!

Chapter XI

GENTLEMEN PREFER BLONDES

> For why? Because the good old rule
> Sufficeth them,—the simple plan,
> That they should take who have the power,
> And they should keep who can.
> —William Wordsworth, *Rob Roy's Grave.*

We have to interview several pretty ladies in the course of this chapter, but the one to whom we shall pay the most attention is chlorine.

Chlorine, you will be sorry to hear, is not a *good* girl like oxygen. Her reputation has never been spotless, and she has recently acquired additional unenviable notoriety in consequence of her appearance as a poison gas during the World War. In spite of all this, chlorine has a great many saving graces, as you will see. To quote the words of the old ballad,

> She is more to be pitied than censured,
> She is more to be helped than despised,

since the sad fact must be admitted that she started out in life under a cloud, and as usual " a man was the cause of it all " ! The villain in chlorine's case was our old acquaintance, Antoine Laurent Lavoisier.

The late alchemists had found that when common salt was heated with " oil of vitriol " (sulphuric acid) a colourless gas was given off which had a very sharp odour. This gas, called " marine air," was very soluble in water, and the strong acidic solution which resulted was given the name of muriatic acid (from the Latin *muria*, brine). In 1774 a poor Swedish apothecary, Carl Wilhelm Scheele, heated muriatic acid with pyrolusite (the black oxide of the metal manganese) and obtained a greenish-yellow gas with a most disagreeable and irritating smell. Scheele thought that this new gas was marine air freed from hydrogen (phlogiston), and he accordingly named it " dephlogisticated marine air." Now this was a terrible name for a young lady to be encumbered with, and Lavoisier, with the best intentions in the world, went to work to get it changed. You will remember that Lavoisier believed oxygen to be an indispensable part of every acid. Muriatic acid—we now call it hydrochloric acid—was accordingly regarded by Lavoisier as a compound of oxygen with some unknown element, and chlorine was assumed to be a higher oxide of this same element. The new name which Lavoisier proposed to bestow upon " dephlogisticated marine air " was " oxymuriatic acid," and for many years chlorine had to struggle along under this disguise, enjoying none of the rights and privileges of an element, while chemists all the world over, following Lavoisier's lead, were striving their utmost to effect her decomposition. It was not until 1810 that Davy finally convinced his colleagues of the inherently simple nature of the gas, and gave her her legal name " chlorine " (from a Greek word meaning *green*).

Before we proceed to discuss the habits of chlorine more critically, let us spare a short paragraph for Scheele, her discoverer. Scheele was undoubtedly the greatest chemist who ever kept an apothecary's shop. In his obscurity in distant Scandinavia he had little opportunity to learn what advances were being made by other scientists of his time, but it is worthy of note that not only did he win the honour of being the first man to prepare chlorine, but he also discovered hydrogen and oxygen quite independently of Cavendish, Priestley, and Lavoisier. He literally lived in his research work, as may be seen by an extract from a letter which he wrote to his friend Gahn, in 1775:

> Oh, how happy I am! No care for eating or drinking or dwelling, no care for my pharmaceutical business, for that is mere play to me. But to watch new phenomena—this is all my desire, and how glad is the inquirer when discovery rewards his diligence; then his heart rejoices.

Unfortunately, Scheele's health broke down under the strain of continuous work and exposure to noxious gases in the poorly furnished shed which served him as a laboratory, and he died in 1786 before he had reached his forty-fourth year.

For many years Scheele's original method was employed in the preparation of chlorine, but cheaper ways of releasing the gas from common salt have now been devised, and at present practically all of the chlorine used in industry is manufactured by the electrolysis of brine. The essential mechanism of this process has already been described. The chlorine, which bubbles off from the anode, is carefully dried and compressed to liquid form in iron cylinders. It is an interesting fact that chlorine is considerably more circumspect in her behaviour when she is placed in a perfectly dry environment. Moist chlorine attacks almost all metals very rapidly. A solution of chlorine in water, indeed, will even dissolve gold—a property which is used in the extraction of that element from low-grade ores. Bonedry chlorine, however, can be stored for years in iron cylinders without damaging the metal at all.

As might be suspected, chlorine is in general a faster worker than oxygen, under similar conditions; she gets her man with promptness and dispatch. Powdered metals, such as copper and iron, catch fire at once when thrown into a jar of the gas and the chlorides of the metals are formed:

$$Cu + Cl_2 \rightarrow CuCl_2$$
$$2Fe + 3Cl_2 \rightarrow 2FeCl_3.$$

A mixture of the gas with hydrogen does not need to have a match put to it to make it explode; it detonates violently in sunlight, forming gaseous hydrogen chloride:

$$H_2 + Cl_2 \rightarrow 2HCl.$$

Many other non-metals, such as phosphorus and sulphur, also combine directly with chlorine. With a limited supply, phosphorus forms the liquid trichloride PCl_3; if an excess of chlorine is present, the solid pentachloride PCl_5 is obtained. Sulphur gives the compound S_2Cl_2, a liquid used in vulcanizing rubber. With carbon monoxide, chlorine combines to form a volatile liquid $COCl_2$, known as phosgene. Phosgene may be remembered as one of the most virulent of the poison gases employed in the World War. At present it is manufactured in large quantities in connection with the dyeing industry.

Chlorine itself also enters into many useful industrial applications. A solution of chlorine in water functions as a very efficient bleaching agent, and this property is utilized on a large scale in the manufacture of paper from wood-pulp and other raw materials, as well as in the cotton and linen trades. In spite of its irritating properties when inhaled in quantity, chlorine in minute traces is good for the human system, since one one-hundredth of the concentration hurtful to us is sufficient to kill the organisms which cause common colds. This fact was first noticed during the influenza epidemic of 1918, when it was found that workers in chlorine factories were practically immune. Bacilli in drinking-water and in swimming tanks are also very rapidly destroyed by the addition of minute traces of chlorine. One of the best antiseptics used on wounds during the World War was Carrel-Dakin solution, which is prepared by passing chlorine into a solution of sodium carbonate (washing soda). In compounds like sodium chloride, we need scarcely mention, chlorine is thoroughly domesticated and no traces at all are observable of the dangerous properties which she exhibits in the free state.

Let us next introduce you to some of chlorine's relations. She belongs to a family of elements which chemists call the *halogen family*, because the most important property of all its members is that of forming salts with metals. (The word halogen is derived from the Greek and means *salt-producing*.) First of all, we want you to meet the two elder sisters of chlorine, bromine and iodine. These two elements were not, it is true, discovered until after chlorine, but they are much more mature in their behaviour. Bromine is red-headed, and iodine is a decided brunette. In more scientific terms, bromine is a brown-red liquid, and iodine is a solid which is almost black. When bromine is warmed it is converted into a dark-red vapour which is exceedingly corrosive and poisonous. Iodine on being heated gives a violet vapour. We always write the formulas for these two elements as Br_2 and I_2 to express their behaviour in the vapour state.

This procedure is not strictly consistent with our action in writing sulphur as S and phosphorus as P, but it has the advantage of enabling us to consider all of the halogen family on an equal footing.

Compounds of bromine and of iodine are by no means so plentiful in nature as those of chlorine, nor are they so important. Certain salt deposits contain minor quantities of bromides admixed with the chlorides, and traces of iodine compounds are present in seaweed. Crude Chile saltpetre $NaNO_3$ also contains a small percentage of combined iodine. Potassium bromide KBr is used in medicine as a sedative—perhaps you remember the old question, "Are you a bromide?"—while silver bromide $AgBr$ is employed in the manufacture of photographic films and plates. A solution of iodine in alcohol painted over the skin reduces swellings and inflammation, and thyroxin, a complex organic compound containing iodine, is used in the cure of goitre.

Bromine and iodine may be obtained in the free state either by electrolysis of a solution containing one of their salts or, still more simply, by bubbling chlorine through such a solution. The reactions which occur in the latter case may be written:

$$2NaBr + Cl_2 \rightarrow 2NaCl + Br_2$$
$$2NaI + Cl_2 \rightarrow 2NaCl + I_2.$$

These reactions, you will note, are examples of *displacements* similar to those which occur when an acid acts upon a metal. Chlorine functions in each case in the rôle of a home-wrecker; she entices sodium away from his less attractive mate and leaves her elder sister to the mercy of a cold world. Do not feel too sorry for bromine in this connection, however, because she has no hesitation in displacing iodine in exactly the same way whenever she gets an opportunity. Thus, when we add bromine to a solution of sodium iodide, we obtain the reaction :

$$2NaI + Br_2 \rightarrow 2NaBr + I_2.$$

Of course, we do not have to use the sodium salt in these displacements ; any other soluble salt will do equally well. In other words, all gentlemen behave alike ; they are all equally fickle and they will desert a brunette for a blonde at the slightest temptation.

It will be instructive for us to probe these domestic triangles a little more closely. Bromides and iodides, like chlorides, are typical polar compounds, and in water solution their negative ions Br^- and I^- are essentially free to act as they please, independently of the nature of the positive ion present, as we have seen in considering the way in which an electric current is conducted through a solution of sodium chloride. We may therefore write the equations for the displacement of bromine from bromides by chlorine, or for the displacement of iodine from iodides by bromine, in the following simplified way, leaving out of consideration altogether the positive ion, which is unchanged in the course of the reaction :

$$2Br^- + Cl_2 \rightarrow 2Cl^- + Br_2$$
$$2I^- + Br_2 \rightarrow 2Br_2^- + I_2.$$

A glance at these equations will expose to you the true mercenary character of the halogen elements. What they actually care for is not the metal that is keeping them company, but the electron extracted from that metal. In other words, it is not the man with whom they are in love, but his money, and when one of them supplants another in the affections of a metal she is careful to annex all his loose electrons as well.

Remorseless fate, however, is hot at chlorine's heels, and she can be made, on occasion, to appreciate the fact that " the punishment fits the crime." Besides her two elder sisters, chlorine has a younger one named fluorine, and fluorine, in the words of Elinor Glyn, is a flapper positively saturated with " IT." Chlorine's golden locks have become somewhat tarnished with green, but fluorine is a genuine canary-yellow blonde, and when she comes on the scene even chlorine finds herself forced to take a back seat.

Davy recognized more than one hundred years ago that this younger sister of chlorine was present in fluorspar (calcium fluoride) and other minerals, but nobody succeeded in getting an actual glimpse of her for many years. The difficulty was that the element is so irrepressible that it immediately reacts with everything with which it comes in contact, corresponding almost, in point of fact, to the " universal solvent " of the alchemists. In 1886, however, Moissan (the same man who manufactured synthetic diamonds) constructed a special platinum apparatus in which he electrolysed a mixture of liquid hydrogen fluoride and potassium fluoride at a very low temperature, thereby obtaining a gas which turned out to be the missing element. This gas is so active that it forms compounds with practically every other companionable element except chlorine, whose aloofness may be ascribed to jealousy, and oxygen, who is too prudent to risk contamination.

The most important compound of fluorine is hydrogen fluoride, which dissolves in water to give hydrofluoric acid. This acid readily attacks substances like glass and sand, which are not affected by other acids. It is consequently used for etching glass and for cleaning the exteriors of granite and sandstone buildings.

The family history is not yet complete. There is a fourth and much older sister of chlorine wandering about somewhere or other in the universe, but no one has yet succeeded in tracing her. We might begin to advertise for information regarding her whereabouts as follows:

Lost, stolen, or strayed: One stout Coloured Lady—

but perhaps it will be safer to substitute scientific terms!

Missing: An element of the halogen family; atomic weight, approximately 218; jet-black solid with a typical metallic lustre; density about 8; melting point about 230° C.; practically insoluble in water; not very active, easily displaced from its compounds by iodine; hydrogen compound a strong acid but very unstable; sodium compound resembles common salt but melts at a lower temperature (about 600° C.).

This description will be sufficient for the present. We might continue by saying that the element is exceedingly radioactive and has an atomic number of 85, but these are points which we have not yet taken up in our discussions. You will probably wonder how it is possible for us to be so positive in our identification marks, since nobody has ever enjoyed an opportunity of examining the element. The explanation, however, is very simple. All of the properties of the halogen elements and of corresponding compounds of these elements change in regular steps as we proceed from one member of the family to the next. Consequently, when we have learned the characteristic features and habits of two or three, we can predict how the rest would look and behave under similar circumstances.

As an illustration of this, you may note that bromine stands almost exactly half-way between chlorine and iodine in all of its physical and chemical properties. Even its atomic weight (80) is practically the mean of the atomic weights of chlorine (35·5) and iodine (127). Indeed, when a sample of bromine was first sent to the laboratory of the famous German chemist Liebig for examination, he thought it was an unstable compound of chlorine and iodine and did not trouble to study it farther. A few years later a young Frenchman, Balard, announced the isolation of a new halogen element, and Liebig went back to the vessel which he had put aside and found that his preconceived ideas had prevented him from gaining the glory of the discovery.

The question naturally comes up: "Can we group other elements into families in this way?" If we could, it would greatly simplify the study of chemistry. Were all of the ninety elements entirely distinct in their behaviour, we should be forced to remember a terrible number of isolated facts and we should not be able to deduce any general laws regarding reactions; every single detail would have to be tested experimentally. On the other hand, if we could succeed in grouping our ninety elements into a few harmonious families, then our task would be considerably lightened.

Fortunately, it has been found possible to make such a classification. For example, oxygen, our charming Cinderella, possesses three ugly sisters—sulphur, selenium, and tellurium—which imitate her conduct to the best of their ability. Nitrogen, the other gas present in the atmosphere, is the first member of another family, which also includes phosphorus, arsenic, antimony, and bismuth. The physical properties in this family, as frequently elsewhere, differ widely as we go from element to element, but corresponding compounds are of similar character and exhibit regular gradations in properties throughout. Carbon, finally, is the first member of a fourth family in which silicon is the only other non-metallic element that finds a rightful place, although a third non-metal, boron, tries to tag along.

Most of the elements which we have mentioned in the last paragraph are of considerable importance both in themselves and on account of the many useful compounds which they form. We shall not have space in this volume to consider them in detail, but the deficiency will (we trust) be remedied later.

Metals also may be grouped into a limited number of distinct families, in every one of which a regular gradation of physical and chemical properties of the elements and of their corresponding compounds may be similarly observed. We are thus enabled to collect all of our ninety elements into a much smaller number of coherent groups, each filling a distinct place in the whole system. To introduce an analogy which we shall amplify later, we may imagine that the ninety men and women constituting our chemical community seek to occupy a block of flats, or (still better) a typical Apartment House, such as is common in Continental cities and in America, each family leasing (or sub-leasing) a separate suite of rooms.

In subsequent chapters we shall inspect this Apartment House from top to bottom, paying a short visit to the families in each story. Before we proceed to indulge in social calls of this kind, however, we have one important piece of business to dispose of. So far we have taken into consideration only the material changes which accompany chemical reactions ; energy changes, which are of equal importance, have not been discussed at all. These energy changes and their significance will form the contents of our next chapter.

Chapter XII

" PEP " AND POPULARITY

Once I saw a chemist take a pinch of powder
—Simple dust it seemed—and half-unstop a phial :
—Out dropped harmless dew. " Mixed nothings make "—
quoth he—
" Something ! " So they did : a thunderclap, but louder—
Lightning-flash, but fiercer—put spectator's nerves to trial :
Sure enough, we learned what was, imagined what might be.
—Robert Browning, *Ferishtah's Fancies.*

As we have had occasion to remark once before, the trouble with these literary people is that they are so terribly indefinite. Browning, of course, was a worse offender in this respect than most of his colleagues, and even the experts still find it difficult to fathom his meaning in places. However, if you read through the above quotation very attentively several times the fact will gradually dawn upon you that somebody took " a pinch " of something and added to it " a drop " of something else, whereupon there resulted a very good imitation of a Fifth of November display. The formula of his mixture Browning is careful not to disclose.

Many chemical reactions do proceed with a tremendous bang. You can illustrate this for yourself by taking a dry milk bottle containing a few generous " pinches " of finely divided aluminium, blowing into it a rapid current of oxygen so as to stir up the dust (as shown in the diagram, *Fig. 38*), and then applying a match to the mouth of the bottle. It is advisable to protect yourself against broken glass by surrounding the bottle with a thick wire screen before applying the match, since, although the bottle usually survives the shock, the result is always sufficiently startling to lead you to resolve never to try the experiment again. In order to express the complete story of this reaction, you must evidently write its equation as follows:

$$4Al + 3O_2 \rightarrow 2Al_2O_3 + \text{energy}.$$

All reactions, however, do not proceed so vigorously or so rapidly as this. When we manufacture water-gas by blowing superheated steam over white-hot coke, we find that the temperature falls rapidly to a point where the reaction will not continue at any reasonable rate unless air is blown through the mass to raise it to a white heat again. Evidently we have here a chemical change in which heat is consumed, and we must therefore write the full equation thus:

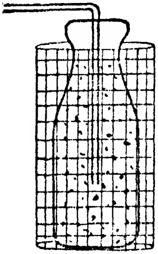

FIG. 38.—Explosion of Aluminium Dust in Oxygen.

$$C + H_2O + \text{energy} \rightarrow CO + H_2.$$

The energy involved in chemical reactions does not always appear, nor is it always used up, in the form of heat. You may convince yourself of this in the following way. First of all, take a stick of ordinary commercial zinc and immerse it in a beaker containing dilute sulphuric acid (*Fig.* 39). Bubbles of hydrogen soon begin to come off, the zinc goes into solution as zinc sulphate, and a large amount of heat is developed. But now change your experiment as follows. Before putting the zinc rod in the acid, connect it with a wire to one post of a galvanometer (an instrument for showing the passage of an electric current), connect a platinum plate similarly to the other post, and immerse the two metals in the acid, taking care that they do not touch (*Fig. 40*). Exactly the same chemical change goes on as before, but the deflection of the pointer on the galvanometer shows that now it is *electrical energy* that is being produced.

The practical importance of energy in chemical changes scarcely needs to be emphasized. We do not burn coal in order to manufacture carbon dioxide gas; it is the heat we want. We do not buy petrol in order

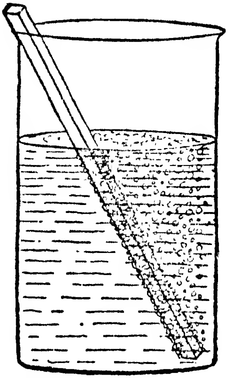

FIG. 39.—Action of Sulphuric Acid on Zinc. Development of Heat Energy.

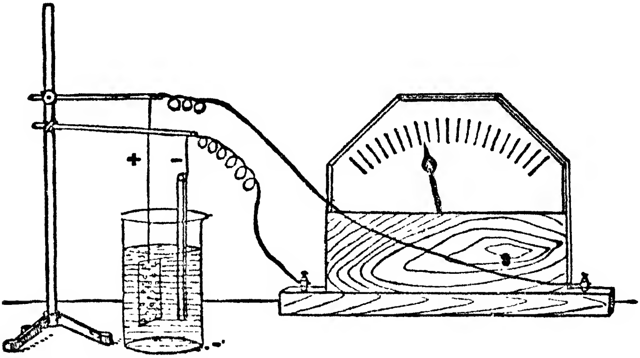

FIG. 40.—Action of Sulphuric Acid on Zinc. Development of Electrical Energy.

to collect various gases from the exhaust of our motor-cars; we really pay for the mechanical energy which is made available by its combustion. The same thing holds when we eat food or use a flash-light; the chemical reactions that occur in each case are of only accidental interest.

The early chemists had no suspicion that there was any relation at all between the different forms of energy—heat, light, friction, electricity, etc. Even Robert Boyle, as we have seen, subscribed to the current view that heat was a form of matter. The first man to combat this opinion successfully was Benjamin Thompson, Count Rumford. Born in Massachusetts in 1753, Thompson held a commission in the American army during the first stages of the Revolution, but was compelled, because of quarrels with his fellow-officers, to leave the country. Wandering through Europe as adviser to various potentates after the war was over, he happened to become interested, in St. Petersburg, in the enormous amount of heat developed during the boring of cannon, and in 1798 he presented to the Royal Society of London a paper which denied that heat was a material substance, postulating instead that it was a form of motion which could be excited by friction. It is a pity that Rumford, with his extraordinary genius, did not follow up these investigations intensively; but soon after this he married the widow of Lavoisier and led such an exceedingly uncomfortable life with her that domestic friction excluded all other varieties from his mind.

It was not until 1840 that Joule—a Manchester brewer who was a pupil of Dalton and made scientific research his main recreation—showed in a quantitative way the connection between heat and electrical energy. Shortly thereafter he put forward the law of the *transformation of energy*, which stated that any one kind of energy could be transformed directly or by steps into any other kind. All types of machinery which we employ at the present time, it may be noted, are devices for effecting such transformations. Energy is thus seen to be work and anything which can arise from or be converted into work. A little later another law was recognized—the law of the *conservation of energy*—according to which no gain or loss of energy is ever observed in an isolated system. Where energy is apparently lost in machinery which is not 100 per cent efficient, the missing part really dissipates itself in a non-usable way—for example, in frictional heat.

How can we bring all of this into harmony with the facts of chemical reactions ? What is the source of the energy obtained when aluminium combines with oxygen, and where does the energy go when water-gas is produced from steam and coal ? To be consistent we must assume that the substances which enter into and are produced by chemical reactions contain energy as well as matter. This energy may conveniently be called *internal energy*.

Some elementary substances evidently contain a large amount of internal energy which is readily available for transformation into other forms of energy. A very good example is atomic hydrogen, as we have seen in an earlier chapter. Substances of this character will be in general extremely active, and when they combine with other elements a large evolution of heat is likely to occur. Their compounds, on the other hand, in consequence of the large loss of energy involved in their formation, will be in general correspondingly inactive. Thus the great majority of the hydrocarbons, or compounds of carbon and hydrogen, being formed from their elements with evolution of heat, are extremely inactive in the chemical sense.

Conversely, elementary substances like nitrogen, which must be assumed to contain relatively little available internal energy because of their inertness, generally require energy to be supplied to them before they can enter into combination with other elements. Thus nitrogen does not combine with oxygen at all at low temperatures, but when the two gases are heated together in the electric arc small quantities of nitric oxide NO are formed. Finally, compounds of inert elements are likely to be extremely reactive in consequence of the energy which they can release in slipping back to the low internal energy level of their constituents. Indeed, the majority of explosives in common use, such as nitro-glycerine, gun-cotton, and trinitrotoluene (or TNT), are nitrogen compounds.

Why should certain pairs of elements, such as hydrogen and chlorine, unite readily to form compounds like hydrogen chloride, while other pairs cannot be persuaded to combine under any circumstances? Chemists used to answer this question by saying that hydrogen and chlorine had a great affinity for one another, and that non-reactive pairs, such as oxygen and fluorine, had little or no affinity. A reply of this kind, however, is obviously only an ingenious method of concealing ignorance, and we are now in a position to give a better explanation.

Attention has already been called several times to the fact that the atoms of most elements contain as an outermost shell a number of loose electrons. In the course of chemical reactions of the ordinary types which we have so far considered—combination, decomposition, and displacement—the interior part of each atom is entirely unaffected; only the outermost electrons change their positions. The atoms of certain elements commonly cast off all their loose electrons in entering into combination. The atoms of other elements take on extra electrons, while in certain cases pairs of electrons may be shared in common between two atoms. Now this shifting of electrons from their original arrangement entails a change in internal energy; work may be done or work may be expended in moving to the new stations.

The most important and at the same time the simplest type of reaction that we can examine in this connection is that of the combination of a typical metal, such as sodium, with a typical non-metal, such as chlorine, to give a polar compound, such as common salt. The magnitude of the energy change here involved evidently depends upon two factors: (1) the desire of the sodium atom to lose an electron, and (2) the desire of the chlorine atom to gain an electron. In this particular case the mutual desires are obviously quite strong, since their satisfaction leads to the formation of a very stable compound. Expressed quantitatively, the *heat of formation* of one gramme molecular weight of sodium chloride from its elements is 97,700 calories.[1] When the other halogens combine with sodium, the heat of formation is different. We have already discussed in detail how bromine and iodine are inferior to chlorine in their power to hold the

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[1] The chemical calorie is the average quantity of heat required to raise one gramme of water one degree in temperature, between 0° and 100°. The food calorie, with which you are probably more familiar, is one thousand times this quantity.

affections of the metals, while fluorine surpasses her elder sister in this respect. We would expect the fact that bromine and iodine grab sodium's loose electrons with less avidity than chlorine does to show up in a lower heat of combination, while the flaming youth of fluorine should lead to a higher heat effect. The actual figures are: for sodium bromide, 85,800 calories; for sodium iodide, 69,100 calories; and for sodium fluoride, 109,300 calories. The greater popularity enjoyed by the younger members of the halogen family is thus explained by their greater "pep."

So far, in looking into domestic disputes, we have put all of the responsibility upon the shoulders of the ladies. This, however, is hardly a fair thing to do. While it is true that gentlemen prefer blondes, it is just as true that blondes prefer gentlemen, and the metals are by no means all equally agreeable partners. Some, like sodium, act like born aristocrats; they distribute their money in an open-handed fashion. Others, like platinum and gold, are terribly tight-fisted, and their spare funds can be extorted from them only with extreme difficulty. The relative tendencies of the various metals to part with their loose electrons have already been illustrated in the chapter on hydrogen. We saw there that metals like zinc and iron displaced hydrogen from dilute acids without difficulty, and we wrote the reactions as follows:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
$$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2.$$

In the light of what we have learned since, we can now simplify these reactions to read:

$$Zn + 2H^+ \rightarrow Zn^{++} + H_2$$
$$Fe + 2H^+ \rightarrow Fe^{++} + H_2.$$

Still more active metals, such as sodium and potassium, will displace hydrogen even from a less polar compound, like water. On the other hand, canny plutocrats like silver, platinum, and gold will not be cajoled into a display of generosity of this kind at all.

Order of Activity.
Metals.

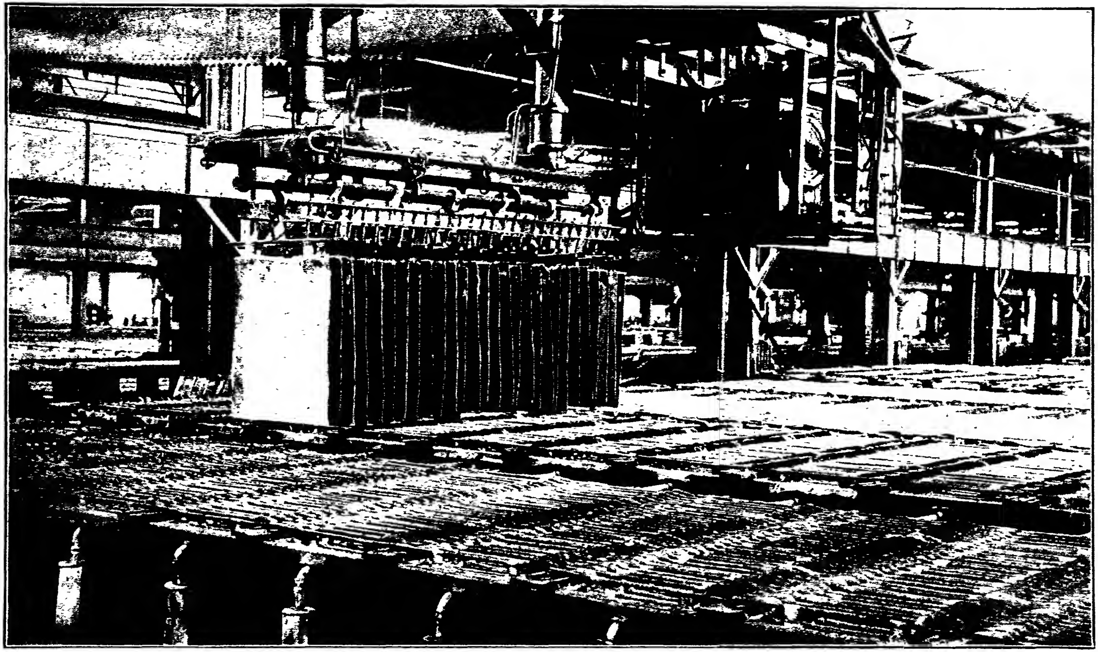
Potassium
Sodium
Calcium
Magnesium
Aluminium
Manganese
Zinc
Chromium
Iron
Nickel
Tin
Lead
Hydrogen
Antimony
Bismuth
Copper
Mercury
Silver
Platinum
Gold

Comparing all of the eligible manhood in our chemical community as far as possible under similar conditions, we can draw up what the ladies might call a "Popularity List," in which the metals are arranged in the order of their liberality with respect to handing out electrons. The chemist calls this list the *order of activity* of the metals. Every metal above hydrogen in the list will relinquish electrons to hydrogen ions, the reaction being more energetic the higher the position of the metal.

This last remark gives us a good opportunity to clear up a point which we left unexplained in the chapter on valence—namely, why the Na^+ ions, which are attracted to the negative electrode in the hope of securing electrons when a solution of sodium chloride in water is electrolysed, are not discharged thereon as metallic sodium when they reach their goal. There are always some H^+ ions present in any aqueous solution, and although the Na^+ ions do practically all the hard work of carrying the current, yet, as these Na^+ ions come sauntering up toward the cathode to receive their reward in the shape of the extra electrons which this carries, the disrespectful little H^+ ions dodge in ahead of them, grab the electrons, and are discharged as hydrogen gas, while the sodium ions get nothing tangible for their labour at all. However, each H^+ ion which is discharged leaves its original partner in the water molecule, a hydroxyl ion OH^-, to keep company with each disappointed sodium ion, and the solution around the cathode consequently becomes increasingly rich in sodium hydroxide $NaOH$ as the electrolysis proceeds. This is, indeed, the actual method employed in the commercial production of sodium hydroxide at Niagara Falls, the other valuable products of the electrolysis of the brine being chlorine (discharged at the positive electrode) and hydrogen (discharged at the negative electrode).

Similar side reactions occur at the cathode when we electrolyse a water solution of a salt of *any* metal high up in the activity list. Instead of the metal being plated out on the electrode, hydrogen is evolved. If our metal is below hydrogen in the activity list, however, there is no hindrance to its deposition, and large quantities of metals like copper and silver are refined electrolytically by taking advantage of the fact that they may be plated out successfully in this way, while the more active metals present as impurities remain in solution.

Hydrogen, however, is not a unique element on the activity list, and we may use any other element as a basis for matching activities. We find, indeed, that experiment bears out our expectation that any element higher up in the activity list will displace any element lower down in the list from its compounds. The metals at the very foot, platinum and gold, are particularly easily displaced from solutions of their salts, and this property is utilized in recovering small traces of these valuable metals from solution. The method employed may be understood by



© *Anaconda Copper Co.*

ELECTROLYTIC REFINING OF COPPER. LIFTING CATHODES OUT OF TANK considering a more familiar case. If we take a solution of copper sulphate and immerse in it a rod of zinc, we find that the rod slowly disintegrates, the zinc going into solution and its surface being covered with a red coating of metallic copper. The reaction which takes place here is:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu;$$

or, more simply, since the sulphate ion is unchanged throughout:

$$Zn + Cu^{++} \rightarrow Zn^{++} + Cu.$$[1]

This same reaction can be carried out in a slightly modified way so as to give electrical energy. The set-up shown in *Fig. 40* at the beginning of this chapter may be employed, substituting a rod of copper for the platinum plate and copper sulphate solution for sulphuric acid. This particular combination is used in practice under the name of the Daniell Cell, which delivers an electric current of a little over one volt. If we substitute for zinc a metal lower down on the activity list, such as tin, a lower voltage is obtained, while if we substitute for copper a metal farther down the list a higher voltage results. The activity list is consequently frequently called the *electromotive series* of the metals.

A similar series can be drawn up for the most important non-metals, showing the order of their adroitness in attracting electrons to them. Any element higher

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[1] Do *not* interpret this equation to mean that zinc has stolen positive charges from copper; what has actually occurred is that each atom of zinc has donated to each copper ion two electrons as "heart-balm" for the alienation of a sulphate ion's affections.

ORDER OF ACTIVITY.
NON-METALS.

Fluorine
Chlorine
Bromine
Oxygen
Iodine
Sulphur

up on this list will displace any element lower down on the list from its polar compounds.

While the metals which exhibit the most " pep " and are most prodigal in disposing of their electrons are most popular with the non-metallic elements, that fact does not make them most popular with or most useful to the human race. Indeed, the more active a metal is, the more difficult it is to isolate and the harder it is to keep it out of combination. The most active metals, potassium and sodium, must be preserved under petroleum in order to prevent their immediate combination with the oxygen of the air. As we go down the activity list, however, the metals become more stable and oxidize with greater and greater difficulty. The metals at the very foot of the list do not tarnish in the air at all, and their oxides must be obtained by indirect methods. These oxides are very unstable and readily decompose when heated. As we go up the list again, the corresponding oxides become more and more stable, until at the top we encounter compounds like quicklime which cannot be decomposed by heat at all, and must be broken up by the use of more intense electrical energy.

The value of the " noble metals," like platinum and gold, lies in their passivity and permanence. Copper and iron share this inertness to a lesser degree, and their comparative susceptibility is compensated for by their cheapness. It is interesting to note that, if we read through the activity list upward, we have approximately the historical order in which the metals came into use. The metals which lie significantly above hydrogen on the list are too active to be found free in nature, and have been pressed into the service of humanity only in recent years.

The same thing holds for the non-metals. Fluorine may be more popular than oxygen or sulphur, but she is not nearly so useful. Many compounds also owe their usefulness to their inactivity. Materials of this type are sandstone, granite, brick, porcelain, glass, and, to a limited extent, water. All of these substances are fully oxidized and are therefore fireproof. Every so often the claim is put forward in the newspapers that some genius has discovered the secret of using water instead of petrol as an automobile fuel, and the fact that water contains a highly combustible gas, hydrogen, is adduced as a proof of the possibility of burning water in this way. Unfortunately, the hydrogen that exists in water is already completely burned. It is true that the reaction $2H_2 + O_2 \rightarrow 2H_2O$ involves the release of 116,200 calories for the production of two gramme molecular weights (or 36 grammes) of water, but before we can decompose water and get hydrogen from it we must first *supply* this same tremendous quantity of energy to every 36 grammes of water that we decompose. To try to use water as a fuel is, consequently, just as sensible as to try to lift one's self from the ground by one's boot-straps.

The preceding paragraphs must not be taken to imply that substances which are active are not useful to mankind. Very frequently it is their activity that makes them of service to us. The good work that oxygen performs in this connection has already been described in detail. Ozone, which is obtained from oxygen with the absorption of energy, is able to bring more energy into a reaction than oxygen itself does and is therefore more active and less stable. Explosives like nitro-glycerine and TNT are of value solely because of their extreme instability, which can be made use of either constructively for blasting and mining projects, destructively for war purposes, or instructively for the edification of poets like Browning.

Chapter XIV

THE APARTMENT HOUSE

> Chemistry marches towards its goal and towards its perfection by dividing, subdividing, and subdividing still again.
>
> —J. B. Dumas.

Some busybody is bound to ask at this point, " What has become of Chapter XIII ? " If you will look at the title at the top of this page, you will understand at once why Chapter XIII was skipped. Nobody deliberately tempts fate by numbering a flat thirteen ; indeed, the taller apartment houses in America even omit the thirteenth floor. A scientist, of course, should not be superstitious, but in view of the fact that the elements themselves carefully avoid the atomic weight of thirteen (carbon being twelve and the next element, nitrogen, fourteen), we have been courteous enough not to injure their feelings here by herding them together under the handicap of a so-called unlucky number.

Let us see how the elements were persuaded to take a lease on the highly desirable property which we shall shortly describe. Of course, this is an incident which does not go back very far into history, since the modern apartment house is a fairly recent invention.

The grouping of the ninety elements into a limited number of distinct families was clearly only a preliminary step to the more important task of organizing them into a complete community. The first real effort to achieve this latter aim was made by an English chemist named Newlands in 1863. Newlands noted the surprising fact that if the elements were arranged in the order of ascending atomic weights, every successive eighth element was a " kind of repetition " of the first. " In other words," said Newlands, " members of the same group of elements stand to each other in the same relation as the extremities of one or more octaves in music. This peculiar relationship I propose to provisionally term *the law of octaves*."

Besides splitting his infinitives, Newlands was handicapped by the doleful fact that he did not know that several notes on his chemical keyboard were missing. A number of elements had not yet been discovered at that time, and consequently, when Newlands put his fingers on two notes an octave apart, he did not always get the expected harmony. In order, therefore, to show the full value of Newlands' idea, we shall not reproduce his original octaves, but shall amend them by omitting hydrogen (the first note on the scale) and inserting certain elements that have since been discovered. The first three " octaves," then, run thus :

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca	Sc	Ti	V	Cr	Mn.

Looking at these octaves, we see that elements of a similar character, belonging to the same family, fall in the same vertical column, while if we read any octave horizontally a regular and progressive change in properties is observable. Thus, if we take the middle octave, we notice that the first three elements are metals of decreasing positive character, while the last four are non-metals of increasing negative character. The most interesting property to discuss in this connection, however, is valence, and the way in which this varies periodically as we ascend the scale is worth examining in detail. Let us take the liberty of reminding you, before we begin, that hydrogen normally exhibits a positive valence of one, and oxygen a negative valence of two, as in H_2O. The valence of other elements can generally be established from their compounds with these two. Thus the formula Na_2O for the polar compound sodium oxide shows that two atoms of sodium have a total valence equal and opposite to that of one atom of oxygen, whence the valence of sodium is found to be $+1$.

The elements in the first column all exhibit a positive valence of *one*, lithium resembling sodium and potassium in giving an oxide, with the formula Li_2O, which dissolves in water to form a strong alkali. This family is, consequently, called " the alkali metal family." The elements in the second column are unanimous in exercising a positive valence of *two*, beryllium resembling magnesium and calcium in giving an oxide with the formula BeO. The oxides of these metals also dissolve in water, though not very readily, to give alkaline solutions, and inasmuch as compounds of these metals are commonly found in the earth, the family is known as " the family of the alkaline earths." The elements in the third column are not so uniform in their behaviour, the first member, boron, being a non-metal. All of these elements, nevertheless, agree in exhibiting a valence of *three* in their oxides, such as B_2O_3 and Al_2O_3.

The fourth family shows still less harmony in its habits, and non-metallic properties persist farther up the keyboard. The formulas of the typical oxides, carbon dioxide CO_2 and silica SiO_2, might be construed to indicate a positive valence of *four* for the elements in this column, but their compounds with hydrogen, methane CH_4 and silicon hydride SiH_4, suggest equally plausibly a *negative* valence of *four*. As a matter of fact, the compounds of carbon and silicon are predominantly non-polar in character, and we have to deal rather with co-valence in both cases. In the fifth column we have elements forming compounds such as nitrogen pentoxide N_2O_5 and phosphorus pentoxide P_2O_5, which suggest a positive valence of *five*, and also compounds like ammonia NH_3 and phosphine PH_3, which indicate a negative valence of *three*. Here again, however, the compounds of both types are in general only slightly polar.

Coming to the elements in the sixth column, we find that sulphur trioxide SO_3 implies a positive valence of *six*, while water H_2O and hydrogen sulphide H_2S point to a negative valence of *two*. Negative polarity is now becoming more pronounced, and when we arrive finally at the first elements of the seventh column—the halogen family—we note that their most common compounds are those in which a negative valence of *one* is evident, as in hydrogen chloride HCl. Nevertheless, compounds are formed which do indicate a positive valence of *seven*, such as chlorine heptoxide Cl_2O_7 and manganese heptoxide Mn_2O_7. These highly unstable substances have more important derivatives, such as potassium perchlorate $KClO_4$ (a powerful explosive) and potassium permanganate $KMnO_4$ (a useful antiseptic).

All of this long digression may be boiled down into the simple statement that positive valence increases steadily from one to seven as we ascend an octave, while negative valence increases uniformly from one to four as we descend the upper half of an octave. Now this is a most astonishing and gratifying regularity, since heretofore valence to us was merely an accidental number which had to be memorized independently for each particular element. The full importance of this "periodicity" in valence, however, will not appear until a later chapter.

Here we may stress once more the fact that Newlands did not find it possible to work out the scheme which we have given above so harmoniously as it is now developed. He was already in trouble in his first three octaves because of missing notes, and when he came to the heavier elements the discords which he produced were truly terrible. For this reason his work did not receive the recognition which it deserved at once; in fact, he was ridiculed unmercifully.

What a wonderful evening the dear old Tories of the London Chemical Society did enjoy when young Newlands read his paper on "The Law of Octaves" to them at one of their meetings in 1866! Burlington House fairly rocked with their laughter, and when one venerable gentleman rose from his chair and inquired whether the speaker had ever tried arranging elements according to their initial letters, all Piccadilly knew that something very unusual had happened to rouse the wiseacres from their wonted torpor. It was finally agreed that the best interests of science would be served if the article were buried in the archives of the society. Twenty-one years later, however, the laugh was on Newlands' side when he was awarded the Davy Medal of the Royal Society for his discovery.

Newlands himself had been too deeply discouraged to do any further work; the reversal of attitude was due to the light which was thrown upon the subject through the efforts of other investigators, notably a Russian chemist, Dmitri Ivanovitch Mendelejeff. Mendelejeff was born in 1834 at Tobolsk in Siberia, where his father was director of a school. He was the youngest of a very large family—variously estimated as containing from eleven to seventeen children. His father became blind soon after his birth, and it was only after heroic struggles on the part of his mother that he was enabled to obtain a patchwork education. Nevertheless, he displayed such great ability that he was appointed professor of chemistry in the university at St. Petersburg at the early age of thirty-two. In 1869 he published his " periodic system " of the elements, which immediately attracted favourable attention.

Mendelejeff advanced beyond Newlands in three very significant ways. In the first place, he recognized that the octave idea could not apply to the whole list of elements, although the lighter members fitted into it so remarkably well. Instead, Mendelejeff postulated that after the first two groups, each containing seven elements, there followed further groups not of seven but of seventeen (the congestion in his family environment perhaps suggested this number to him !). Each of these groups of seventeen elements, furthermore, could be split up into three sections : (1) a group of seven, similar to the early octaves ; (2) a group of three, all closely similar in character; and (3) a second group of seven, echoing the characteristics of the first. Each group of seventeen elements Mendelejeff called a " long series," to distinguish it from a simple early octave or " short series." There are three complete long series distinguishable among the heavier elements, and a small fragment of a fourth.

The valence changes observable in each of the two octaves in a long series are in general similar to those which have already been discussed for the short series. The echo, however, of which we spoke, is rather indistinct, as may be illustrated by one or two examples. The first elements in the three complete long series are potassium, rubidium, and caesium. These all resemble lithium and sodium very closely indeed; they are typical alkali metals. The eleventh or echoing elements, however, in these same three series are copper, silver, and gold. These form a fairly coherent group and resemble each other quite closely, but their resemblance to the members of the alkali metal family is limited to the fact that they exhibit in certain of their compounds a positive valence of one. In the same way the second elements of the three long series—calcium, strontium, and barium—are extremely similar; the twelfth or echoing elements—zinc, cadmium, and mercury—also show striking resemblances, but the only feature common to all six metals is a positive valence of two.

The same thing holds true through all the seven notes of the octave, so that we are able to recognize two distinct families in each valence group. The three "transition elements" which intervene between the two octaves of each long series are always metals of almost identical character—for example, iron, cobalt, and nickel. The *maximum* valence which can be exhibited by these transition elements (not all of them, it must be admitted, do exhibit it) is, in accordance with their position, eight. Minor differences in behaviour run in parallel through each group of three, so that the first in one long series resembles most closely the first in another long series, and so on.

The second stride forward which Mendelejeff took was to make the bold assertion that apparent inconsistencies in the order of the elements were due not to imperfections in his periodic system but to errors in the accepted atomic weights. The experimental proof of this soon followed in a great many cases and furnished wonderful support for the validity of the whole system. Thus, in 1869, the accepted atomic weights for gold, iridium, platinum, and osmium placed these elements in the order Au 196, Ir 197, Pt 198, and Os 199. This order did not fit in at all with their properties according to the periodic system, and Mendelejeff insisted that they should be ranked thus : Os, Ir, Pt, Au. When careful investigation of the matter was made, he was shown to be perfectly right, as may be seen by reference to the present values given in the table in the Appendix.

With certain elements the changes in atomic weights demanded by Mendelejeff were still more significant. Thus the atomic weight of uranium was thought, at that time, to be about 119, but there is no place in the periodic system for an element with this value. Mendelejeff pointed out that there was a vacant place into which this element would fit admirably if it were assigned the atomic weight of 2 × 119, or 238. Later work showed that this suggestion was completely justified. You may remember that in an earlier chapter we discussed how Dalton got into difficulties through taking the atomic weight of carbon as six instead of as twelve. At that stage we were unable to give any logical reason for prefering one value to the other. Now that each element must occupy a definite place in the system, however, all doubt as to which figure is correct has vanished. Still further proof of this will appear in a later chapter.

The third line of advance started by Mendelejeff was the most daring of all. Instead of acquiescing in discords in his octaves as Newlands did, Mendelejeff moved an unmelodious element one or two notes ahead, if necessary, until it did harmonize. This left quite a number of blank spaces in his keyboard, and Mendelejeff was audacious enough to prophesy that these vacancies corresponded to missing elements which would some day be discovered to have the properties demanded by their positions in the scale. He even went so far as to predict from the properties of the known elements of a family exactly what the properties of its unknown members would be. For example, a gap existed in his keyboard in the space preceding arsenic. Mendelejeff noted that the element which should fill this gap would be a member of the carbon family falling between silicon and tin, and from the properties of its neighbours he predicted in detail the properties of this unknown element, which he called *eka-silicon*. Fifteen years later the German chemist Winkler discovered a new element which he called *germanium*, and the almost incredible fulfilment of Mendelejeff's prophecies is shown in the following table :

	Mendelejeff's Eka-silicon (1871).	Winkler's Germanium (1886).
Atomic weight .	Es 72	Ge 72·6
Density . .	Es 5·5	Ge 5·47
Colour . .	Es dirty grey	Ge greyish white
Density of oxide .	EsO_2 4·7	GeO_2 4·703
Boiling-point of chloride . .	$EsCl_4$ below 100°	$GeCl_4$ 86°
Density of chloride	$EsCl_4$ 1·9	$GeCl_4$ 1·887
Boiling-point of ethide . .	$Es(C_2H_5)_4$ 160°	$Ge(C_2H_5)_4$ 160°
Density of ethide	$Es(C_2H_5)_4$ 0·96	$Ge(C_2H_5)_4$ nearly 1

Similarly successful predictions were made of the properties of many other elements which have since been discovered, until at the present time there remain only two elements missing from Mendelejeff's complete system of 92. One of these, as we have seen, is a member of the halogen family; the other is an alkali metal.

The remarkable way in which Mendelejeff's system brought all of the elements into an harmonious community led to the almost immediate and universal endorsement of his ideas. Perhaps the best way for you to appreciate at a glance the numerous details in Mendelejeff's work, the most important of which we have tried to indicate already, will be to fit them into a plan of a kind which is familiar to you. The guide-books to museums and other extensive public buildings frequently carry sectional diagrams showing the arrangement of the rooms, to prevent visitors from going astray, and since we are going to find it very helpful to imagine that all of our chemical families are attempting to accommodate themselves in a large apartment house, which we may call Mendelejeff Court, a vertical section of Mendelejeff Court is given for your convenience in *Fig. 41*.

According to this diagram, there are eight stories above the street-level. Most of the floors contain six rooms, but toward the top the building regulations compel us to step back from the street a little. These six rooms we may label, for convenience, starting from the left, A, B, C, D, E, and F. Rooms A, B, and F are single rooms; C, D, and E are double rooms. The three bungalows on the roof are each sufficiently commodious to house three elements.

You will notice that two distinct families occupy each of the seven main floors, and we have attempted to mark

A	B	C	D	E	F	
		Ni 58·7 Co 58·9 Fe 55·8	Pd 106 7 Rh 102·9 Ru 101·7	Pt 195·2 Ir 193·1 Os 190·8		ROOF (Maximum valence + 8)
F 19	Cl 35·5	Br 79·9 Mn 55	I 126·9 Ma (99)	? (218) Re (187)		SEVENTH FLOOR (Valence + 7 or — 1)
O 16	S 32	Se 79·2 Cr 52	Te 127·5 Mo 96	Po (210) W 184	U 238	SIXTH FLOOR (Valence + 6 or — 2)
N 14	P 31	As 75 V 51	Sb 121·8 Cb 93	Bi 209 Ta 181·5	U–XII 234	FIFTH FLOOR (Valence + 5 or — 3)
C 12	Si 28	Ge 72·6 Ti 48	Sn 118·7 Zr 91	Pb 207·2 Hf 178·6	Th (232)	FOURTH FLOOR (Valence + 4 or — 4)
B 10·8	Al 27	Ga 69·7 Sc 45	In 114·8 Y 89	Tl 204·4 La 138·9 and 14 others	Ac (226)	THIRD FLOOR (Valence + 3)
Be 9	Mg 24·3	Zn 65·4 Ca 40	Cd 112 4 Sr 87·6	Hg 200·6 Ba 137·4	Ra 226	SECOND FLOOR (Valence + 2)
Li 6·9	Na 23	Cu 63·6 K 39·1	Ag 107·9 Rb 85·4	Au 197·2 Cs 132·8	? (224)	FIRST FLOOR (Valence + 1)
He 4	Ne 20·2	A 39·9	Kr 82·9	Xe 130·2	Rn 222	BASEMENT (Valence 0)
A	B	C	D	E	F	

FIG. 41—Mendelejeff Court.

this fact for you by assigning to each element living in a double room a definite position in it. If you want to visit the members of any one particular family, therefore, you follow a horizontal line from left to right in the diagram. If you want to visit the elements in the order of ascending atomic weights, you start with room A on the first floor and work upward, then do the same with room B, then make two trips upward through room C on each floor, interviewing the elements on the left on your first trip and those on the right on your second trip, and so on through rooms D, E, and F until you reach the element of highest atomic weight, uranium. You will find that there is only one room which is completely vacant (room F on the first floor), while the double room E on the seventh floor has only one occupant.

So far we have not said a word about the basement, and you are probably wondering why. As a matter of fact, nobody knew that the basement was inhabited at all until a great many years after Mendelejeff Court was built. Probably those among our readers who have lived in an apartment house and who have tried unsuccessfully for years to get the janitor, or any of his confederates who are supposed to dwell in the nether regions, to do some little odd job for them will appreciate how this could be the case. Indeed, the elements which were finally discovered in the basement behaved until recently like regular Weary Willies; they refused to do any work at all. How they were ferreted out from their retreat, and how they have now begun to reform, is a story which may fitly conclude this chapter.

In 1892 Lord Rayleigh, the celebrated physicist, observed that, although samples of oxygen and of other gases made purposely from different sources always had the same densities, nitrogen was a peculiar exception. Nitrogen obtained from the air and supposed to be perfectly pure had a density more than one half of one per cent greater than nitrogen prepared by the decomposition of nitrogen compounds. The suspicion arose that some heavier gas was present in atmospheric nitrogen, and Rayleigh appealed to the chemical world for suggestions. Sir William Ramsay, professor of chemistry at University College, London, was the only chemist to respond by actual experiments, and in 1894 he and Rayleigh succeeded in isolating this heavier gas by passing atmospheric nitrogen repeatedly over heated magnesium. The real nitrogen combined with the metal to form a solid, magnesium nitride Mg_3N_2, but the small quantity of gas which remained refused to combine with magnesium, or indeed with any other element. It was first named *aeron* to denote its connection with the atmosphere, but its discoverers received so many letters inquiring when Moses might be expected to turn up that they finally rechristened it *argon* (from a Greek word meaning *lazy*) in order to avoid confusion.

It is of interest to note that Cavendish almost discovered argon a century earlier. Passing an electric discharge through air in order to cause the nitrogen and oxygen to combine, he found that a small bubble of inactive gas remained, but had no means at that time of establishing its true nature.

Argon is so unsociable that not only does it refuse to combine with other elements, but it will not combine even with itself. The atoms of more congenial gaseous elements, like oxygen and hydrogen, as we have seen, like to go around in pairs, but the molecule in the case of argon contains only one atom. This inert character of the gas is now turned to advantage by using it to fill electric light bulbs.

Other hermit elements were soon brought to book. In 1894 Ramsay, searching for new sources of argon, examined a gas which Hillebrand of the United States Geological Survey had obtained by heating an ore of uranium. Hillebrand had thought that this gas was plain nitrogen, although certain of its properties should have told him that it was not. Ramsay was surprised to find that it contained a large proportion of an inert gas almost as light as hydrogen, and was still more astonished to ascertain, when he came to analyse the spectrum obtained on passing an electric discharge through this gas in a vacuum tube, that its most prominent feature was a magnificent orange line (*Fig. 42*) This same line had been detected previously only in the light coming to us from the outermost fringes of the sun and had been ascribed to an unknown element present in its photosphere. The element had accordingly already been named *helium* (Greek, *the sun*). Helium has since been found to exist in minute quantities in our own atmosphere, and in larger amounts in many natural gases, notably in those connected with the oil-fields of Texas and Western Canada. Methods have recently been developed to obtain helium in a pure state from these latter sources cheaply, and the same gas which a few years ago cost £500 a cubic foot is now manufactured in large quantities for filling dirigibles. Its lifting power

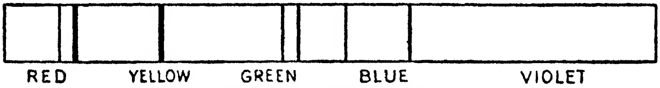

FIG. 42.—Helium Spectrum.

is only a little less than that of hydrogen, and it has the great advantage of being non-inflammable and non-explosive when mixed with air. Helium is also used nowadays to replace the nitrogen in the compressed air supplied to workmen engaged in caisson work—for example, in tunnelling under rivers. It has been found that the lower solubility of helium in water greatly reduces the danger that such workmen will suffer from "the bends" (release of air-bubbles in the blood, frequently with fatal consequences) when they return to the surface.

By liquefying argon at –186° C. and examining carefully the gases evolved at various stages of its re-evaporation, Ramsay was able to show that very small quantities of three other inert elements were also present in the atmosphere. These three gases were named *neon* (Greek, *new*), *krypton* (Greek, *hidden*), and *xenon* (Greek, *stranger*). The penetrating orange-red glow which is obtained when an electric discharge is passed through a vacuum tube containing traces of neon (*Fig. 43*) is now utilized in electric signs and in beacons for aviation fields.

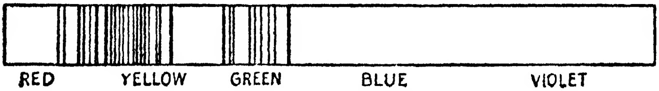

FIG. 43.—Neon Spectrum.

The last element to be found hiding in the basement of Mendelejeff Court was *radon* or *radium emanation*. This gas is employed in hospitals in the treatment of superficial cancerous growths. Its formation from radium will be discussed in the course of the next chapter.

Chapter XV

THE FIRST AND SECOND FLOORS

> We stand to-day towards radioactivity exactly as our ancestor stood towards fire before he had learned to make it. He knew it then only as a strange thing utterly beyond his control . . . so it is that we know radioactivity to-day. This—this is the dawn of a new day in human living.
>
> —H. G. Wells, *The World Set Free.*

There is only one man in history who discovered more elements than Ramsay, and that is our old friend Sir Humphry Davy. In 1807, Davy first obtained sodium and potassium by the electrolysis of their hydroxides. According to his cousin, Edmund Davy, who was his laboratory assistant at the time, he was so excited over his success that he danced about the room in ecstasy. Only a few months later he announced to the world the discovery of five more new elements—magnesium, calcium, strontium, barium, and boron. The first and second floors of Mendelejeff Court would certainly remain rather empty but for Davy's agency.

And yet, as might be expected from his dancing proclivities, Davy was not merely a stodgy scientist with no outside interests. Dalton, who visited him in London, wrote as follows:

> He is a very agreeable and intelligent young man, and we have interesting conversations in an evening. The principal failing in his character is that he does not smoke.

He was fond of travel and sport, and even wrote a book on fly-fishing. His lectures at the Royal Institution were so brilliant that an *Ancient Mariner* like Coleridge went to hear him "to increase his stock of metaphors," and expressed the opinion that "if he had not been the first chemist, he would have been the first poet of his age." As a matter of fact, Davy did write reams of poetry, but we cannot stop to reproduce any of his effusions here. After all, in spite of his own wonderful achievements, his most important discovery was Michael Faraday.

The rise of Michael Faraday from "a bright-eyed errand-boy who slid along the London pavements with a packet of newspapers under his arm" to the recognized leader of scientific research in Great Britain, is one of the finest romances of chemistry. Faraday's own account of how he got his start is reproduced in somewhat condensed form below:

> When I was a bookseller's apprentice, I was very fond of experiment and very adverse to trade. It happened that a gentleman took me to hear some of Sir H. Davy's lectures. I took notes, and afterwards wrote them out more fairly in a quarto volume.
>
> My desire to escape from trade, which I thought vicious and selfish, and to enter the service of Science, which I imagined made its pursuers amiable and liberal, induced me at last to take the bold and simple step of writing to Sir H. Davy, expressing my wishes; at the same time, I sent the notes I had taken of his lectures. Early in 1813 he requested to see me, and told me of the situation of assistant in the laboratory of the Royal Institution, then just vacant.

Now let us hear the other side of the story, as told by one Dr. Gassiot:

On one occasion Sir H. Davy, showing Pepys a letter, said, " Pepys, what am I to do? Here is a letter from a young man named Faraday; he has been attending my lectures, and wants me to give him employment at the Royal Institution—what can I do?" "Do?" replied Pepys, "put him to washing bottles; if he is good for anything he will do it directly; if he refuses, he is good for nothing."

Faraday showed later that he was "good for" the discovery of benzene and the foundation of the laws of electrochemistry, to mention only two of his scientific achievements which have aroused admiration far beyond the borders of pure chemistry itself, as the Prince of Wales and Herbert Hoover have already proved to you in previous chapters.

But we are wandering too far afield from Mendelejeff Court in these historical digressions. Let us, therefore, without further delay, enter upon our task of interviewing the elements on the first floor.

The metals of the alkali family—lithium, sodium, potassium, rubidium, and caesium—are all very much alike, and inasmuch as you have already learned the main facts with respect to sodium and potassium, it will not be necessary to waste much time on their less important brothers. The metals themselves are all so active, so eager to hand over electrons to the first female they meet, that their practical importance lies almost entirely in their compounds. One use of metallic sodium, however, is worth noting. In its presence isoprene, a hydrocarbon with the formula C_5H_8 which occurs in turpentine, and derivatives of isoprene, which can be obtained in quantity from potatoes at a fairly low cost, are converted into products which resemble raw rubber. The best results so far obtained in the manufacture of *artificial* *rubber* have been with the use of methyl-isoprene C_6H_{10}. Artificial rubber cannot yet be made cheaply enough, or of a sufficiently high grade, to compete with the natural product under ordinary circumstances, but in Germany during the war, when the supply of plantation rubber was cut off entirely, considerable quantities of " methyl rubber " were manufactured from methyl-isoprene by this synthetic method.

Compounds of lithium are rather rare, but their occurrence in certain mineral waters (lithia springs) makes these of value for medicinal purposes. Sodium compounds are much more common, the chloride being the most useful of all salts. Sodium nitrate or Chile saltpetre $NaNO_3$ is of great importance in the explosive and fertilizer industries. Sodium carbonate (washing soda) is used in large amounts in glass manufacture and in water softening. Sodium hydroxide has many uses, the chief of which is in soap-making. Many other compounds of commercial importance might be mentioned, but lack of space forbids.

Compounds of potassium are in general similar to those of sodium in their physical properties and chemical action. Potassium chloride is most valuable as a fertilizer. Potassium nitrate is the main ingredient of gunpowder (KNO_3, 75 per cent; charcoal, 15 per cent; sulphur, 10 per cent). Potassium carbonate is used in making hard glass. Potassium hydroxide is employed in the manufacture of soft-soap. Potassium cyanide, an exceedingly poisonous salt with the formula KCN, was long used in extracting gold from its ores and in electroplating.

Now for many purposes sodium salts, which are much more abundant and consequently much cheaper, are just as good as the corresponding potassium compounds. Consequently the shortage in potassium salts, which became very acute during the World War, since the main deposits of potassium chloride are at Stassfurt in Germany, has now been relieved to some extent by educating the industries into substituting compounds of sodium for those of potassium wherever practicable. In some cases the sodium salt has actually turned out to be superior to the potassium salt that it has displaced! Sodium cyanide, for example, is not only less expensive, gramme for gramme, than potassium cyanide, but one gramme molecular weight (49 grammes) of $NaCN$ is just as effective in electroplating, or in gold-extracting, as one gramme molecular weight (65 grammes) of KCN, so that even the quantity required is cut down considerably.

Some compounds of potassium cannot be replaced by sodium salts. Sodium nitrate, for instance, cannot be directly substituted for potassium nitrate in explosives, because it takes up water in damp weather. Similarly sodium chloride will not function as a fertilizer in place of potassium chloride, since plants insist upon plenty of potassium.

Compounds of rubidium and of caesium are merely chemical curiosities, too rare to be of any practical significance. The element that should occupy room F is at present undiscovered, although H. G. Wells has made it the hero of one of his most stimulating books, *The World Set Free*, under the name of Carolinum.

Let us, then, turn to the other occupants of the first floor, copper, silver, and gold. Their association with the alkali metals is not a happy one, since their habits are quite conflicting. Copper, silver, and gold are among the most inactive of all metals; they are very loath to

THE LABORATORY : BY W. HUNT

(often mis-named " Michael Faraday Washing Apparatus for Sir Humphry Davy")

MADAME CURIE IN HER LABORATORY

© *United States Radium Corporation*

CONCENTRATING RADIUM FROM BARIUM RESIDUES BY CRYSTALLIZATION

lose electrons, and their compounds are relatively unimportant and often unstable. Gold, indeed, is such a miser that in a chemical sense it scarcely counts as a metal at all. Copper sulphate is used in small amounts to prevent the growth of organisms in reservoirs and also to kill fungi on grape vines and other plants. The instability of compounds of silver and of gold is taken advantage of in photography.

The three metals themselves, of course, have very numerous and important uses. It would lead us too far astray at present to describe in detail the methods employed in extracting them from their ores and in refining them subsequently.

Ascending, then, to the second floor, we find there one very harmonious family—the alkaline earth metals, calcium, strontium, barium, and radium—and a more heterogeneous group comprising the elements beryllium, magnesium, zinc, cadmium, and mercury. Let us get rid of this second ill-assorted bunch quickly, so that we may linger longer with their more interesting neighbours. The only items worth mentioning in connection with beryllium are that it sometimes tries to masquerade as a non-metal and that it is found, under the name of beryl, in emeralds. Magnesium, zinc, and cadmium are all fairly active metals, but not too active to prevent their being useful in the free state. Mercury, on the other hand, is very inactive. Important compounds of magnesium include the sulphate (Epsom salts) and the oxide (85% magnesia, employed for insulating pipes and boilers). Zinc oxide is used as a filler for rubber in automobile tires and also for filling teeth. Certain salts of mercury are of service in medicine.

The alkaline earth metals approach the alkali metal-in high activity, and consequently dispose of their spare electrons very readily to form stable compounds, many of which are of great importance. Calcium chloride is employed in refrigeration and in the manufacture of cement. Since it attracts water very strongly, the solid salt is scattered on gravel roads in summer to lay the dust. Chalk, limestone, and marble all consist of calcium carbonate, differing only in compactness and in crystalline character. When calcium carbonate is heated, it loses carbon dioxide, and calcium oxide (or quicklime) remains. The hydroxide (slaked lime) is the cheapest and most widely used of all alkalis ; it is employed in hundreds of chemical industries. Calcium sulphate (gypsum) is used as a fertilizer and in making plaster of Paris. Calcium sulphide has the peculiar property of shining in the dark after it has been exposed to sunlight, and this property is now utilized in advertising signs. Strangely enough, the perfectly pure salt does not exhibit this behaviour ; the presence of traces of salts of other metals is necessary. Many other salts of calcium have extensive industrial uses, but we must omit them here. Strontium salts are used in the manufacture of red lights for fireworks and signals, while barium salts give flames a green colour.

We have now reached the most intriguing element of the whole family—radium. A few introductory words regarding its discovery will be opportune.

In 1896 Professor Henri Becquerel of the University of Paris observed that a crystal of a salt of uranium (a rare element extracted from the mineral *pitchblende*) could, in the dark, act upon the silver bromide on a photographic plate, even when a sheet of black paper impervious to light was placed between. Evidently a radiation, different from light, was given out by the salt. Next he discovered that an electrometer (*Fig. 44*), in which the gold leaves had been caused to separate by charging them with electricity, lost its charge rapidly when a salt of uranium was brought near to the knob connected with the leaves. Evidently the salt rendered the air a conductor (or *ionized* the air), and this permitted the escape of the electricity. These discoveries, in the hands of a multitude of observers, have led to the development of an entirely new branch of our science—namely, *radio-chemistry*. Eminent above all others in this field is the name of Madame Curie.

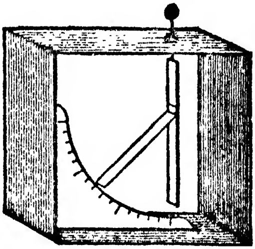

FIG. 44.—Electrometer.

Marie Sklodovska, a Polish refugee stranded in Paris, was first engaged in the physical science department of the Sorbonne to wash bottles (note the resemblance to Faraday !) and to prepare the furnace. Professor Lippmann, the pioneer in colour photography, promoted her to setting up apparatus, and put her to do research work with Pierre Curie, one of his assistants. A very natural result followed. "What a grand thing it would be to unite our lives and work together for the good of science and humanity," said Pierre, so in 1895 Marie became Madame Curie.

In the course of a systematic study of various elements and compounds to determine whether any of them were radioactive like uranium, Madame Curie discovered that a sample of pitchblende from which most of the uranium had been extracted was much more active than uranium itself. After about a ton of the very complex residues had been laboriously analysed, it was found that the radioactivity was now particularly concentrated in the part which, in all ordinary chemical respects, behaved like barium sulphate. From this a product free from barium and more than two million times as active as uranium was finally secured in the form of the bromide of a new metal, which was named *radium.* The chemical properties of this metal were investigated, and its references were found sufficiently satisfactory to entitle it to rent the vacant room F on the second floor of Mendelejeff Court. We are not particularly interested here, however, in the regular chemical behaviour of radium; we want to look into the novel question of radioactivity more closely.

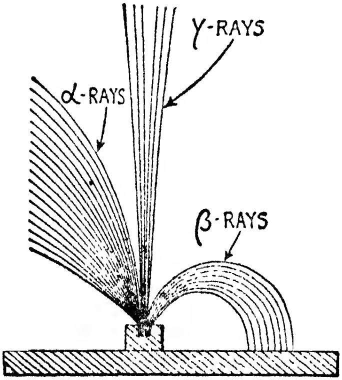

Fig. 45.—Alpha-, Beta- and Gamma-Rays in a Magnetic Field.

The " rays " emitted by compounds of uranium, radium, and other raidoactive elements, such as thorium and actinium, are of three kinds. They are most sharply distinguished from one another when allowed to pass through a powerful magnetic field. The *alpha-rays* are positively charged and are bent in one direction, while the *beta-rays* are negative and are bent in the other. The *gamma-rays* are not affected (*Fig. 45*).

The alpha-rays are atoms of helium, ejected in straight lines with varying initial velocities, depending upon the radioactive element employed, but averaging about one-tenth that of light. Each such atom bears a double positive charge (in other words, it has lost two electrons). The beta-particles are electrons, or unit charges of negative electricity, and are shot out—not just donated, as in ordinary chemical reactions—with a velocity approaching that of light (300,000 kilometres per second). The gamma-rays are identical with X-rays and are presumably produced, like the latter, by the impact of the electrons on the surrounding matter.

The helium atoms are almost all stopped by a sheet of paper or by aluminium foil 0·1 mm. thick. The electrons have greater penetrating power, many passing through gold-leaf, but being practically all blocked by a sheet of aluminium 1 cm. thick. The gamma-rays (X-rays), however, are able to penetrate relatively thick layers of metals and other materials of low atomic weight.

The explanation of this emission of atoms of helium and of electrons from radioactive elements was supplied in 1902 by Sir Ernest Rutherford and Frederick Soddy, then at McGill University, Montreal. These investigators found that the atoms of elements like uranium and radium spontaneously disintegrated into atoms of other elements. When radium emits helium there is also produced another gas, radium emanation or radon. This gas is itself radioactive and undergoes further disintegration, depositing a solid residue, also radioactive, on bodies in contact with it. After long research, the complete story of the series of internal eruptions which the atoms of radioactive elements suffer was disclosed. Radium, it appears, is only an intermediate chapter in a long chronicle which begins with uranium and ends with lead. The accompanying table

(*Fig. 46*) presents in a simplified form the transformations which occur in sequence.

There are several points in this table which require explanation. First of all, if you look at the equations which show the successive steps in the disintegration process, you will notice that a number of " elements "

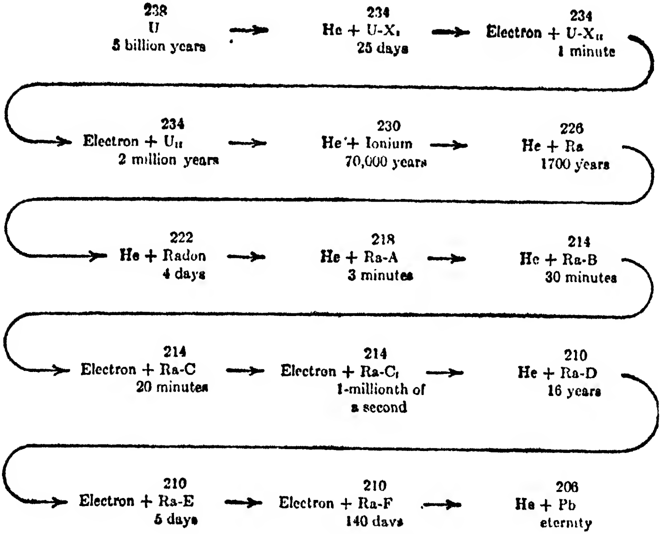

FIG. 46.—Uranium Series of Radioactive Disintegrations.

are included therein which do not apparently find an abode in Mendelejeff Court. The atomic weight of each " element " is placed immediately above its symbol. Do not let the existence of these " elements " and the fact that several of them have the same atomic weight disturb you for the present; we shall explain these matters satisfactorily later on.

Under each element, you will notice, a time is given, varying from one-millionth of a second to eternity. This time gives you an idea of the relative rate at which each step in the process takes place. Only a certain fraction of the atoms of any particular radioactive element disintegrates in a given unit of time. Thus, as the amount which remains diminishes, the amount changing in unit time becomes less and less, and an infinite time would be necessary for disintegration to be complete. For convenience, however, it is customary to calculate for each radioactive element its *half-period*, or the time required for the disintegration of half its original amount, and these half-periods are included in the above table.

It may help you if we explain briefly how the rate at which radium disintegrates can be measured. The apparatus used is called a Crookes' spinthariscope (*Fig.* 47). A sample of radium bromide is placed at B, and some of the ejected helium atoms strike a screen C nearby, which is covered with zinc sulphide, thereby producing faint flashes of light. The lens A magnifies the flashes, which can be seen in a dark room after the eye has become thoroughly rested (15 to 20 minutes are necessary). From the number of flashes noted in a given time, the position of the screen, and the weight of radium bromide taken, the rate of disintegration is easily calculated. The actual figure obtained corresponds to the disintegration of two billion atoms of radium in a gramme of radium per minute. Of course, it is not advisable to start with a whole gramme of radium in an experiment like

FIG. 47.—Crookes' Spinthariscope.

this, since you would have to count the flashes rather rapidly, but by using a very minute sample this difficulty can be avoided. If you look at one of the luminous figures in the dial of a radiolite watch through a strong magnifying glass in the dark, you will probably be able to distinguish the isolated flashes which occur thereon. This is about as near as any one can get at present to seeing single atoms.

Unlike ordinary chemical change, the rate of a radioactive disintegration is not affected by conditions. It can neither be started nor stopped at will. It is no more vigorous at a white heat than at the temperature of liquid air. Other changes occur between atoms, these *within each atom.*

Before we close the chapter one duty must not be avoided. The appalling thought has probably struck you before now : " How can uranium, radium, and other radioactive ' elements ' be called elements when they are continuously decomposing ? " There are two courses open to us in dealing with this difficulty. The easiest way out, of course, would be to deny radioactive substances any claim to the title of " elements." This, however, would scarcely be fair, since uranium and radium have rented rooms in Mendelejeff Court, and their chemical properties indicate that they have a legal right to resist eviction. The only other thing that we can do is to modify Lavoisier's definition of an element, which has served us so well for more than a century, so as not to exclude substances like uranium and radium. Chemists have made several attempts to accomplish this.

The first new definition proposed ran thus : " Substances which we are not able to decompose into, or to build up from, other substances *at will* are called simple substances or elements." The two little words *at will* were considered to dodge the difficulty of radioactive substances, since we can neither hasten, retard, nor stop their disintegration by any change in conditions. Each radioactive substance decomposes at its own sweet will and in its own sweet way.

A few years ago, however, Rutherford made some experiments which shot this definition all to pieces. Rutherford discovered that a number of common elements, such as nitrogen and aluminium, could be disrupted by exposing them to the action of the swiftly moving helium particles ejected from radium-C in the course of its disintegration. One of the products of this disruption was always positively charged hydrogen atoms. Now this decomposition of " elements " like nitrogen and aluminium is certainly brought about *at the will of the experimenter* ; the atoms of these substances are perfectly stable until he decides to bombard them with alpha-rays.

The definition was therefore changed once more to read as follows : " Substances which we are not able *by the ordinary types of chemical change* to decompose into, or to build up from, other substances are called simple substances or elements." The only fault that can be found with this definition is that it suffers from the unfortunate defect that what is regarded as very unusual to-day may be an ordinary experience to-morrow. As we shall see in a later chapter, it is quite possible that methods will be found in the near future to transmute one common element into another by the use of electricity or thermal energy, although enormous advances must first be made to develop the tremendous intensities necessary to effect such transmutations.

We are therefore left, for the present, without any perfectly satisfactory definition. Listen to the words of Frederick Soddy, speaking before the Chemical Society of London in 1918:

> Nemesis, swift and complete, has indeed overtaken the most conservative conception in the most conservative of sciences. . . .
>
> I am not much concerned with definitions, but I think the Chemical Society might safely offer a prize of a million pounds to any one of its members who will shortly and satisfactorily define the element and the atom for the benefit of and within the understanding of a first-year student of chemistry at the present time.

We are going to devote the greater part of the rest of this book to an endeavour to win the million pounds. We do not expect ever to collect the money, but it furnishes us (and you, we hope, also) with an incentive to carry on to the bitter end. Let us pretend that we are modern alchemists, attempting to get gold out of the other elements. In our search we shall, of course, have some rough stretches of road to travel, but we trust that you will make it a point of honour, having accompanied us so far, not to desert us now that we are at last nearing our goal, whatever difficulties we may encounter together in the final stages of our long journey.

Chapter XVI

THE ALCHEMISTS WERE RIGHT

"The great things are discovered," wrote Gerald Brown in his summary of the nineteenth century. "For us there remains little but the working out of detail." . . . And already Chemistry, which had been content with her atoms and molecules for the better part of a century, was preparing herself for the vast next stride that was to revolutionize the whole life of man from top to bottom.

—H. G. Wells, *The World Set Free.*

In an earlier chapter we disposed rather casually of the ancient notion that all matter was composed of four elements—fire, water, earth, and air—and of the belief of the alchemists that one substance could be transmuted into another.

The development of the idea of specific elements and the general acceptance of the atomic theory of Dalton did not, however, entirely displace the original conception of a single parent substance. In 1815 William Prout, an English physician, put forward a definite hypothesis that the atoms of all elements were more or less complex aggregates of atoms of *hydrogen*. According to this view, the atomic weights should evidently all be expressed by whole numbers on the basis $H = 1$. At that time accurate atomic weight determinations were few and far between, but among the common elements there were so many with atomic weights near to whole numbers that most of the great authorities of the period were disposed to accept Prout's hypothesis as correct. As more exact analytical methods were developed, however, the unsound character of the hypothesis gradually became evident, and in 1860 the Belgian chemist Jean Servais Stas, who had himself redetermined many atomic weights with the greatest possible refinements of method and apparatus, apparently wrote its doom in the following words: "I have arrived at the absolute conviction, the complete certainty, so far as it is possible for a human being to attain to certainty in such matters, that the law of Prout is nothing but an illusion, a mere speculation definitely contradicted by experience."

The change to the scale Oxygen = 16, however, brought the question before chemists once more. While it is perfectly true that some of the atomic weights on this scale are far removed from whole numbers, yet it is extraordinary how large a proportion of the common elements possess atomic weights which lie very close indeed to simple integers. Mention may be made of carbon, 12·000; nitrogen, 14·008; and sodium, 22·997, as examples. The number of such elements is much greater on the scale O = 16 than on the scale H = 1. This congregation of atomic weights around simple whole numbers may, of course, be due merely to chance, but Sir William Ramsay, from an examination of the whole list of elements, arrived at the conclusion that the odds were many billions to one against so many atomic weights being *fortuitously* so very nearly integral.

The general acceptance of Mendelejeff's periodic system, as might have been expected, also served to revive interest in the possibility of the genesis of the elements from a single source. Mendelejeff himself, however, frowned upon such speculations and regarded work along this line as an abuse of his ideas. The fact that the fundamental unit had been successively lowered from

1 to 0·5 (to take care of chlorine), then to 0·1 (to include other elements), and so on, as atomic weight determinations became more and more precise, seemed to justify this attitude.

The discovery of the electron and the study of radioactivity finally indicated the solution of the difficulty. It was shown by Sir J. J. Thomson that the beta-ray or electron (the unit of negative electricity) was the same, whatever element it sprang from, its mass being 0·00054 on the ordinary scale of atomic weights (O = 16). For the time being, let us assume that this mass is negligible. Now since atoms are electrically neutral they must contain *something* besides units of negative electricity. The most natural hypothesis is to assume that units of positive electricity also exist, and that such positive units in quantity sufficient to neutralize the electrons must constitute an essential part of the structure of the atom. These " wanted " positive units, however, proved to be by no means so accessible to investigation as the electron. They lurked in the interior of the atom and did not participate in the ordinary types of chemical reactions. After a prolonged search, nevertheless, the unit of positive electricity was at last brought to light—and, surprisingly enough, it turned out to be an old friend. It is, in fact, a hydrogen atom which has lost its lone electron—in other words, a naked hydrogen ion. This unit is now called a *proton* (from the Greek for *first*). Its electrical charge, of course, is equal and opposite to that of the electron, and its mass has been shown by direct measurement to be 1·007 on the oxygen scale. For our present purposes, let us disregard the decimals and assume it to be exactly 1.

The suspicion that the atoms of all elements might be built up of aggregates of protons and electrons led to a tremendous amount of research work on all aspects of the problem, both from the physical side and from the chemical side. Sir Ernest Rutherford first formulated in 1911 a number of principles regarding atomic structure which are still the basis of our current views. The atom of any element was postulated as consisting of a very minute central nucleus, positively charged, surrounded by planetary electrons. These planetary electrons are at relatively large distances from the nucleus, just as the planets are at relatively large distances from the sun in our solar system. The central nucleus contains *all* the positive electrical units or protons and is consequently responsible for practically the entire *mass* of the atom. The external electrons, on the other hand, are responsible for practically the entire *volume* occupied by the atom. Only in radioactive transformations does the nucleus change; ordinary chemical reactions affect merely the external electrons.

With this model as a basis, the most wonderful advances ever made in the history of our science have resulted during the last seventeen years. For a time the difficulty of fractional atomic weights was disregarded, since such rapid strides were being made in other directions, but now a complete explanation of this point also has been obtained. It will pay us to consider here certain experimental results which encouraged chemists to ignore this difficulty temporarily.

The atomic weight of ordinary lead is always 207·20, whatever the source of the material employed in the analysis. In 1914, however, Professor Theodore William Richards of Harvard University found that the lead present in uranium ores, and presumably resulting from the decomposition of the uranium series of radioactive elements, possessed an atomic weight which was significantly smaller than 207·20. Samples of lead extracted from the pitchblende mines in Bohemia, indeed, gave values as low as 206·05. On the other hand, Soddy discovered that lead resulting from the decomposition of radioactive elements of the thorium series might have an atomic weight as high as 208·4. All of these forms of lead were *identical* chemically! This was taken to indicate that atomic weight is, after all, *not* a fixed and fundamental property and that some explanation of its variability might ultimately be arrived at which would bring elements like chlorine (all natural samples of which give the very awkward value of 35·457) into correspondence with a proton-electron structure.

After all, chemists had already had good warning that atomic weights were not infallible. If you refer back to the picture of Mendelejeff Court, you will notice three cases where the natural order of ascending atomic weights is disregarded. Argon (39·9) is placed before potassium (39·1). Cobalt (58·9) comes ahead of nickel (58·7). Tellurium (127·5) anticipates iodine (126·9). In each of these three instances the properties of the elements are such as to necessitate reversing the customary order. It would clearly be absurd to plant potassium among the inert gases, or argon among the alkali metals. As we have seen, Mendelejeff succeeded in explaining other apparently similar cases of this kind by showing that the expected atomic weights were in error, but in spite of all the painstaking work that has been done to bring these three pairs of elements into line (the researches on tellurium alone would fill many volumes), the figures persist in running the wrong way.

If atomic weights, then, which Richards once characterized as " the most significant set of constants in the universe," are not a fundamental property of the elements, what can we use in their place ? The answer to this question was supplied through the work of a brilliant young English physicist, H. G. J. Moseley, who introduced the idea of *atomic numbers.* Moseley's promising career was unfortunately cut short by a Turkish bullet in the trenches at Gallipoli, a tragedy in science equalled only by the death of Lavoisier. Orders for his recall from active service to undertake important scientific war work at home arrived just too late to prevent the unnecessary sacrifice of his genius.

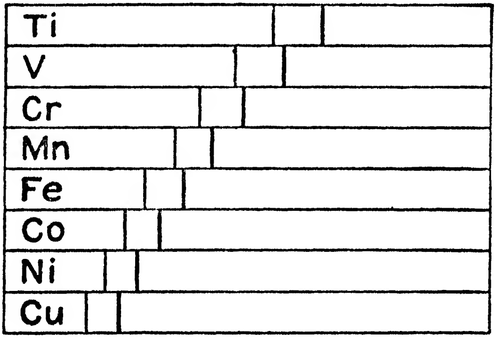

FIG. 48.—K Series of X-ray Spectra (diagrammatic only).

Moseley's work requires a brief preliminary introduction. We have already discussed in an earlier chapter how X-rays may be used to analyse the arrangement of atoms in crystals. The fact that the atoms in crystals act as a diffraction grating for X-rays may be utilized also in the reverse direction. As we have seen, X-rays are produced in an evacuated tube (*Fig. 36*), through which an electrical discharge is passed, when the streams of electrons emanating from the cathode C (cathode rays) strike the anti-cathode A. With different elements on the anti-cathode, X-rays of slightly different wave-lengths, and therefore giving different X-ray spectra when analysed, are produced. The characteristic radiations for each element appear in two distinct series, known as the K (hard) and L (soft) series, but each series is made up of a strong line and a weaker one a little distance from it, as shown in *Fig. 48* for a number of successive elements in the Mendelejeff periodic table, ranging from titanium to copper. There is evidently some simple connection between the position of these lines and the position of each element in the periodic system, but if we attempt to correlate the wave-lengths of the lines with the atomic weights of the elements, we run at once into trouble. This will be immediately evident from the fact that the atomic weights of nickel and of cobalt are in the wrong order, whereas their spectra step up uniformly with those of the other elements.

Atomic weights, then, which show *very divergent* intervals between successive elements, are not of service in elucidating the *regular* variation in the position of their X-ray spectra. Moseley showed, however, that the position of these lines can be related directly, by calculations founded on strict theory, to *atomic structure*. The simple rule thus brought to light states that the greater the positive charge on the nucleus of the atom, the shorter should be the wave-length of the characteristic X-rays. Indeed, when the elements are arranged in the order of the periodic system, successive whole numbers (*atomic numbers*) can be assigned to each which are inversely proportional to the square roots of the wave-lengths of corresponding lines in their X-ray spectra (see *Fig. 49*). You will understand what this diagram means, perhaps, a little more clearly from the statement that increasing the atomic number four times means decreasing the wave-length to one half.[1]

This law is of use in several ways. In the first place, every detail in the spectrum of an element can be predicted from the spectra of its neighbours. In the second place, when we list the elements in order of increasing nuclear charge, we find this most important fact to be true. Starting with hydrogen (atomic number, 1), we can call the roll of all the elements which occupy Mendelejeff Court, and every known element up to and including uranium will answer promptly, " 2," " 3," " 4," etc., up to " 92," according to its position on the list. Only two elements, with atomic numbers 85 and 87, are missing. The atomic numbers thus obtained for each element are

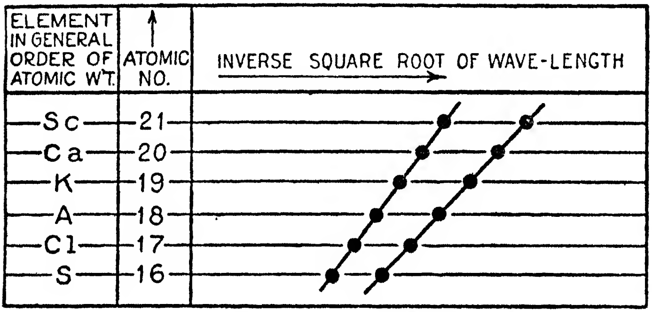

FIG. 49.—Partial Illustration of Moseley's Law.
The two dots opposite each element are derived from the two lines in the K series.

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[1] The square root of four is two, and the inverse ratio changes this two to one-half.

included in the table in the Appendix. These numbers, it may be mentioned, have been confirmed by other independent methods.

The superiority of atomic numbers over atomic weights may not be apparent to you at first, but as we proceed you will gradually be convinced that *the atomic numbers determine all the properties of each element and are more fundamental than the atomic weights.* The latter are secondary properties, in most cases modified by other factors and in a few cases actually thrown out of order by such factors.

It is important that you should remember very clearly what the atomic numbers mean. *They represent the nuclear charge on the atom of each element.* The hydrogen atom, with a nucleus consisting of one single proton, has an atomic number of 1. As soon as we come to helium, however, the situation gets a little more complicated. The nucleus of the helium atom obviously cannot be completely represented by two protons, since the atomic weight of helium is four. The answer to this riddle is reached when we recognize that the helium nucleus, besides its two *free* protons, also contain two extra protons associated, or *bound*, with an equal number of electrons. The complete " picture " of the helium atom is therefore as given in *Fig. 50.* The alpha-ray in radioactive disintegrations is the nucleus of this helium atom without the two external electrons.

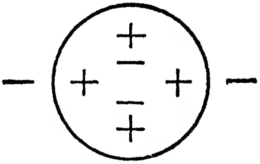


FIG. 50.—Helium Atom (diagrammatic only).

A similar state of affairs occurs in the atoms of all the heavier elements. We have a number of *free protons* in the nucleus represented by the atomic number, and we have a number of *bound protons*, associated with *electrons*, equal to the difference between the atomic weight and the atomic number. (For the present, remember, fractions are being disregarded.) Since the atom as a whole is neutral, we have, finally, a number of *external electrons* equal to the number of free protons in the nucleus. An example may serve to illustrate all of these points. The atomic weight of sodium is 23; its atomic number is 11. The sodium atom, therefore, has a nuclear charge of 11. It contains 11 free protons in the nucleus. Its nucleus also contains 23 minus 11, or 12, protons bound up with an equal number of electrons. And, finally, there are 11 electrons outside the nucleus to make the picture complete.

We are now almost ready to face the difficulty of fractional atomic weights, such as 35·457 for chlorine. In order to indicate the way in which this difficulty has been solved, however, we must return for a moment to the series of radioactive disintegrations which we studied in the last chapter. This series begins with uranium and ends with lead, and at each step a helium nucleus or an electron is ejected from the atom. Now, when an atom loses a helium nucleus, it also loses two free positive charges, and its atomic number is consequently *reduced by two*. For example, the atomic number of radium is 88 and that of radon is 86. On the other hand, when a radioactive change takes place involving the loss of an electron, a positive charge in the nucleus, previously bound, is set free and the atomic number is therefore *increased by one*.

With these facts in mind, an examination of the uranium disintegration series (*Fig. 46*) discloses that several elements (for example, radium-B, radium-D, and lead) must exist which possess *the same atomic numbers but different atomic weights.* To get from radium-B to radium-D, or from radium-D to lead, we lose two electrons and one helium nucleus, so that the atomic number is unchanged. You will be able to distinguish other examples of the same type if you go through the whole table.

Elements which have the same atomic number but different atomic weights are known as *isotopes.* The one essential fact regarding isotopes is that they are *identical in all of their chemical properties*, although they differ in atomic mass. This shows conclusively that atomic number and not atomic weight is a fundamental property.

So long as this matter of isotopes remained restricted to radioactive elements it was obviously an interesting theoretical point but of no everyday significance. Recently, however, it has been shown by Dr. F. W. Aston of the Cavendish Laboratory of Cambridge University that *many common elements also are isotopic*, or contain chemically identical atoms of different weight. The method employed by Aston to demonstrate this was that of " positive ray analysis."

It was discovered in 1886 that rays with a positive charge travel through a Crookes vacuum tube in the reverse direction to the more familiar cathode rays. The nature of these rays was very fully investigated by Sir J. J. Thomson, who showed that in their simplest form they were atoms which had lost an electron. By introducing small quantities of any element which we wish to examine into a tube of the Crookes type, rays consisting of positively charged atoms of this element can be obtained.

Aston used the apparatus shown in *Fig. 51.* The positive rays from the element under study are sorted out into a thin ribbon by means of the two parallel slits S_1 and S_2, and are then passed between the oppositely charged plates P_1 and P_2. Because of their positive charge, the rays are deflected toward the negative plate P_2 and are spread out into an electric spectrum. A portion of this spectrum deflected through a given angle passes through a narrow opening in the diaphragm D and then travels between the circular poles of a powerful electromagnet O, the field of which is such as to bend the rays back again to fall on a photographic plate placed as shown. If all the rays with a single positive charge have the same mass, they will converge to a focus at F. If, however, the rays are derived from an element which consists of a mixture of isotopes, each isotope is distinguished by a separate band on the photographic plate, and from the positions of these bands the mass of the atom to which each corresponds can be deduced.

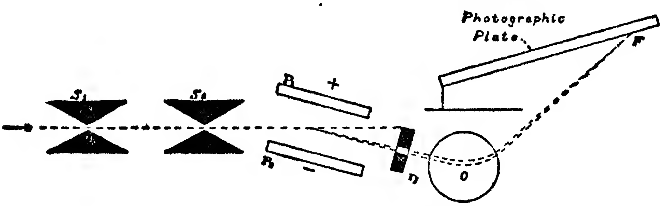


FIG. 51.—Aston's Mass-Spectrograph.

Chlorine (atomic weight, 35·457), examined in this way, showed itself to be a mixture of two isotopes with atomic weights *exactly* 35 and *exactly* 37. Bromine (atomic weight, 79·92) gives isotopes with atomic weights *exactly* 79 and *exactly* 81. Argon (39·9) is mostly 40, with a trace of 36; potassium (39·1) is mainly 39, with a very little of 41. The lighter species of argon is lighter than the lighter species of potassium ; the heavier potassium is heavier than the heavier argon. The inverse order of the atomic weights of these two elements in the periodic system is therefore completely explained—it is due merely to the unfair way in which the proportions of their two isotopes are allotted by nature. Again we see how atomic weights are delusive, and inferior to atomic numbers in their significance.

Mercury (atomic weight, 200·6) appears to exist in as many as six forms, with atomic weights ranging from 197 to 204. Tin (atomic weight, 118·7) is a mixture of no fewer than eleven isotopes, with atomic weights ranging from 112 to 124 ! Many other elements show similar diversities, but some—such as oxygen, nitrogen, and iodine—give no indications of isotopes.

The fundamental atomic weights obtained by Aston seemed at first to be all whole numbers, within the limits of experimental error. Subsequent measurements with more delicate methods of observation have shown that slight deviations from whole numbers do exist in many cases, but still the only glaring exception among the lighter elements is hydrogen (atomic weight, 1·008). It looks as if we are straining at a gnat when we call this an exception, but the deviation of 8 parts in 1000 is really of vital importance, as we shall show later.

All elements, then, which apparently have a fractional atomic weight very far removed from a whole number actually consist of a mixture of two or more species of atoms, each species having an integral atomic weight. Thus our girl-friend, chlorine, whose unruly nature is shown even in her atomic weight, 35·457, turns out after all to be " chemical twins," one a perfect 35, the other a perfect 37. The atomic number of chlorine is 17, so that we may picture these two kinds of chlorine as follows :

Atomic Weight.	Nucleus.	Exterior.
Chlorine 35	17 free protons ; 18 bound protons and 18 electrons	17 electrons
Chlorine 37	17 free protons ; 20 bound protons and 20 electrons	17 electrons

Why does all natural chlorine check up as 35·457 ? Because all samples contain the two isotopes in exactly the proportions necessary to give this average value. Why cannot we separate them ? Chemists are beginning to achieve success in this direction, but it is an exceedingly difficult task, since isotopes have identical chemical properties and differ only very slightly indeed in certain physical properties. By diffusing nearly a ton of gaseous hydrogen chloride through a long series of porous porcelain tubes (the stems of Scotch churchwarden smoking-pipes), Professor W. D. Harkins of the University of Chicago has succeeded after several years' labour in " fractionating " the substance so that one small sample contained chlorine with an average atomic weight as low as 35·42 and another small sample contained chlorine with an average atomic weight as high as 35·52. If this diffusion could be continued long enough—several geological eras would probably be required—a complete separation would undoubtedly be effected. Then it would be possible, when your neighbour at dinner asked you to pass the salt, to inquire, " Which kind do you prefer, the perfect 35 or the perfect 37 ? " Both kinds, however, would possess exactly the same properties ; the only difference would be that one would contain several per cent less chlorine than the other. It is possible, of course, that some ingenious trick will be devised to separate isotopes more rapidly in the near future, but present indications are not hopeful.

Why do isotopes possess identical chemical properties ? Because chemical properties depend only upon the exterior electrons, and these are the same in number for the different isotopes. Ordinary chemical reactions do not involve the nucleus at all, and it is only in the nucleus that the isotopes differ.

What effect do the discoveries discussed above have upon our old beliefs—Dalton's atomic theory and the law of definite proportions, for instance ? Of course, we have to reframe our definitions somewhat ; but as we shall find in the next chapter, we do not need to cast away anything of importance.

Let us conclude this chapter by calling your attention again to the fact noted in its title—namely, that *the alchemists were fundamentally right.* The only error they made was that they did not go far enough when they switched from the four " elements " of Aristotle—fire, water, earth, and air—to the three " principles "—mercury, sulphur, and salt. They should have reduced the number of essential substances for the building-up of all forms of matter to two—*protons* and *electrons.* If we could only pick atoms to pieces and rearrange their protons and electrons as we liked, we would obviously be able to transmute *anything* into gold ! How far this dream is from realization at the present time will appear in due course.

Chapter XVII

REWRITING THE CONSTITUTION

> And we know now that the atom, that once we thought hard and impenetrable, and indivisible and final and—lifeless—lifeless, is really a reservoir of immense energy. That is the most wonderful thing about all this work. A little while ago we thought of the atoms as we thought of bricks, as solid building material, as substantial matter, as unit masses of lifeless stuff, and behold! these bricks are boxes, treasure boxes, boxes full of the intensest force. . . . The energy we need for our very existence, and with which Nature supplies us still so grudgingly, is in reality locked up in inconceivable quantities all about us. We cannot pick that lock at present, but—we will.
>
> —H. G. WELLS, *The World Set Free.*

THE atom, which was formerly regarded as sacred and placed in the holy of holies of the chemical temple, has now lost a great deal of its mystery. We have, to use a slang phrase, " got its number." We know that it is by no means all it was once cracked up to be—one, indivisible, indestructible, eternal, and so on. Atoms of radioactive elements explode of their own accord ; atoms of common elements can be disrupted by the use of radioactive energy. Indeed, the jocular proposal has recently been made that the atom, which was so named by Democritus because he considered it to be indivisible (Greek, *a-* not, *tom* divided), be now re-christened "Tom" to celebrate its dismemberment. This suggestion is, perhaps, not so foolish as it sounds, since we do need some such change of title to remind us that the atom, which formerly was dead, is now *alive.*

So many of our old ideas have fallen by the wayside in the course of the last two chapters that it will be well for us to stop for a time and take stock. We began our chemistry by defining the term substance. Let us see, then, first of all, what has happened to this concept during our recent shake-up.

If you want to have some fun, go to one of your friends who fancies himself a chemist and ask him whether chlorine, with its two isotopes, is one substance or two. Whatever answer he makes you can tell him that he is wrong, because ordinary gaseous chlorine is now *three* substances. The reason for this is that the chlorine molecule Cl_2 may be made up either of two atoms with mass 35, or of two with mass 37, or of one of both kinds. Each of these three types of molecule differs in certain physical properties such as density and rate of diffusion. The chemical properties, however, are identical. In the same way radioactive isotopes, like radium and mesothorium (a step in the thorium disintegration series), are identical in their ordinary chemical properties, but their rates of disintegration are quite different. The half-period of radium is 1700 years, while the half-period of mesothorium is only about seven years. Mesothorium is much cheaper than radium and consequently is used instead of radium on the luminous dials of watches. Its effectiveness does diminish noticeably with time, but nobody expects to carry an Ingersoll for 1700 years.

Perhaps you would like us to give a new and improved definition of a substance to take care of all these facts. We respectfully beg to be excused. Why should we run the risk of arousing the derision of our colleagues when there is no prize of a million pounds offered for this particular feat?

Seriously, there is no need of a new definition. It is only natural that the more we get to know about things (or persons), the harder we find it to formulate a short and satisfactory outline of their salient characteristics. We can make very definite statements regarding people whom we have met casually, or even regarding people whom we have never met at all, but as soon as we start to discuss our intimate friends we find that all sorts of qualifying clauses begin to enter into our analysis. So here; substances are no longer strangers to us but bosom companions. We know much more about them than previously, but we cannot put all of our knowledge into a definition that is both brief and convincing.

Let us consider next the second stride in our early progress, the law of definite proportions. It looks at first as if this law would have to be discarded completely, since three different compositions have already been obtained experimentally for lead chloride, depending on the source of the lead, and Harkins has now started to supply us with two kinds of sodium chloride containing different percentages of chlorine.

The change in our point of view, however, is not so revolutionary after all. The law of definite proportions does not have to be rejected as useless; it is rather expanded into a wider generalization which includes both the old facts and the new. Our insight into the interior structure of the atom, indeed, really serves to make our understanding of the law of definite proportions much clearer and more concrete. Although the compositions of the two possible sodium chlorides differ, each contains one atom of sodium to one atom of chlorine, and instead of a complicated weight relationship of 22·997 to 35·457 for the two elements, we now have a simple integral relationship between the number of protons and electrons in the constituent atoms of each variety. In one kind this relationship is 23 to 35; in the other it is 23 to 37. The fractions have vanished entirely, and we have now not only a law of definite proportions but also a law of *integral proportions.* The laws of multiple proportions and of combining proportions are similarly expanded.

We approach now the crux of the whole argument, the definition of the term "element." The fundamental nature of the changes in our ideas with respect to the element will have become apparent to you in the course of the last chapter. Elements are *not* incapable of decomposition, for the fact that their atoms are all built up of the same protons and electrons shows that it is *possible* to transmute any element into any other, provided the proper procedure is followed to effect the necessary rearrangement. We have not yet discovered how to bring about any building-up of heavier atoms from lighter ones, and only a few atomic disintegrations on a very small scale (such as those performed by Rutherford on nitrogen and aluminium with the help of the alpha-ray ejected from radium-C) have been carried out under human control, but that is merely a consequence of our lack of skill and not a sign that ordinary "elements" cannot be converted into one another.

It is certain, at any rate, that their atoms are not *indivisible.* Furthermore, the atoms of the same element are *not* all necessarily similar to one another, since they can differ in mass. Not only may atoms of the *same* element have *different* masses, but atoms of *different* elements may possess the *same* mass! Selenium for example, with an atomic weight 79·2, is found by Aston's mass-spectrograph to consist of a mixture of six isotopes, with atomic masses 74, 76, 77, 78, 80, and 82 respectively. Krypton, with an atomic weight 82·9, also consists of a mixture of six isotopes, with atomic masses 78, 80, 82, 83, 84, and 86 respectively. The three heavier isotopes of selenium, therefore, are identical in mass with the three lighter isotopes of krypton. Atoms of different elements which possess the same mass are called *isobares.* Examples are very common among the radioactive elements, as we have already found in our uranium disintegration series.

Although isobares have the same mass, they do *not* have the same chemical or physical properties; in fact, they differ absolutely in every respect *except* atomic mass. That this is the case is due to the fact that the arrangement of electrons in their nuclei and in the exterior parts of their atoms are essentially different. The *atomic number* of selenium is 34 and that of krypton is 36. To take a specific example, then, an atom of selenium with mass 78 contains 78 protons and 44 electrons in its nucleus, and 34 electrons in its exterior; while an atom of krypton with the same mass contains 78 protons and 42 electrons in its nucleus, and 36 electrons in its exterior. We have already seen that it is only the outermost electrons which are concerned in ordinary chemical reactions and which, consequently, determine the chemical properties of the atom. Therefore, since the number of loose electrons is different in isobares and the same in isotopes, isobares are quite diverse in chemical behaviour, while isotopes are identical.

The above illustrations suggest a tentative definition of the term element in accordance with the demands of modern theory. An *elementary substance* may be defined as a substance which consists of atoms all possessing the same atomic number, or nuclear charge. A *compound*, on the other hand, consists of atoms possessing different nuclear charges, in definite and simple numerical ratios.

We do not know at present whether the above definition will be considered worthy of the award of a million pounds. In any case, our duty is but half done, since we still have to define the term " atom." It is hardly sufficient to state that an atom is an aggregate of protons and electrons, and we shall wait until we have studied the way in which these arrange themselves in different elements more intimately before we venture to claim the money.

We come then, finally, to the two laws of the conservation of mass and the conservation of energy. Let us start on these by considering some of the new features of energy involved in radioactive transformations. As we have stated previously, we cannot yet control in any way the *rate* of disintegration of radioactive elements. If, however, some method of inducing or hastening radioactive changes on a large scale is discovered in the future, a wonderful new source of power will be put into our hands—namely, *atomic energy*.

The energy change in radioactive disintegrations is enormously greater than in ordinary chemical reactions. The heat evolved from one gramme of radium through the emission of rapidly moving alpha- and beta-particles and their stoppage by collision with surrounding matter amounts to about 120 calories per hour, and this would continue at a gradually decreasing rate for centuries. The total heat available would be over 2,000,000,000 calories per gramme, whereas a gramme of carbon burning to carbon dioxide gives only 8070 calories. The disintegration of a pound of uranium salts would furnish enough power to drive an ocean liner across the Atlantic, but five billion years is entirely too long to wait for the completion of the trip. Chemists are already looking forward, nevertheless, to the possibility of using the enormous stores of energy here available so soon as means of hastening its emission are devised.

Another interesting by-product of this subject is the calculation that the heat given off by the disintegration of the radioactive elements existent in the earth (radon is found in the soil and in well waters) is sufficient *alone* amply to account for the maintenance of its temperature. A globe the size of the earth, possessing originally only heat energy and cooling from a white-hot condition to the temperature of interstellar space, would have passed through the stage of habitable temperatures in a much shorter time than that which geological deposits and fossils show to have actually elapsed. The discovery of the enormous but gradually released disintegration energy of elements such as uranium and radium enables us now to explain the prolonged period during which life has existed on the earth.

This energy is derived *from within the atom itself*, by rearrangements of the *protons* and *electrons* of which it is constructed. How enormous the energy involved in such rearrangements may be can best be illustrated by considering what would happen if we could transmute hydrogen, our simplest element, into helium, the next simplest. The hydrogen atom consists of one exterior electron and one proton in the nucleus. The helium atom consists of two exterior electrons, with two free protons and two bound proton-electron pairs in the nucleus. The helium atom therefore contains the same material as four hydrogen atoms, *arranged in a different way.* But the mass of the helium atom is not $4 \times 1{\cdot}008 = 4{\cdot}032$; it is 4·00. By rearranging our protons and electrons we have lost 0·032 units of mass. How is this possible?

The answer is: *What we have lost in mass we have gained in energy.* The fact that all matter has an ultimate electrical structure (protons and electrons) forces us to conclude that matter and energy are not distinct phenomena, but interchangeable. Matter is potential energy; energy is potential matter. The law of conservation of mass holds within our limits of weighing, for all ordinary chemical reactions, only because the energy changes involved are extremely small. But radioactive disintegrations have already demonstrated to us that energy changes in reactions " within the atom " may be enormously great. Correlating our laws of conservation of mass and conservation of energy into one wider law—*conservation of mass-energy*—we can calculate what change in mass corresponds to any particular energy change, and vice versa. We shall not bother you with the details of these calculations, but shall merely state our results. We find that the mass change in all ordinary chemical reactions is infinitesimal, and that even in radioactive disintegrations it is exceedingly minute. Conversely, if the mass change is appreciable, as in the case of hydrogen-helium, the energy change must be stupendously large. The transmutation of one gramme of hydrogen to helium, indeed, would liberate energy equivalent to (in terms of heat) 166,000,000,000 calories or (in terms of work) 200,000 kilowatt-hours.

Here at last is a source of energy sufficient to account for the heat of the sun, and if this tremendously exothermal reaction, once started by some clever chemist, were to run out of control and proceed explosively, all of the hydrogen on the earth might be transformed at once and the success of the experiment signalled to the rest of the universe as a new and exceedingly bright star!

Fortunately, it is not very likely that this will occur in the near future. Claims of a successful transmutation of one common element into another have been made by several investigators, but on careful repetition of their work it has always been found that each claim was premature. Thus the conversion of hydrogen into helium was asserted many years ago by Ramsay, and more recently by Paneth and Peters. Ramsay also vouched to have effected the transmutation of copper to lithium, while several workers have lately published preliminary notices regarding the change of mercury into gold. The trouble with all this kind of work is that only exceedingly minute quantities of transmuted material are to be expected under the best conditions, and our chemical tests are now so delicate that they are far too likely to deceive the over-optimistic investigator. The infinitesimal amount of helium which will leak from the outside air through a sealed glass bulb into its interior is sufficient to give a spectroscopic test for helium. The trace of lithium in cigarette smoke is enough to lead to a positive test for lithium, and it is suspected that the gold obtained from mercury really came from the gold-rimmed spectacles worn by the experimenter! While scientific inventiveness will no doubt finally succeed in demonstrating actual transmutations, yet it is to be expected, for a long time at least, that these will be only on a very small scale and very expensive. Synthetic gold, therefore, is not a problem which the economist needs to worry about at the present moment. The practical utilization of atomic energy is also, in all probability, a vision of the distant future.

Why should the rearrangement of the protons and electrons which occurs when hydrogen is transmuted into helium involve such a stupendous energy change? This will be a very profitable question for us to examine, since it will enable us incidentally to clear up most of the points which we have left hanging over during the discussion in the last two chapters.

Small as the atom is, the electron is only a minute speck in its centre. One writer has likened the atom to a cathedral, and on this basis the nucleus would correspond to a bee buzzing about the centre aisle. When we shoot at an atom with bullets of alpha-rays, only one in many million shots hits the nucleus. Now it is in this almost inconceivably restricted space occupied by the nucleus that all of the protons are congregated. But positively charged units of electricity should repel one another, and if the nucleus contained nothing but protons the atoms of all elements except hydrogen (where only one proton is present) would have only the most transitory existence. The salvation of such atoms lies in the fact that the nucleus in all cases except hydrogen also contains a number of electrons. These electrons have very aptly been styled *cementing electrons*, since their function is to hold the discordant protons together. The number of cementing electrons is never less than one-half the number of protons in the nucleus. In the lighter elements, as in the case of helium (*Fig. 50*), it is just one-half; as we go up the series the ratio tends to get a little higher. Nevertheless, the difficulty of holding the protons in the congested nucleus becomes ultimately too great for the cementing electrons to be perfectly successful in their task, and we find that the atoms of the elements of highest atomic weight are all likely to explode spontaneously —in other words, they are radioactive. Elements with a higher atomic weight than uranium are presumably too unstable to exist at all.

The fact that protons and electrons are packed together in such a narrow compass in the nucleus is evidently going to put them under a considerable strain, and this strain will necessitate so great a change in their energy that it will even affect their mass appreciably. The mass of all atoms except the hydrogen atom will consequently be modified by what has been called a *packing effect.* The early work of Aston indicated that this packing effect was practically constant throughout the whole periodic system, since all of the fundamental atomic weights on the basis oxygen = 16 seemed to be exactly integral. Only in the case of hydrogen, where the packing effect was absent, did a fractional value (1·008) appear.

In view of this you will see how fortunate we have been in our decision to make oxygen = 16, rather than hydrogen = 1, our fundamental basis for atomic weights. Since all atoms are built up out of protons and electrons, and since the hydrogen atom contains only one of each kind, it would have been only natural for chemists to have considered returning to a hydrogen basis *but for the existence of this packing effect.* It is rather unfair to hydrogen that it should be the only element to which a basic atomic mass is assigned which deviates significantly from a whole number, when the hydrogen atom is really the only atom which acts normally, but justice has to give way to convenience in this case as in many others.

The later and more accurate work of Aston has demonstrated, it must be admitted, that the packing effect is *not quite* constant. It varies very slightly and in a regular way as we go up the periodic system, so that, after all, atomic masses may diverge from whole numbers by very small amounts. This will probably serve to relieve the feeling of resentment rankling in the breast of hydrogen, but it is only a matter of minor importance to us.

We have tarried a long time on the second floor of Mendelejeff Court, but our visit with radium was certainly well worth while. We shall now proceed to ascend to the third floor.

Chapter XVIII

THE THIRD FLOOR—AND A HURRIED EXIT

" AL'S HERE ! "
—DON MARQUIS, *The Old Soak.*

AL, in the play, is a bootlegger of a retiring disposition whose presence off-stage is invariably indicated by an impressive whisper. Our chemical Al, aluminium, had until recently the same habit of always keeping in the background. As a matter of fact, only forty years ago, if you wanted to see a piece of aluminium, you had to go to a museum and look at a small sample of it shut up in a bottle on a shelf with other chemical curiosities. Nowadays, however, aluminium is not a curiosity ; it is not even a luxury ; it is a necessity. We are not referring here merely to its familiar use in kitchen utensils ; without aluminium and alloys containing aluminium, the modern aeroplane and modern dirigible would be non-existent.

Davy attempted to isolate aluminium by the same methods that he employed for obtaining other metals, but failed. In 1827, however, the German chemist Friedrich Wöhler heated aluminium chloride with potassium and achieved the following displacement :

$$AlCl_3 + 3K \rightarrow Al + 3KCl.$$

The metal obtained in this way cost about £30 a pound. Henri St. Clair Deville in 1854 reduced the price to £3 by employing the cheaper metal sodium instead of potassium, and in 1886 Hamilton Y. Castner, a young student at Columbia University, invented a new process for manufacturing sodium which lowered the cost of aluminium to sixteen shillings a pound. Aluminium has so many valuable properties that a large plant was immediately erected to prepare the metal on a commercial scale, using Castner's process, but these plans came to an end when an electrical method was discovered by Charles M. Hall by means of which aluminium compounds could be decomposed directly at a cost of only four shillings a pound.

Hall was born in Ohio in 1863 and graduated from Oberlin College in 1885. He gained his first knowledge of chemistry from reading a textbook which his father studied at school. It lacked both cover and title-page, so he never knew who the author was, but the book said something about aluminium that interested him. After graduation he came across the statement that every clay bank was a mine of aluminium, but that the metal itself was as costly as silver. He soon began to think out processes for making aluminium cheaply. Most of the apparatus which he used in his first experiments on an electrolytic method of preparation was borrowed from his old professor. To make his battery-cells, however, he had to chop the wood and cast the zinc plates with his own hands. In a few months this boy, working under such handicaps, discovered and perfected what had escaped Wöhler, Deville, and many other world-renowned chemists who had been busy upon aluminium over a period of half a century.

Hall's process, which was patented in 1889, depends on the electrolysis of the oxide (bauxite) in solution in the molten fluoride (cryolite). Iron tanks (*Fig.* 52) are lined with a compressed mixture of coke and tar which is afterward baked. This carbon lining forms the cathode, while the anodes consist of a series of rods of carbon about three inches in diameter which are attached to copper clamps. The cryolite is first melted (1000°) by the arcs struck by the carbon rods when the current is applied. The anodes are then raised somewhat, the aluminium oxide is added, and some coal (which floats) is thrown in to cover the surface and obscure the blinding glow due to the liberation of oxygen at the carbon anodes. From time to time more of the oxide is added and the molten aluminium is tapped off from the floor of the cell.

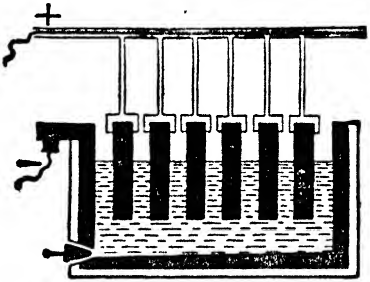


FIG. 52.—Manufacture of Aluminium.

Fifty years ago the whole world's production was only about forty kilogrammes (less than a hundred pounds) a year. To-day Great Britain alone consumes 40,000 metric tons annually, the world's production being approximately five times that quantity.

The metal melts at 660°, but is not mobile enough to make castings. It is exceedingly light, and in tensile strength excels the other metals, with the exception of iron and copper. It is malleable, and the foil is taking the place of tin-foil for wrapping foods. It has a silvery lustre, and tarnishes very slightly, the firmly adhering film of oxide first formed protecting its surface. Although, comparing cross-sections, it is not so good a conductor of electricity as is copper, yet *weight for weight* it conducts better. It is difficult to work on the lathe or to polish, because it sticks to the tools, but the alloy with magnesium (about 2 per cent) called *magnalium* has admirable qualities in these respects. *Aluminium bronze* (from 5 to 12 per cent aluminium with copper) is easily fusible, has a magnificent golden lustre, and possesses mechanical and chemical resistance exceeding that of any other bronze. The metal and its alloys are used for making cameras, opera-glasses, cooking utensils, the metal parts of dirigibles and aeroplanes, and other articles requiring lightness and strength, as well as for the transmission of electricity. The powdered metal, mixed with oil, is used in making a silvery paint.

In consequence of its very urgent desire to lose electrons, aluminium displaces from their oxides the metals below it in the electromotive series. Thus, when a mixture of aluminium powder and ferric oxide (*thermite*) is placed in a crucible and ignited by means of a piece of burning magnesium ribbon, aluminium oxide and iron are formed:

$$Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe.$$

A temperature of 3000° to 3500° is produced by the action, the molten iron collects at the bottom, and the immiscible molten aluminium oxide floats to the top. This very simple method of making pure specimens of metals like chromium, uranium, and manganese, whose oxides are otherwise hard to reduce, is called by Goldschmidt, the inventor, *aluminothermy*. By preheating the ends of steel rails with a gasolene torch, firing a mass of thermite in a crucible above the joint, and allowing the iron to flow into the joint, perfect welds are made. In the same way large castings, like propeller shafts, when broken, can be mended.

The bulk of the aluminium now manufactured is used by steel-makers. When added in small amount (less that 1 part in 1000) to molten steel, it combines with the gases dissolved therein, and gives sound ingots free from blow-holes.

Salts of aluminium and minerals of which it is a part are of very great importance in commerce. We need only mention here the fact that building materials, such as brick, cement, and concrete, are essentially compounds of aluminium. The china and porcelain trades are also aluminium industries.

The other occupants of the third floor of Mendelejeff Court are by no means so significant to us. Boron, the only non-metal, is of some interest in connection with boric acid H_3BO_3 (used as an antiseptic) and borax, a complex sodium salt of this acid (used in water-softening and in making enamels). The metals which complete the list of trivalent elements are all more or less museum specimens, and most of them have no special uses either in the free state or in combination.

However, we cannot omit discussing them here on account of a scandalous state of affairs which confronts us as soon as we look in on lanthanum in room E. Lanthanum, which ought to occupy this double room with thallium alone, has invited no fewer than fourteen companions to share its quarters, and all these elements are cooped up in one room together! Now this is a situation which is clearly beyond a joke and it cannot be allowed to exist indefinitely. Chemists have made innumerable efforts to find alternative accommodations for these fourteen extra elements in Mendelejeff Court, but they simply refuse to fit into the general system.

These metals of the rare earths, as they are called, have really been a source of serious annoyance ever since the periodic system was introduced. Everywhere else throughout the list we have a uniform change in valence and chemical properties as we proceed from one element to another. Suddenly, and for no apparent reason, we find fifteen elements with beautiful names—lanthanum, cerium, praseodymium, neodymium, illinium, samarium, europium, gadolinium, terbium, dysprosium, holmium, erbium, thulium, ytterbium, and lutecium—which *all* prefer to exhibit a valence of three and which are so similar in their general chemical properties that they almost approach being isotopes. They occur all together in nature, notably in the monazite sands of North Carolina, Brazil, and Travancore. They can be separated one from another only with great storm and strife by taking advantage of slight differences in the solubilities of corresponding salts.

What can we do with these rare earth elements? They cannot be permitted to impose permanently upon the kind nature of lanthanum, since the Ministry of Health will not tolerate such congested conditions for ever. As we have seen, they cannot be accommodated elsewhere in Mendelejeff Court. There is no proper place for them, and they would disturb the harmony of the elements which follow. Chemists in the past have usually done one of two things. Either they have, like the ostrich, shut their eyes to the fourteen superfluous elements and pretended that they were not there, or they have shunted them off into a shed outside of the main building and left them to their own devices.

Now this is manifestly unjust, since, after all, the rare earth metals are unquestionable elements; they have atomic numbers and everything. If, then, we cannot find suitable places for them in Mendelejeff Court, and if we do not want to split our chemical community into two sections, what other alternative can we devise? The only course to take is to move *all* of the elements out of Mendelejeff Court and to discover better quarters for them somewhere else.

As a matter of fact that is exactly what we are going to do. We have been rather loath to throw any bricks at Mendelejeff Court before now, since it has justified its erection by housing most of our chemical community fairly comfortably for a great many years. But apartment houses grow antiquated very quickly, and according to modern standards it must be admitted that Mendelejeff Court is horribly out of date. Apart from the chronic overcrowding in room E of the third floor, there are many inconveniences which its occupants have to suffer. Each is restricted to one floor and often finds that certain of the elements there are quite uncongenial, while certain elements on other floors in whose company life would be more pleasant are too remote for easy association. In other words, many elements exhibit more than one valence, but in Mendelejeff Court they are arbitrarily confined to one especial valence. Thus copper, for example, would be much happier with mercury than with silver and gold, because their compounds are similar, both elements forming two series of salts in which their valences are one and two respectively. Iron is more complex still in its behaviour. It forms one series of salts in which it has a valence of two, and these ferrous salts resemble those of the magnesium family. It forms another series of salts in which it has a valence of three, and these ferric salts resemble those of the aluminium family. Finally, it forms ferrates, such as sodium ferrate Na_2FeO_4, in which it exhibits a valence of six, and these ferrates resemble the sulphates, such as sodium sulphate Na_2SO_4. Yet iron is put up on the roof, far away from its natural companions.

This is bad enough, but the roaming tendencies of the two metals *chromium* (used as an alloy with iron for special purposes, such as "stainless steel" cutlery and "chrome-steel" armour-plate) and *manganese* (also a useful alloy for iron in "manganese steel," employed in making burglar-proof safes) are still more restricted. Chromium has a valence of two in chromous salts, a valence of three in chromic salts and chromites, and a valence of six in chromates and dichromates. Manganese has a valence of two in manganous salts, a valence of three in manganic salts, a valence of four in manganites, a valence of six in manganates, and a valence of seven in permanganates. The periodic system stresses in each case one class of compounds exclusively and ignores the rest. We shall return to discuss this point in a later chapter. At the present moment we shall merely make the mentioning of these metals an excuse to request each of them to rally around it all its companions on the upper floors, so that the exodus from Mendelejeff Court may be complete.

Indeed, it is impossible to put up with conditions there any longer, for the worst feature of all in this typical city apartment house still remains to be exposed. *Children are not allowed!* Poor little hydrogen, our " Nice Baby," cannot enter its portals. This fact alone confirms us in our determination to evacuate and to move our chemical community out of the cruel city into the calm and peaceful open spaces where there will be ample room for everybody and no objection will be raised to Nice Baby.

Chapter XIX

MOVING OUT INTO THE COUNTRY

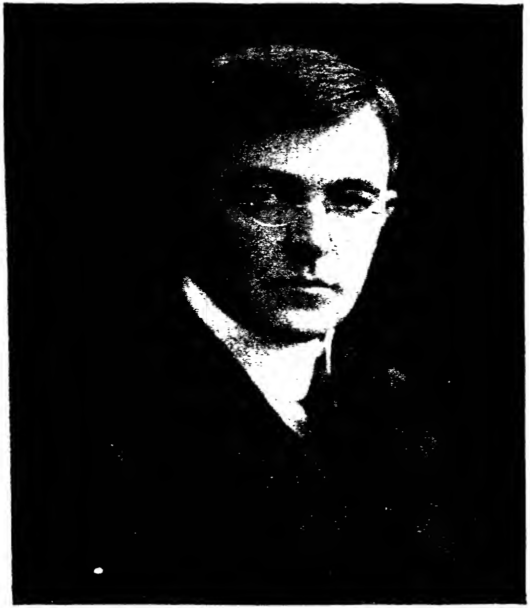
> Nobody is going to make money directly out of the modern theories of the constitution of the atom. Their interest is profound ; they give us a knowledge of the world in which we live, of which our forefathers, of which the men of science when I was a young man, never dreamed, but they are not going to fill anybody's pockets, they are not going to lead to any quotations on the Stock Exchange, and they are not going to be the parents of great dividends.
>
> —The Earl of Balfour, *Messel Memorial Lecture,* 1926.

As a confirmed bachelor, of course, Lord Balfour was not supposed to know anything about babies or about the peculiar habit which they have of growing up. It may be true that in the time of Faraday theoretical progress was not usually accompanied by immediate practical applications—the growth of Faraday's babies was a slow and steady affair. This twentieth century, however, has speeded things up considerably, and many of our recent theoretical advances have originated from the laboratories of great industrial corporations, the directors of which are evidently fully alive to the financial advantages which can be made to accrue therefrom.

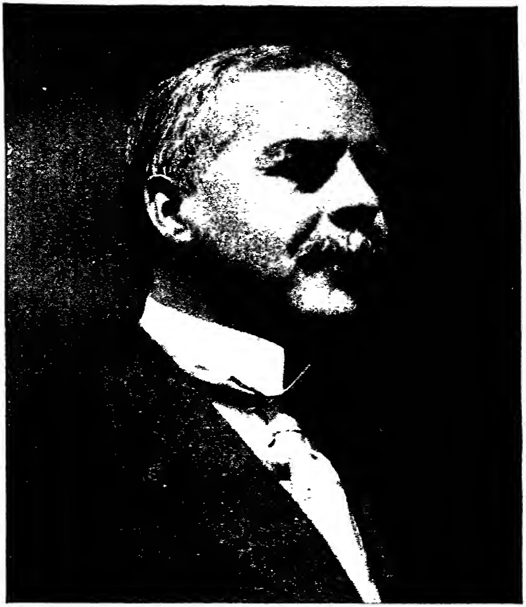
A great part of the work, indeed, which we shall describe in this chapter has issued from the research laboratories of the General Electric Company at Schenectady, where Dr. Irving Langmuir has been working on the atom for the last twenty years. We have already noted the use of the atomic hydrogen torch for welding purposes—a direct result of the theoretical investigations by Dr. Langmuir on the dissociation of the hydrogen molecule. Researches connected with this topic have led to still more significant practical consequences in connection with the development and improvement of electric-light bulbs. Dr. Whitney, the director of the laboratories, has estimated that the work of Langmuir and his colleagues in this field has effected a saving of two hundred thousand pounds a night. This nightly two hundred thousand pounds, of course, has not found its way into the pockets of the research staff of the General Electric Company, neither has it all gone into "great dividends." The principal beneficiaries have been the public at large, including (we hope) the late Lord Balfour.

Let us return now to our atoms. Before we begin to assemble the elements in their new environment, and before we assign to each a definite location in the garden city which we propose to erect, it would be well for us to learn a little more about their structure, so that we may not make any more mistakes in grouping them in an uncongenial way.

So far we have studied mainly the *nucleus* of the atom. However, inasmuch as changes in the nucleus do not occur except in radioactive transformations, whereas in ordinary chemical reactions it is the exterior electrons that are affected, the *arrangement of these exterior electrons* with respect to the nucleus is an important point which we must not omit to consider. Two alternative theories have already been developed in this connection, and a third is just getting under way. While these theories are in many points contradictory, yet each supplies us with certain ideas which are necessary to fill in the complete picture, and we may assume that ultimately some



Irving Langmuir



Gilbert N Lewis

"WHAT NEXT?" ASKED EDISON

On being shown recently some of the latest applications of scientific research in the laboratories of the General Electric Company at Schenectady, even Thomas Edison is said to have exclaimed, "What next?"

Mr. Edison is here shown standing with Dr. Irving Langmuir, twice winner of the Nichols Medal.

single generalization will be devised which will bring all of the three into perfect harmony. In view of its priority of development, we shall take up first what is known as the *octet theory* of Lewis and Langmuir. This theory represents what may be called the " construction period " in the organization of our new colony. Perhaps you have noticed the little cubicles that are erected for the shelter of the labourers engaged in a large building project in an isolated district. Well, the octet theory provides just such temporary shelters for the elements to occupy while their more permanent abodes are being prepared.

The " cubical atom " was first conceived by Professor Gilbert N. Lewis of the University of California in 1902. It was not until much later, however, that a detailed account was first published. To assist you to understand the arrangement of the exterior electrons which this cubical atom postulates, let us recall to your attention Newlands' octaves. Omitting hydrogen and helium, we have eight elements, starting with the alkali metal lithium and ending with the inert gas neon, which are now comprised within the first octave, and the next eight elements, starting with the alkali metal sodium and ending with the inert gas argon, make up a second and precisely similar series.

Since the atom is always electrically neutral, the atomic number of an element represents not only the number of free protons in the nucleus, but also the number of exterior electrons. In the case of hydrogen (atomic number, 1) we have one solitary exterior electron. The helium atom (atomic number, 2) has two exterior electrons, and Lewis supposed these to be situated on opposite sides of the nucleus—a very stable arrangement (see *Fig.* 50). Succeeding electrons are assumed to be built up outside of these two in concentric shells, and throughout the first two octaves the mean positions of the electrons in these shells are considered to be at the corners of imaginary cubes, as shown in *Fig. 53*.

To avoid crowding the diagrams in this picture the nucleus and the inner shell of two exterior electrons, which are common to all of the heavier atoms, have been omitted. Indeed, as soon as we begin to fill in electrons at the corners of our cube, the inner shell of two electrons becomes quite inert so far as ordinary chemical reactions are concerned and can therefore be left out of consideration. Only the electrons in the outermost shell can be added to or lost in interactions with other atoms, and for all of the eight elements from lithium to neon the tendency is either to lose electrons until none are left in this shell, or to gain electrons until a complete ring of eight is formed.

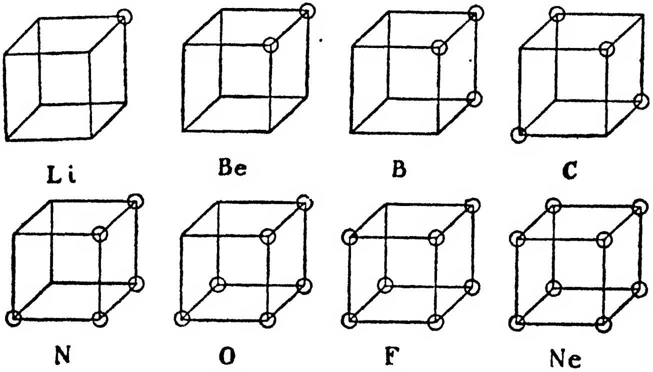


FIG. 53.—Octet Theory—Lithium to Neon.

The valence of any one of these elements is therefore represented either by the number of electrons that can be lost (*positive valence*) or by the number required to form a stable ring of eight (*negative valence*). The next eight elements (sodium to argon) exhibit exactly the same behaviour, a second concentric shell of eight electrons being supposedly built up in a cubical form around the first. The electron arrangements for the succeeding elements in the periodic system, when we leave the short series of Newlands and arrive at the long series of Mendelejeff, become rather less obvious, and we shall therefore postpone the detailed consideration of the positions assumed by the exterior electrons in heavier atoms until later.

One point can be noted immediately. However complex the atom may be, its outermost shell *never* (we really mean never—not " hardly ever " !) contains more than eight electrons. All of the electrons within this outermost shell are dormant, so that we never have valences exhibited by any element throughout the system greater than eight. Indeed, if you refer back to the picture of Mendelejeff Court, you will notice that when elements exhibit both positive and negative valence the arithmetical sum of these is always eight. When the maximum positive valence is exercised, the outermost shell is stripped of electrons ; when the maximum negative valence is exercised, this shell is made complete. The inert gases already have their outermost shells filled up and therefore do not take part in chemical reactions at all.

An element which necessarily deviates from this rule of eight is hydrogen. Hydrogen usually reacts as a positive element by losing its lone electron, leaving a positively charged hydrogen ion H^+ ; but it is also possible for hydrogen to react as a *negative* element, acquiring one electron to fill up its outer shell to the helium maximum of two. This can happen, clearly, only when Nice Baby is fortunate enough to meet another atom which is particularly eager to donate an electron to it. Such a requirement is filled by the alkali metals, and we find in fact that hydrogen combines with these to form *hydrides*, like lithium hydride LiH, in which the hydrogen exists as a negative hydride ion H^-. Notice how careful we have been in the past always to refer to hydrogen as "it." Usually Nice Baby prefers to pose as a potential metal, but, as we have just seen, under special circumstances the element actually behaves as if it were one of the blondes preferred by gentlemen.

Other elements exercising positive valence frequently choose to keep two of their outermost electrons rather than to strip their shell completely. They thereby bring their atoms to an imitation of the helium structure, rather than to a copy of the inert gas nearest to them. Thus we have seen that tin and lead, whose atoms have four electrons in their outermost shell, may form with chlorine the compounds $SnCl_2$ and $PbCl_2$ as well as the compounds $SnCl_4$ and $PbCl_4$. This behaviour is, indeed, very common among the heavier elements.

When combination takes place between two atoms (for example, Li and F) by loss and gain of electrons, we are left with a compound, such as lithium fluoride LiF (*Fig. 54*), in which the constituent atoms are apparently separate, being held together only by the attraction of their opposite charges. An atom may also complete its stable ring of eight, however, by sharing electrons with another atom, as in the case of the fluorine molecule F_2 (*Fig. 55*). The common habit of the atoms of gaseous elements to associate in molecular groups like O_2 and Cl_2

is thus explained. Similarly in carbon disulphide CS_2, we may assume that all three atoms have completed their "octets," each sulphur atom holding two pairs of electrons in common with the central carbon atom (*Fig. 56*).

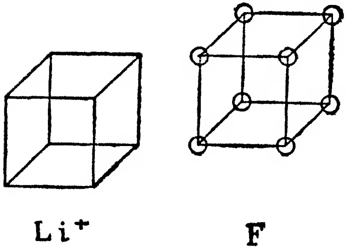


FIG. 54.—Typical Polar Combination.

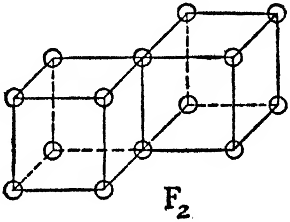


FIG. 55.—Typical Non-Polar Combination.

We have already discussed this sharing of pairs of electrons between atoms in a previous chapter under the name of *co-valence.*

Compounds like lithium fluoride or sodium chloride, in which an electron has already passed from one atom to another, are evidently potentially ionized, and if we can diminish the attractive forces between the two charged atoms sufficiently to enable them to break away from one another, we obtain immediately the free ions (as in aqueous solution). Such compounds are consequently *polar*. In substances like carbon disulphide, on the other hand, where electrons are held in common, the molecule

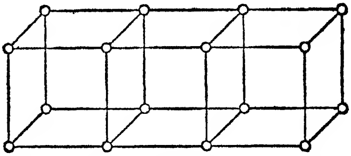


FIG. 56.—Carbon Disulphide Molecule (Octet Theory).

will not tend to break up in this way. Such substances do not, therefore, ionize in solution, and are *non-polar*.

In cases where combination between two elements produces a compound, such as sodium chloride, in which the arrangement of electrons is much more stable than in the original substances, we shall clearly obtain a considerable diminution in the internal energy of the system as the result of the interaction. Chemical activity or affinity hence appears to depend, finally, upon atomic structure. Where the electron arrangement is extremely stable, as with the helium family, the element will be inert. Where electrons are readily gained or lost, as with the halogens or the alkali metals, the element will be active, and will form very stable compounds with such elements as assist most readily in the interchange.

We may now proceed to dicuss the later elements in the periodic system and the correlation of their electron arrangements with valence. For this purpose we shall return for a moment for a last look at the deserted corridors of Mendelejeff Court, in order to draw some information from the plan on which its apartments are arranged that will be of assistance to us in grouping the elements amicably in their new settlement.

When we examine the organization of the rooms in Mendelejeff Court, we see that the range of applicability of Newlands' law of octaves was extended by lengthening the third and fourth series from *seven* to *seventeen* (or, including the inert gases subsequently discovered, from *eight* to *eighteen*) elements. But still there remained a point of difficulty at either end of the system. Hydrogen was left out in the cold altogether. The elements of the rare earths made the fifth series unduly congested, and we either had to throw our whole system out of harmony or omit them too.

Suppose now we attempt to arrange the elements in a different way, not trying to force them into series of fixed lengths, but simply writing them down in successive columns, taking them in the order of their atomic numbers,

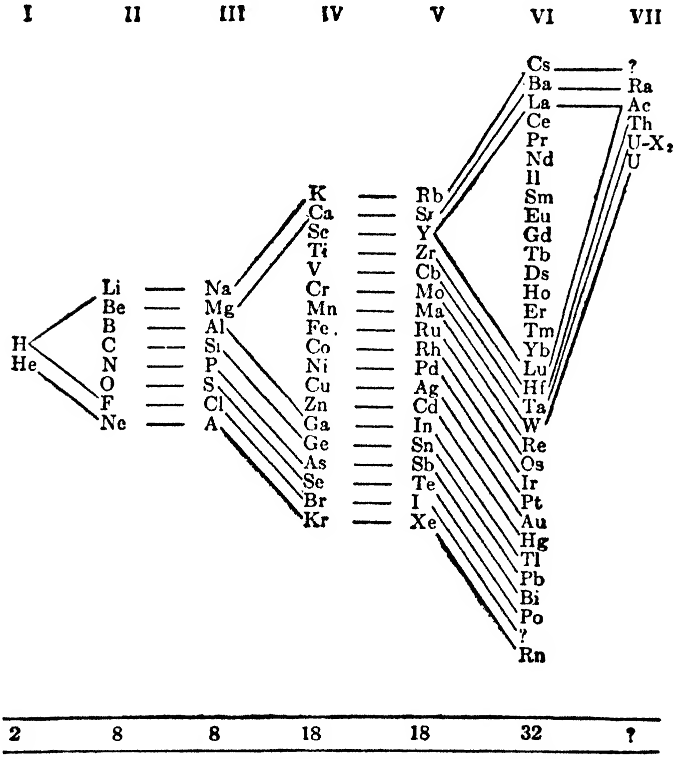


FIG. 57.—Plan of Bohrville (see next chapter).

leaving blanks for elements still to be discovered and concluding each column with an inert gas. We obtain in this way the plan shown in *Fig. 57*. Let us count up the number of elements in each column. The first column contains 2; the second 8; the third also 8; the fourth 18; the fifth again 18; the sixth 32; the seventh is obviously incomplete.

Now we can express these numbers in a regular mathematical series:

$$2 \times 1^2;\ 2 \times 2^2;\ 2 \times 2^2;\ 2 \times 3^2;\ 2 \times 3^2;\ 2 \times 4^2.$$

This series is far too stiking to be due to chance. Evidently, then, *the inert gases occur at regular sequences in the list of elements*. Their atomic numbers are expressed by the ascending series:

$$2(1^2 + 2^2 + 2^2 + 3^2 + 3^2 + 4^2 + \ldots)$$

How about the other elements? Can we connect them up at all? Let us try to do so by drawing lines between those exhibiting similar properties. If an element resembles two others, let us draw lines to them both. But do not let us draw so many lines as to confuse our diagram, because we shall need to refer to it frequently in the course of the next chapter.

Giving *Fig. 57* merely a casual examination at this point, we see that it fits the facts a great deal better than the Mendelejeff arrangement. There must be some fundamental principle underlying it. What that principle is we are now in a position to elucidate.

Chapter XX

BOHRVILLE

> For the world was built in order
> And the atoms march in tune.
>
> —Ralph Waldo Emerson.

As an atom is to a man in size, so is a man to the sun. It is therefore an interesting coincidence that man, in seeking to discover the structure of that almost infinitesimally small unit, the atom, should have developed a precisely similar model for its component parts as was previously arrived at for that almost infinitely great unit, the sun with its attendant planets.

The first theory of atomic structure which was discussed in the preceding chapter was essentially a *chemical* theory. It explains very clearly a number of points in connection with the formation of compounds between atoms, but the picture which it draws of the atom itself is very patchy and incomplete. Physicists have always raised objections to what they called the " loafing electron." A system of electrons sitting around at the corners of cubes turns out not to be a stable system when it is subjected to mathematical analysis. Of course, chemists never really believed that the electrons *did* sit still at the corners of the cubes, any more than they believed that the cubes themselves had an actual material existence. The corners of the cubes merely represented the mean, or average, positions occupied by the electrons. Even this concession, however, did not satisfy the physicists, so the latter have gone ahead and invented a second theory of atomic structure which, it must be admitted, is vastly superior to the first in most respects, even although it restricts itself primarily to single atoms and does not attempt to explain all types of interactions between atoms.

According to this second theory of atomic structure, developed mainly by a talented Danish physicist, Niels Bohr, the arrangement of the exterior electrons is not on a static " cubical " basis. The electrons are assumed to be in rapid motion around the nucleus, each travelling in a regular orbit, just as the planets in our solar system journey in orbits around the central sun. The orbits described become more and more complex the greater the number of electrons, and we have not space here to go into all the details of the Bohr theory, fascinating and comprehensive though they are. It will suffice at this point to state that the orbits may be divided into classes, which we shall call (for reasons that will appear shortly) " orbits of principal quantum number 1," " orbits of principal quantum number 2," and so on. These successive classes correspond, in a general way, with the electron shells of the octet theory, although the ratios of the major axis to the minor axis in the orbits of different electrons of the same principal quantum number are not all identical. This fact is illustrated in *Fig. 58*, which shows the various types of orbits within each class. Each orbit may be distinguished by assigning to it, besides a principal quantum number, a subordinate quantum number. Thus we may speak of a 4_4 orbit (the largest circle in *Fig. 58*), a 2_1 orbit (the smallest ellipse in the same figure), and so on. Representations of the hydrogen atom (*Fig. 59*), the helium atom (*Fig. 60*), the lithium

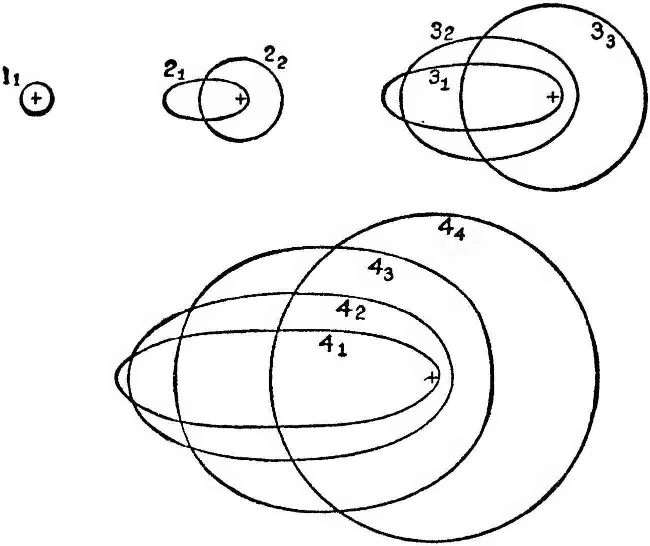


FIG. 58.—Electron Orbits according to Bohr Theory.

atom (*Fig. 61*), the neon atom (*Fig. 62*), and the sodium atom (*Fig. 63*), according to the Bohr theory, are given

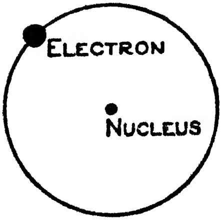


FIG. 59.—Hydrogen (Atomic No. 1).

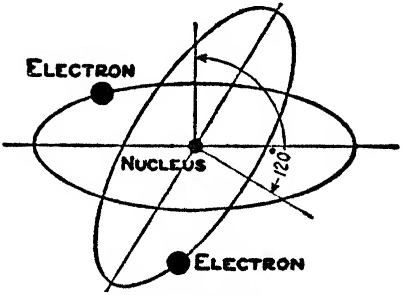


FIG. 60.—Helium (Atomic No. 2).

below, and will be found to repay careful analysis in connection with the later discussion.

Two minor points may be noted here. First of all, as will be seen in *Fig. 61*, the orbits of all outer electrons in elements which follow helium are *precessing* ellipses (that is, the orbit rotates in space), owing to the perturbing effect of the two inner electrons. Secondly, the compactness of the neon atom as contrasted with that of sodium suggests immediately the reason for the inert character of the former element, and for the ease with which the latter rids itself of its widely roaming outermost electron to become a positively charged sodium ion. The same contrast holds, of course, for helium and lithium, and

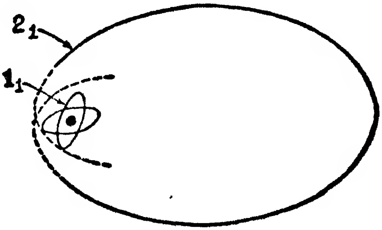


FIG. 61.—Lithium (Atomic No. 3).

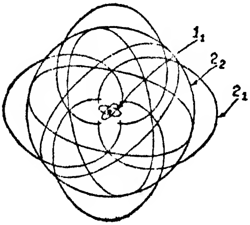


FIG. 62.—Neon (Atomic No. 10).

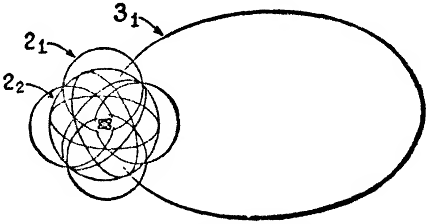


FIG. 63.—Sodium (Atomic No. 11).

indeed for the family of the inert gases and the alkali metal family in general.

These atomic pictures must not, of course, be taken too literally. They are merely attempts to represent the main essential facts, and minor points such as scale are sacrificed for the sake of clearness. Thus if the hydrogen atom were correctly represented, the nucleus would be invisible. Imagine a hydrogen atom magnified until it became as big as the Wembley Stadium. The nucleus would then be equivalent in size to a golf ball in the centre, while the electron might be represented by a man in an aeroplane circling over the top ring of seats. The atom is therefore mainly emptiness after all.

Here we come to a ticklish point in the Bohr theory. A negative electron revolving around a positively charged nucleus should, according to all the experts, radiate energy. Classical dynamics absolutely demands that this should be the case, but if our atoms really did radiate out energy in this way, the exterior electrons would fall slowly but surely into smaller orbits and finally flop into the nucleus. We know that this does not happen, since the atom is a stable affair and does not collapse with time. Consequently Bohr calmly decided to disregard classical dynamics, so far as an electron keeping in a single orbit is concerned.

But now watch what follows. Under suitable persuasion an electron can be urged to jump from its original orbit to another orbit of higher or lower quantum number, and classical dynamics demands that when such a jump takes place a definite amount of energy shall be used up or evolved. This definite amount of energy is called a *quantum*, and we must digress for a moment to consider this quantum. The German physicist Planck put forward in 1900 the hypothesis that, just as matter was not continuous but made up of definite minute particles or atoms, so energy also was not continuous but existed only in definite indivisible units or quanta. Of course, these units must be very small, but their magnitude can be calculated by methods which we shall not attempt to describe, and the application of Planck's quantum theory to modern physics has been just as fruitful in its results as was the application of Dalton's atomic theory to chemistry a century ago. Planck has been very aptly styled, indeed, the Dalton of the twentieth century.

Now it is precisely because energy can be absorbed or emitted only in quanta, or jerks, that the electrons are unable to change their orbits gradually, but are restricted to definite jumps from one of the orbits shown in *Fig. 58* to another. Bohr calculated, *using classical dynamics*, what wave-lengths the energy emitted or absorbed in such orbital changes for atoms of different elements would possess, and found that the calculated values were *in exact agreement* with the lines on the spectrum (see, for example, *Fig. 42* for helium in Chapter XIV) in each case. This furnishes the main experimental evidence in favour of Bohr's theory and constitutes the chief basis for the approval which it has enjoyed. Why classical dynamics is applicable to it in one place but not in another, however, is still a mystery.

Some rules regarding the arrangement of the electrons in their orbits may now be developed. The *maximum* number of electrons in each quantum class is limited as follows:

First class; orbits of quantum number 1: 2 electrons.
Second class; orbits of quantum number 2: 8 electrons.
Third class; orbits of quantum number 3: 18 electrons.
Fourth class; orbits of quantum number 4: 32 electrons.

Note that these numbers follow the mathematical series: 2×1^2; 2×2^2; 2×3^2; 2×4^2. The first rule, then, in the Bohr theory is that the maximum number of electrons of quantum number n is given by the expression $2n^2$. To this first rule we must immediately add a second, which reads as follows: The outermost class of orbits cannot, however, contain more than 8 electrons. The next to the outermost, similarly, cannot contain more than 18.

These two rules restrict the electron arrangements to the following periods:

Number of Period.	Last Element in Period.	Atomic No. of Last Element.	Number of Electrons in Successive Classes. I	II	III	IV	V	VI
1	He	2	2	..	..	..	..	..
2	Ne	10	2	8	..	..	..	..
3	A	18	2	8	8	..	..	..
4	Kr	36	2	8	18	8	..	..
5	Xe	54	2	8	18	18	8	..
6	Rn	86	2	8	18	32	18	8

If you refer back to *Fig. 57* at the end of the last chapter, you will begin to get a general plan of the new garden city which has now been erected as a country retreat for our elements, replacing Mendelejeff Court. This garden city evidently needs a name, and in honour of the man who has done most in promoting the villa colony we shall venture to call it Bohrville.

Bohrville, you will notice, occupies a plot of land which is in shape an irregular triangle. Each element owns a villa on a particular avenue, and communicates with its friends on other avenues by means of side streets connecting. First Avenue contains only two elements, hydrogen and helium. Second Avenue and Third Avenue each accommodates eight elements. Fourth Avenue and Fifth Avenue are longer and each has room for eighteen, a simple explanation being thus provided for the first time for the introduction of Mendelejeff's long periods into the periodic system. Furthermore, these two avenues with eighteen elements are followed by Sixth Avenue, which has space for no fewer than thirty-two. Sixth Avenue, therefore, includes *perfectly* (since $32 = 18 + 14$, as even *Pooh* could tell us by this time) the fourteen rare earth elements for which no provision could be made in Mendelejeff Court. Seventh Avenue, finally, has only six occupants, but that is merely because the elements after uranium are too unstable to exist.

So far as positive and negative valence relationships are concerned, the Bohr theory deals with the elements living on the first three avenues precisely as was done under thetoctet theory. Each element seeks either to rid itself of is outermost class of electrons entirely or to complete the quota of its class to the maximum. Co-valence, it must be admitted, has not so far been satisfactorily explained. Probably the pairs of electrons shared between atoms assume a new type of intermediate orbit, but the Bohr theory has not so far been extended to cover molecular structure.

When we make an inspection of Fourth Avenue, however, some compensation appears. For the elements in this and succeeding avenues a new rule has been formulated which runs as follows: During the change of an inner class of orbits from a stable group of eight to one of eighteen, or from eighteen to thirty-two, there occurs a transition series of elements which can have more than one structure. This rule accounts for the *variable valence* which we find in elements such as chromium, manganese, iron, copper, etc., as we shall now see.

Let us start from argon, the last element of the third period. Its electron arrangement is 2,8,8. The next element, potassium, is a typical alkali metal; its electron arrangement must be 2,8,8,1. Calcium, similarly, gives 2,8,8,2. But as we proceed farther we find a more and more pronounced tendency for additional electrons to exercise the choice which is available to them either to continue filling up the fourth class of orbits to its maximum of eight or to slip into the third class of orbits and start to fill that up toward its maximum of eighteen. Thus we arrive at alternative possible electron arrangements for the same element. Some illustrative examples are given below.

Chromium: 2,8,8,6 (chromates); 2,8,11,3 (chromic salts); 2,8,12,2 (chromous salts).
Manganese: 2,8,8,7 (permanganates); 2,8,9,6 (manganates); 2,8,11,4 (manganites); 2,8,12,3 (manganic salts); 2,8,13,2 (manganous salts).
Iron: 2,8,10,6 (ferrates); 2,8,13,3 (ferric salts); 2,8,14,2 (ferrous salts).
Copper: 2,8,17,2 (cupric salts); 2,8,18,1 (cuprous salts).

This rule clarifies a great many chemical relationships. Chromates, manganates, and ferrates are derived from atoms with shells of 2,8,8,6; 2,8,9,6; and 2,8,10,6 electrons respectively. They all have the same outer number of electrons, and we meet with a similar arrangement in sulphur (2,8,6). Hence chromates, manganates, and ferrates show close resemblances with each other, and also with sulphates. Note that it is always the number of electrons in the outermost set of orbits that is of primary importance. These are the *valence electrons.*

Element.	Atomic Number.	Electron Arrangement.						
Lanthanum	57	2	8	18	18	8	3	
Cerium	58	2	8	18	18	8	4	(ceric)
		2	8	18	19	8	3	(cerous)
Praseodymium	59	2	8	18	20	8	3	
Neodymium	60	2	8	18	21	8	3	
Illinium	61	2	8	18	22	8	3	
Samarium	62	2	8	18	23	8	3	
Europium	63	2	8	18	24	8	3	
Gadolinium	64	2	8	18	25	8	3	
Terbium	65	2	8	18	26	8	3	
Dysprosium	66	2	8	18	27	8	3	
Holmium	67	2	8	18	28	8	3	
Erbium	68	2	8	18	29	8	3	
Thulium	69	2	8	18	30	8	3	
Ytterbium	70	2	8	18	31	8	3	
Lutecium	71	2	8	18	32	8	3	

The same method of analysis may be applied to elements in the higher avenues. The table given above indicates how beautifully the rare earth elements fit into the new periodic system by the filling-up of the fourth shell from 18 to 32 electrons, and how the predominance of trivalency and the extreme similarity of chemical properties throughout the whole series are thus completely explained.

It is of importance, also, to mention that the elements which possess incomplete inner shells with labile electrons are precisely those same elements which give coloured ions. Rare earth elements such as praseodymium and neodymium, as well as chromium, manganese, and copper, are all in this class. The capacity of their compounds to absorb light energy with wave-lengths in the range of the visible spectrum is clearly connected with their ability to allow an electron to slide easily from one orbit to another.

The protons and electrons inside the nucleus of an atom are also, presumably, in rapid and ordered motion, and recent work by Rutherford is beginning to give us a glimpse into this very heart of things. When the nucleus is very complex it may happen occasionally that the congested units therein existent will get themselves into such a terrible tangle, in the course of their gyrations, that the cementing electrons no longer cement and part of the nucleus is ejected. Here we have the most plausible explanation of radioactivity. How frequently such a catastrophe will occur depends, of course, upon the nature of the atom. Its average life may be short or long, as we saw in the uranium disintegration series.

Chemistry " within the atom " is still in its infancy, but it cannot be doubted that its development will lead to results of the greatest importance in the near future. The facts presented in this volume have almost all been derived on a purely experimental basis, and the construction of hypotheses to correlate and explain these facts has been a long and painful process. On the basis of atomic structure, however, the next generation of chemists will be able to calculate in advance, in greater and greater detail, the necessary properties and reactions of all forms of matter. A multitude of new facts will thus be brought to light, and many new applications of chemistry to industry will become evident.

Since we began this chapter by talking about stars, let us conclude by discussing a very interesting astronomical fact. Sirius, the Dog Star, which is one of the two brightest stars in the sky, appears under the telescope to be double. This is a very frequent occurrence, even in a prohibition era, but there is something very unusual about the tiny faint companion of Sirius. Its density is many thousand times that of water, and far beyond the density of any element which is known on the earth.

How can this be explained? We have already seen that an electron may be driven from one orbit to another in the same atom by the absorption of a certain amount of energy. If we can get an electron very excited (for example, by bringing chlorine near to an element like sodium, or by bombarding a less impressionable element with alpha-rays), we may drive it away from its parent atom altogether. Now if we could repeat this process for *all* the exterior electrons, we should be left with a very minute residue—the nucleus—which would still contain practically the whole mass of the atom. The density of such nuclei tightly packed together would obviously be almost inconceivably great in comparison with our ordinary terrestrial materials. A cubic inch of the companion of Sirius, indeed, would weigh almost one ton, and similar stars are known which have a density eight times greater than this!

Chapter XXI

WHAT OF THE FUTURE?

> Beauty is truth, truth beauty,—that is all
> Ye know on earth, and all ye need to know.
> —John Keats, *Ode on a Grecian Urn.*

The beauty of the Bohr theory of the atom is so evident that a bitter feeling of disappointment is arising in many quarters as the conviction is gradually forced upon us that, after all, its foundations are not firmly fixed. Our chagrin is comparable, indeed, to that of the small child who has just been told that there is no Santa Claus. We are not referring here to minor modifications in the original idea of Bohr, such as have been made by Main-Smith and Stoner. While these modifications are of considerable interest, we have not the space to discuss them in this volume. For the sake of completeness, however, it is necessary for us to indicate as far as possible the main features of a third theory of atomic structure, which threatens to supersede the Bohr concept.

We likened the octet theory to the construction period in the organization of our garden city. The Bohr theory, in the same way, represents what may be called the "settling-down period." Whether the third period upon which we are now entering will merely be a "cleaning-up period," which will leave our ninety elements still in occupation of a better and brighter Bohrville, or whether it will involve their migration to an entirely different environment is a question which remains, as yet, unanswered.

With all these new theories coming in and going out, the impression has become prevalent among the public at large that modern science is advancing *too* rapidly. Recently, in fact, the Bishop of Ripon made the suggestion that it would be well if our scientific leaders would take a ten-year vacation in order to give ordinary mortals an opportunity to catch up. The field is so fascinating, however, that it is highly improbable that his advice would be followed even if workers in pure science were so well paid as to be able to afford such an extended holiday.

When we look into the matter, the current situation is not without parallel. Robert Boyle was solicited more than two hundred years ago to withdraw from the Royal Society because of the evil effect which the work of its members was having upon the general community. A certain Dr. Henry Stubbe wrote to him as follows:

> I beseech you, sir, consider the mischief it [the Royal Society] hath occasioned in this once flourishing kingdom, and if you have any sense, not only of the glory and religion, but even of the being of your native country, abandon that constitution. It is too much that you contribute to its advancement and repute: the only reparation you can make for that fatal error is to desert it betimes. Do not you apprehend that all the inconveniences that have befallen the land, all the debauchery of the gentry, will be charged on your account? . . . It will be impossible for you to preserve your esteem but by a seasonable relinquishing of these impertinents.

The present-day protest of the worthy bishop seems quite tame in comparison. Of course, there are still divines who talk in the above strain regarding scientific progress (and especially evolution) nowadays, but we shall not give them the publicity for which they pine by mentioning any names.

Our third theory of atomic structure is called the Schrödinger wave theory, after a brilliant young Austrian physicist who has been prominent in its development. This theory, as we shall see when we come to it, is not at all easy for the layman to understand, yet popular presentations are beginning to appear in the well-known magazines. In connection with one such article a great American daily recently saw " fit to print " the following comment:

> The next time you go shopping for atoms insist upon the Schrödinger sort. Don't let the dealer fool you with any last year's Bohr stuff. And after this keep a sharp eye out for changes in fashion.

Now this kind of statement may be very clever, but it is likely to do a great deal of harm. The man in the street who reads it is certain to get the idea that our scientific laws and theories are just wild guesses which may have to be discarded at any moment, and consequently will immediately make up his mind that he is not going to waste any time trying to understand them. Even persons who are highly educated along non-scientific lines often betray the fact that they have acquired an entirely wrong idea of the nature of scientific laws and theories. They admit that this is a scientific age, but they openly bewail the impermanence of scientific truths as contrasted with the eternal beauties of literature and art. They are especially pained by the fact that the scientist of to-day looks down with scorn upon the " laws of nature " reverenced by his forefathers, whereas they themselves do not dare to dream that they can ever aspire to the heights attained by Shakespeare or by Michelangelo. Let us stop to analyse this lament for a short time. We can return to Schrödinger with a lighter heart after we have justified our general method of procedure.

In the first place, there is no such thing as a law of nature known to us at present. All of our laws in chemistry, or in any other science, are " man-formulated " laws. They represent the general result of all the observations which man has made in a given field, and record in the fewest possible words man's interpretation of what he has found to happen. In a previous chapter we have discussed how a hypothesis grows up into a theory, and how a theory grows up into a law ; but however firmly established a law may seem to be, it is always *man-made* and not nature-made. It expresses what we have ascertained so far to be true, but we have no means of enforcing obedience to such a law when (as invariably happens sooner or later) exceptions are unearthed. Thus the law of definite proportions, which remained " unbroken " for over a century, was not a law of nature ; it was a law of Joseph Proust. And when T. W. Richards discovered in 1914 that lead chloride derived from radioactive sources did not " obey " this law, he did not propose any penalty for the transgression ; he did not attempt to persuade the compound to mend its conduct ; he rather sought to amend the law. Science, then, is always looking ahead and advancing nearer and nearer toward final truths, but never expects to reach the point where farther progress is impossible.

This constitutes, indeed, the real joy of science. Every student in a chemistry class to-day can (if he wishes) know more about the subject than Boyle, Lavoisier, or

Faraday ever imagined possible. The point of view is forward instead of backward. Furthermore, although our opinions change, they do not change in quanta or jerks; they change logically and progressively. If Schrödinger can see farther than Bohr, it is merely because he is standing on Bohr's shoulders, just as Bohr was standing on the shoulders of Mendelejeff. Let us substantiate our own statements by a few words from one who has done much to reconcile science with humanity. Professor R. A. Millikan of the California Institute of Technology, the Nobel Prize winner in physics in 1923, has recently observed, in connection with the work of Einstein on the laws of gravitation, regarding which you have probably read a lot and understood very little, as follows:

> One of the greatest contributions of science to life is the discovery that progress is in general made by the evolutionary process. *There are no revolutions in science.* Einstein does not replace Newton, he merely supplements him. In so far as Newtonian mechanics was a body of experimental facts, it is eternally true. The whole of Newton is incorporated in Einstein. Let the revolutionary reformer ponder well that fact.

This same situation holds, it is clear, with regard to chemical laws like the law of definite proportions. New discoveries do not lead to the repeal of the law as unconstitutional; they result instead in its restatement in an expanded form. In the same way with the law of evolution in biology. Fundamentalists are fond of deriding this law because Darwin's original statement of it has been modified. Of course, *they* never modify any of their views, and are quite incapable of recognizing the fact that the law of evolution proves itself by itself evolving.

Having eased our mind fully on all these points, let us at last return to Schrödinger's wave theory. The main idea behind this theory is, as you will see, rather nebulous, but we shall describe to you its evolution as intelligibly as is possible at the present time.

The trouble started about three years ago when a French physicist, Louis de Broglie, attempted to apply Einstein to the Bohr type of electron. De Broglie found that, in order to get his equations to fit, it was necessary to assume that an electron revolving in its orbit around the nucleus was attended by a group of waves. We do not need to worry ourselves as to what the waves consist of; as a matter of fact, we do not know yet what light-waves consist of—they are merely, as Lord Kelvin said, " the nominative of the verb to undulate." These waves are born a little behind the electron and die out a little way ahead. But the electron moves so fast, and its orbit is so small, that the possibility had to be considered that the wave series might persist all around the orbit. If this should be the case, then the orbit would have to be such that the waves would not annihilate one another; crest must meet crest, and trough meet trough. When the necessary calculations were made with this point in mind, de Broglie found that the only possible orbits meeting such conditions were the orbits already postulated by Bohr.

This was all very promising and seemed so far to strengthen the Bohr theory, but now Schrödinger entered the field and said, " If the waves work so well, why a material electron at all ? " Diving deep down into the waves with all the latest mathematical equipment, Schrödinger has finally emerged with some involved equations in his grasp which show that the corpuscular electron is, indeed, unnecessary. In place of it we have the nucleus surrounded by a vibrating halo, an undulating aura, a pulsating nebula, or whatever else you like to call it. This is electrical in its nature, and the density of its charge fluctuates at different points, but the halo is brightest, the aura is most pronounced, or the nebula is densest around the path of the little hard electron that, as such, isn't there any more. Theoretically, a single atom extends out into infinite space; practically, of course, its intensity is negligible beyond a very minute radius. The nucleus degenerates into an electrical hole in the centre of the aura.

Of course the proof of the pudding is in the eating. The Schrödinger atom explains everything which the Bohr atom explains, and it also accounts for a great many facts with which the Bohr theory has not been brought into agreement. Unlike the Bohr atom, it does not conflict with classical dynamics, but it still has certain troubles of its own connected with its hypothetical waves. Perhaps when we have discovered what these are actually saying our ideas will become a little more tangible. Of course, there is no guarantee that the waves will be permanent, although we hope, for the bishop's sake, that no new theory will appear to complicate the situation until the Schrödinger atom and the Bohr atom have fought matters out to a finish. At the present moment, in spite of the hard blows which it has recently received, the Bohr theory is by no means at its last gasp.

There exists, indeed, with respect to the electron to-day the same situation which has existed on and off with respect to light for more than two centuries. We have two theories concerning the nature of light, the corpuscular theory of Newton and the wave theory of Huyghens.

These theories are absolutely contradictory; each is perfectly satisfactory for half of the facts and quite inapplicable to the other half. As Sir Oliver Lodge has said, the two concepts are like a shark and a tiger—each supreme in its own domain but helpless in that of the other. The physicists, accordingly, make a virtue of necessity and use both without hesitation. Sir William Bragg, the world's greatest authority on crystal structure, has summarized the matter thus:

> There must be some fact of which we are entirely ignorant, and whose discovery will revolutionize our views of the relation between waves and ether and matter. For the present we have to work both theories. On Mondays, Wednesdays, and Fridays we use the wave theory; on Tuesdays, Thursdays, and Saturdays we think in streams of flying energy, quanta or corpuscles. That is, after all, a very proper attitude to take. We cannot state the whole truth since we have only partial statements, each covering a portion of the field. When we want to work in any one portion of the field or other, we must take out the right map. Some day we shall piece all the maps together.

Although it is rash to make predictions, yet we may expect that chemists in the same way will continue for a long time to develop both the Bohr atom and the Schrödinger atom, mutually contradictory though they be, until the time arrives when some genius will succeed in correlating them both into some broader generalization.

When all is said and done, very few theories in science ever go completely out of existence. They merely become embodied as part of some subsequent theory. The only noteworthy exception in chemistry is the phlogiston theory. Even here there are resourceful people who argue that, after all, when a metal is converted into rust by heating, it does lose something—it loses *electrons*. If we are willing to accept this quibble, then the electron becomes the modern equivalent of the old-time phlogiston. The electron has also been suggested as a suitable candidate for admission to the list of elements, being the element with atomic number zero. Some philosophers even get quite agitated over speculations regarding the missing element with atomic number $\sqrt{-1}$! It seems a wicked waste of time, but perhaps it keeps them out of worse mischief.

There is little likelihood that any of our present-day theories will prove to be merely an encumbrance to the advancement of science, as the phlogiston theory was for over a century. For every person who used to philosophize over phlogiston, there are now a hundred experimenting on the atom, and any false notions are disproved and abandoned almost as soon as they are announced. It is a shame, nevertheless, that it is too late to impose upon George Ernst Stahl, the protagonist of the phlogiston theory, any fitting punishment—something lingering, with boiling oil or molten lead in it !

Before we close, we must not forget our one million pounds. To win this, it is still necessary that we frame an acceptable definition of the term atom. If the Bohr theory had been left undisturbed, this would not have been such a difficult thing to do. Under the Schrödinger theory, however, the nearest we can come to defining the atom is to call it " an aura surrounding a cavity." We doubt very much that this attempt will be considered worthy of the million pounds, but we hereby promise that, if the money ever does come into our hands, each purchaser of this volume who will send us a sworn affidavit that he has followed our footsteps faithfully from beginning to end will receive the sum of one hundred pounds (free of Income Tax) therefrom as a recompense for his pluck and perseverance.

We have now guided your faltering feet to the summit of the first hills leading upward toward the high mountain-peaks of present-day chemical discoveries, and there we shall have to leave you for the time being. We hope, however, that you will always retain pleasant memories of your brief excursion into chemistry, and that you have appreciated the beauties of the subject sufficiently to understand the words of the old alchemist:

> The chymists are a strange class of mortals impelled by an almost insane impulse to seek their pleasure among smoke and vapour, soot and flame, poisons and poverty, yet among all these evils I seem to live so sweetly, that may I die if I would change places with the Persian King.

As Robert Louis Stevenson might say, if he had to rewrite his Law of Felicity in modern free verse:

> The world is so full of protons and electrons,
> I'm sure we should all be as happy as chemists—

a quotation which brings us right back to where we started and with which, therefore, we may very properly close.

APPENDIX

I. THE METRIC SYSTEM

THE metric system is a decimal system of measures and weights, of which the metre is the fundamental unit: first adopted in France, but now generally used, especially for scientific purposes. The names of the various units of this system are formed by combining the words *metre*, *litre*, *gramme*, etc., with the following numerical prefixes:

milli-	= one-thousandth	or	·001
centi-	= one-hundredth	or	·01
deci-	= one-tenth	or	·1
deca-	= ten	or	10·
hecto-	= one hundred	or	100·
kilo-	= one thousand	or	1000·

Length. 1 metre (1 m.) = 10 decimetres = 100 centimetres (100 cm.) = 1000 millimetres (1000 mm.).

1 kilometre = 1000 metres (1000 m.) = 0·6214 mile.

1 decimetre = 0·1 m. = 10 centimetres = 3·937 inches.

1 metre = 1·094 yards = 3·286 ft. = 39·37 in.

Volume. 1 litre = 1000 millilitres (1000 ml.).

1 millilitre = volume of 1 gramme of water at 4° Centigrade.

1 litre = 0·03532 cu. ft. = 61·03 cu. in. = 1·136 quarts (Brit.) or 1·057 quarts (U.S.) = 35·3 oz. (Brit.) = 34·1 fl. oz. (U.S.).

1 ounce (Brit.) = 28·4 ml. 1 cu. ft. = 28·32 litres.

Weight. 1 gramme (g.) = wt. of 1 millilitre of water at 4° Centigrade. 1 kilogramme (kg.) = 1000 g.

1 gramme = 10 decigrammes = 100 centigrammes = 1000 milligrammes (1000 mg.).

1 kilogramme = 2·205 lb. avoirdupois.

1000 kilogrammes = 2205 lb. = 1 metric ton.

1 lb. avoird. = 453·6 g.

1 oz. avoird. = 28·25 g. 100 g. = 3·5 oz.

II. INTERNATIONAL TABLE OF ATOMIC WEIGHTS

	Symbol	Atomic Number	Atomic Weight
Actinium	Ac	89	226·
Aluminium	Al	13	26·97
Antimony	Sb	51	121·77
Argon	A	18	39·91
Arsenic	As	33	74·96
Barium	Ba	56	137·37
Beryllium	Be	4	9·02
Bismuth	Bi	83	209·00
Boron	B	5	10·82
Bromine	Br	35	79·916
Cadmium	Cd	48	112·41
Calcium	Ca	20	40·07
Carbon	C	6	12·000
Cerium	Ce	58	140·25
Caesium	Cs	55	132·81
Chlorine	Cl	17	35·457
Chromium	Cr	24	52·01
Cobalt	Co	27	58·94
Columbium	Cb	41	93·1
Copper	Cu	29	63·57
Dysprosium	Dy	66	162·52
Erbium	Er	68	167·7
Europium	Eu	63	152·0
Fluorine	F	9	19·00
Gadolinium	Gd	64	157·26
Gallium	Ga	31	69·72
Germanium	Ge	32	72·60
Gold	Au	79	197·2
Hafnium	Hf	72	178·6
Helium	He	2	4·00
Holmium	Ho	67	163·4
Hydrogen	H	1	1·008
Illinium	Il	61	147
Indium	In	49	114·8
Iodine	I	53	126·932
Iridium	Ir	77	193·1
Iron	Fe	26	55·84
Krypton	Kr	36	82·9
Lanthanum	La	57	138·90
Lead	Pb	82	207·20
Lithium	Li	3	6·940
Lutecium	Lu	71	175·0
Magnesium	Mg	12	24·32
Manganese	Mn	25	54·93
Masurium	Ma	43	99·
Mercury	Hg	80	200·61
Molybdenum	Mo	42	96·0
Neodymium	Nd	60	144·27
Neon	Ne	10	20·2
Nickel	Ni	28	58·69
Nitrogen	N	7	14·008
Osmium	Os	76	190·8
Oxygen	O	8	16·000
Palladium	Pd	46	106·7
Phosphorus	P	15	31·027
Platinum	Pt	78	195·23
Polonium	Po	84	210·
Potassium	K	19	39·096
Praseodymium	Pr	59	140·92
Radium	Ra	88	225·95
Radon	Rn	86	222·
Rhenium	Re	75	187·
Rhodium	Rh	45	102·91
Rubidium	Rb	37	85·44
Ruthenium	Ru	44	101·77
Samarium	Sm	62	150·43
Scandium	Sc	21	45·10
Selenium	Se	34	79·2
Silicon	Si	14	28·06
Silver	Ag	47	107·880
Sodium	Na	11	22·997
Strontium	Sr	38	87·63
Sulphur	S	16	32·064
Tantalum	Ta	73	181·5
Tellurium	Te	52	127·5
Terbium	Tb	65	159·2
Thallium	Tl	81	204·39
Thorium	Th	90	232·15
Thulium	Tm	69	169·4
Tin	Sn	50	118·70
Titanium	Ti	22	48·1
Tungsten	W	74	184·0
Uranium	U	92	238·17
Uranium-X$_{II}$	U-X$_{II}$	91	234
Vanadium	V	23	50·96
Xenon	Xe	54	130·2
Ytterbium	Yb	70	173·6
Yttrium	Y	39	88·9
Zinc	Zn	30	65·38
Zirconium	Zr	40	91·

INDEX OF NAMES

SUBJECT INDEX

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