

# Ionization Potentials in the Teaching of Elementary Chemistry

HARRY H. SISLER and CALVIN A. VANDERWERF

*University of Kansas, Lawrence, Kansas*

## INTRODUCTION

**I**N A previous publication<sup>1</sup> the authors presented arguments for the application of recent advances in chemical theory to the teaching of elementary inorganic and organic chemistry. In the present paper an attempt has been made to describe in some detail the use of one group of data, namely, the gaseous ionization potentials of the elements, in a course which purports to follow the objectives set forth in the reference mentioned above.

No characteristics of an atom are more important than those which determine the manner in which the atom will enter into combination, *i. e.*, the type of chemical bonds which it will form with other atoms. The close dependence of the properties of a compound upon the nature of the bonds which hold its atoms together makes the knowledge of the important factors which affect the nature of these bonds of primary importance to the student in the "modern" chemistry course.

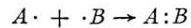
Let us consider the compound *AB* in which each atom has at least one valence electron. The bond between the atoms *A* and *B* may be of one of several types. If the compound is electrovalent the reaction between *A* and *B* may be represented by the equation,



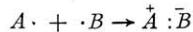
On the other hand, *AB* may be a completely nonpolar

<sup>1</sup> SISLER, H. H., AND C. A. VANDERWERF, "Modern theory: A tool in teaching elementary college chemistry," *J. Chem. Educ.*, 20, 479-83 (1943).

covalent compound in which case the two atoms share a pair of electrons equally as is shown by the equation,



Finally, *AB* may be a covalent molecule with an appreciable degree of polarity. In such a case, the pair of electrons is shared but unequally, giving what is known as a polar covalent bond.



The most important factor which determines to which of these types the molecule *AB* will belong is the relative attractions which the atoms *A* and *B* have for electrons. If there is a wide difference in this respect electron transfer will take place and an electrovalent compound will result. If, however, the atoms differ appreciably in their attraction for electrons but not sufficiently to allow for electron transfer, a polar covalent compound in which the pair of electrons are shifted toward the atom having the greater attraction for electrons will result. Finally, if the electron attractive powers of *A* and *B* are equal a nonpolar covalent molecule will be obtained.

## MEASURES OF ELECTRON ATTRACTION

Such terms as electronegativity, electron affinity, and electron attraction have, in the past, been used almost interchangeably and various methods have been proposed for their measurement. The lack of accurate definitions of these terms resulted in a considerable amount of confusion. This confusion, however, has

been largely eliminated in recent years, and it is now generally recognized that there are three different characteristic sets of data which, to a greater or lesser degree, may be used as indications of the relative attractions which atoms have for electrons, or for the relative ease with which atoms lose electrons. These three are: (a) electron affinities; (b) electronegativities; and (c) ionization potentials.

The electron affinity of an atom  $A$  is equal to the energy evolved in the reaction,  $A + e^- \rightarrow A^-$ . This quantity is particularly useful in discussing electrovalent compounds of the active nonmetals such as the halogens. In the case of the metals and the inert gases, however, the electron affinities may be considered to be zero and the general application of electron affinity data is, therefore, limited.

The electronegativities of atoms are more difficult to define but, as derived by Pauling, may be said to consist of numbers which represent the relative attractions which the various atoms have for electrons when the atoms are in stable covalent molecules.<sup>2</sup> The derivation of these quantities is based upon thermal data and upon theoretical concepts which are difficult for the elementary student to understand; their usefulness in an elementary course is therefore limited.

The gaseous ionization potentials of the elements are the energies required to bring about the removal of electrons from the atoms of the elements. The first ionization potential may be defined as the energy required to remove the least tightly bound electron from the unexcited atom and corresponds to the process,  $A \rightarrow A^+ + e^-$  where  $A$  is in its normal or unexcited state and is in the form of a dilute gas. The second ionization potential is the energy required to bring about the process  $A^+ \rightarrow A^{++} + e^-$ , the third ionization potential refers to the change  $A^{++} \rightarrow A^{+++} + e^-$ , and so on for the fourth, fifth, sixth, . . . and  $n$ th ionization potentials.

This concept is not a difficult one for the elementary student to understand nor are the methods by which ionization potentials are measured<sup>3</sup> beyond the comprehension of the student in general chemistry. Furthermore, the extent to which these experimental quantities may be correlated with the chemical properties of the elements and with their electronic configurations is a source of much satisfaction to the thinking student; it does much to remove the stigma of "just memory work" from elementary inorganic chemistry.

#### FACTORS AFFECTING IONIZATION POTENTIALS

In order that the significance of the variation in ionization potentials of the elements with their position in the periodic system may be clearly understood by the student the atomic characteristics which determine the ionization potentials must be discussed. A rigorous

discussion of this question would involve a wave mechanical treatment which is out of the question in an elementary course. However, a fairly simple, though admittedly approximate, approach to the problem is attained through a discussion of the following factors: (a) the charge on the atomic nucleus; (b) the shielding effect of the inner electron shells; (c) the atomic radius; and (d) the extent to which the least tightly bound electron may be considered to penetrate the more or less spherical smear of negative electrical charge constituted by the inner shells of electrons.

The relationship of the first of these factors to the ionization potential is fairly obvious for, as the positive charge on the nucleus of the atom is increased, the attractive force acting upon the negatively charged electrons will be augmented. It is likewise readily apparent that the existence of inner shells of electrons between the least tightly bound electron and the nucleus will "shield" the electron from the nucleus to a certain extent, for the attractive force of the nucleus for that electron will be partially counterbalanced by the repulsion effect of the previously added electron shells. The effect of increased atomic radius on the attractive force between the nucleus and the outer electrons presents no great difficulty to the elementary student.

The fourth factor, *i. e.*, the effect of penetration by the outer electron into shells of previously added electrons is not so apparent, but even this does not offer an insurmountable obstacle to the student's understanding. In the discussion of the electronic configurations of the elements and the periodic system the student has already become acquainted with the quantum numbers  $n$ ,  $l$ , and  $m$ , and their significance in terms of the Bohr-Sommerfeld picture of the atom. He has thus already been told that the shape, *i. e.*, the eccentricity, of an elliptical electron orbit depends upon the quantum number  $l$ —the smaller the value of  $l$  for a given value of  $n$ , the greater the eccentricity of the orbit. Thus for a given shell of electrons, the eccentricity of the orbits in the  $s$ ,  $p$ ,  $d$ , and  $f$  subshells (corresponding to values for  $l$  of 0, 1, 2, and 3 respectively) decreases in the order  $s > p > d > f$ . Since the nucleus is at one of the foci of an elliptical orbit it is apparent that for a given electron shell, the more eccentric the orbit of an electron, *i. e.*, the smaller the value of  $l$ , the closer the electron will come to the nucleus at one point in its orbit. This is illustrated in Figure 1.

It follows, therefore, that an electron in an  $s$  orbit will penetrate inner electron shells to a greater extent than a  $p$  electron in the same shell. Likewise, the penetration of inner electron shells by a  $d$  electron will be less than that of a  $p$  electron and more than that of an  $f$  electron in a given electron shell.

The relationship of the ionization potential of an element to the extent to which its least tightly bound electron penetrates inner electron shells arises from the fact that, while as long as the electron remains completely outside these inner electron shells it is shielded from the nucleus by them, as soon as the electron penetrates these inner shells this shielding effect is reduced,

<sup>2</sup> PAULING, L., "The Nature of the Chemical Bond," Cornell University Press, Ithaca, New York, 1940, Chapter II.

<sup>3</sup> GLASSTONE, S., "Textbook of Physical Chemistry," D. Van Nostrand Co., Inc., New York, 1940, pp. 49-52.

TABLE 1\*

Electron Configuration		Ionization Potentials (in Volts)							
		1st	2nd	3rd	4th	5th	6th	7th	8th
H	1s <sup>1</sup>	13.53							
He	1s <sup>2</sup>	24.48	54.14						
Li	1s <sup>2</sup> 2s <sup>1</sup>	5.37	75.28						
Be	1s <sup>2</sup> 2s <sup>2</sup>	9.28	18.14	153.10					
B	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>1</sup>	8.33	23.98	37.75	258.1				
C	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>2</sup>	11.22	24.28	47.7	64.19	389.9			
N	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>3</sup>	14.48	29.47	47.17	77.0	97.43	546.7		
O	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>4</sup>	13.55	34.93	54.88	76.99	109.19	137.48	733.0	
F	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>5</sup>	18.6	34.6	62.35	86.72	113.67	156.37	184.26	945.8
Ne	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup>	21.47							
Na	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>1</sup>	5.12	47.0						
Mg	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup>	7.61	14.97	79.72					
Al	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>1</sup>	5.96	18.75	28.32	119.37				
Si	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>2</sup>	8.12	16.27	33.30	44.95	165.6			
P	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>3</sup>	11.1	19.81	30.04	51.1	64.74	.....		
S	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>4</sup>	10.31	23.30	34.9	47.08	67.0	87.67	.....	
Cl	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>5</sup>	12.96	23.70	39.7	53.16	67.7	88.7	113.7	.....
A	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>2</sup> 3p <sup>6</sup>	15.69							

\* Values taken from SHERMAN, J., *Chem. Revs.*, 11, 138 (1932), and "The Chemistry & Physics Handbook," 23rd Edition, The Chemical Rubber Publishing Company, Cleveland, 1939, pp. 1528-9.

i. e., the net force of attraction of the nucleus for the electron is increased. Therefore, the ionization potential of an atom whose least tightly bound electron is in an *s* subshell is greater than if that electron were in a *p* subshell; similarly a *p* electron is more difficult to remove than a *d* or *f* electron, other factors being constant.

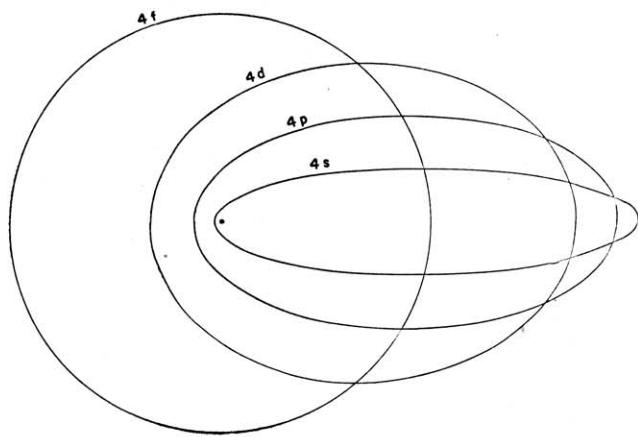


FIGURE 1.—ORBITS OF THE S, P, D, AND F ELECTRONS OF THE 4TH ELECTRON SHELL

Having been given these basic principles, the student is ready to undertake the discussion and interpretation of ionization potential data in relation to the atomic structure and properties of the elements. Ionization potentials for the elements of the first three series of the periodic system are given in Table 1.

#### APPLICATION AND INTERPRETATION OF IONIZATION POTENTIAL DATA

*Variation within a Single Period.* An examination of the data in Table 1 reveals a number of interesting relationships. In the first place it may be observed that there is a general increase in ionization potential as one proceeds from left to right in a given period in the periodic system. Thus, in the third period, sodium, the first element on the left, has the lowest ionization potential and argon, the last element in the period, the highest. This general increase is easily explainable as the result of the fact that, whereas within a given period the electrons which are being added as we go from element to element are at roughly similar distances from the nucleus, the charge on the nucleus is steadily increasing. The energy required to remove these outer electrons increases, therefore, with increasing nuclear charge. This is indicated graphically in Figure 2. Furthermore, it can be seen from this figure that following each inert gas, i. e., at the beginning of each new period in the periodic system, there is a marked drop in ionization potential. This is to be expected for this electron is going into a higher shell and is relatively well shielded by all the previously added electrons, in spite of the fact that, since this electron is an *s* electron, it penetrates these inner shells to a considerable extent.

The increase in ionization potential which takes place as one passes from left to right through a given period may be used as a basis for at least a partial explanation of the marked change in the chemical properties of the elements in this same order. At the left of a period we have elements of low ionization potentials which react

by losing electrons. These elements are, therefore, strong reducing agents and form ionic compounds in which they exist as positive ions. As one proceeds to the right the increase in ionization potentials is accompanied by a decrease in the tendency to lose electrons and an increase in the tendency to form covalent rather than electrovalent bonds. Near the end of the period elements of very high ionization potential which tend to form negative ions by taking up electrons and are, therefore, strong oxidizing agents are encountered. It must be pointed out to the student, of course, that ionization potential is not a measure of tendency to take up electrons. Atoms of very high ionization potential, however, which have unfilled places in their valence shells have, in general, high electron affinities. The trend from metallic to nonmetallic behavior is thus easily correlated with ionization potential data.

This point is emphasized even more decisively when it is pointed out that in order to attain stable inert gas configurations by the loss of electrons, not one but all the valence electrons must be lost. Thus the sodium atom must lose one electron, the magnesium atom two, aluminum three, silicon four, phosphorous five, sulfur six, and chlorine seven. The data in Table 1 indicate that the energy required to remove each successive electron from an atom is higher than for the previous electron. Thus the ionization potential required to re-

move all the valence electrons increases even more rapidly from left to right in a given period than do the ionization potentials for a single electron. This is shown by Table 2.

TABLE 2

IONIZATION POTENTIALS FOR COMPLETE REMOVAL OF ALL VALENCE ELECTRONS (IN VOLTS)

H								He
	13.53							54.14
Li	5.37	18.14	37.75	64.19	97.43	137.48	184.26	.....
Na	5.12	14.97	28.32	44.95	64.74	87.67	113.7	.....

In view of the data in this table it is not surprising that there is little tendency for elements other than those in the alkali and alkaline earth metal groups to form simple positive ions.

It is interesting to note from the data in Table 1 and from Figure 2 that whereas the general trend of first ionization potentials is an increase from left to right in a given period in the periodic system, there are minor retrogressions between beryllium and boron in the second period, between magnesium and aluminum in the third period, and between the corresponding elements in succeeding periods. Likewise, less important retrogressions occur between nitrogen and oxygen, phosphorous and sulfur, arsenic and selenium, and an

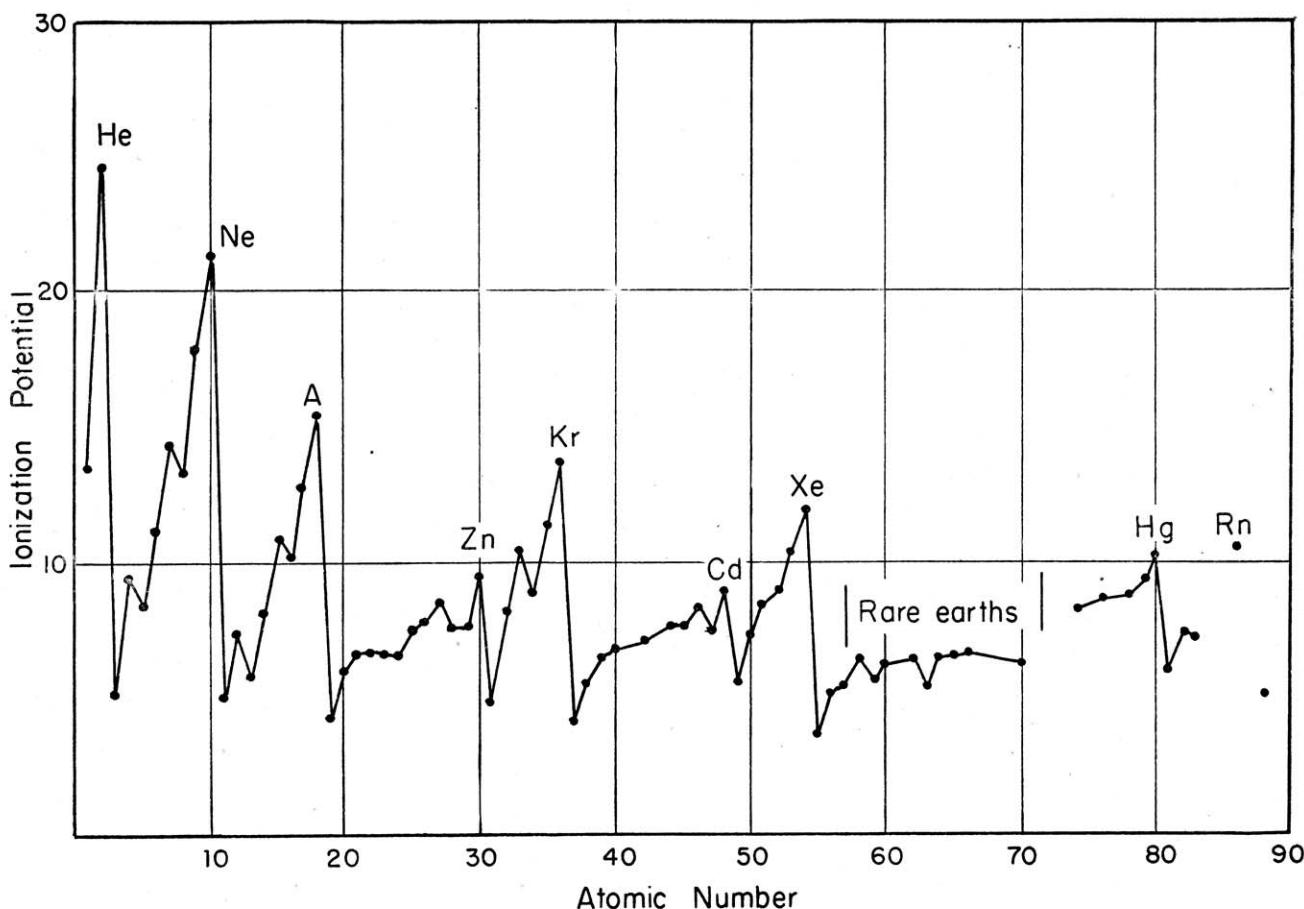


FIGURE 2.—FIRST IONIZATION POTENTIALS OF THE ELEMENTS

irregularity in the curve between antimony and tellurium. The reasons for the latter set of irregularities are obscure, but the former are readily understandable and give additional evidence for the validity of our approach to the problem.

Let us consider, for example, magnesium and aluminum. The magnesium atom has two valence electrons both of which in the normal state of the atom reside in the  $3s$  subshell. Since the capacity of an  $s$  subshell is only two, however, the third valence electron in the aluminum atom must go into the  $3p$  subshell. As has already been pointed out, a  $p$  electron, because of its lesser degree of penetration, is held less tightly than an  $s$  electron having the same main quantum number  $n$ . In this case, the smaller attraction of the nucleus of the atom for the  $p$  electron as compared with that for an  $s$  electron more than overcomes the effect of the increased nuclear charge so that the first ionization potential of aluminum is less than that of magnesium. It is to be noted, however, that this retrogression applies only to the first ionization potentials of these elements, for the second ionization potential of aluminum is definitely higher than the second ionization potential of magnesium. This is as we would expect for here we are dealing with  $3s$  electrons in both elements, and the normal effect of increasing nuclear charge operates. A similar argument applies equally well to beryllium and boron, zinc and gallium, cadmium and indium, and mercury and thallium.

Since the retrogression of first ionization potentials in these instances depends upon the smaller degree of penetration of a  $p$  electron as compared to an  $s$  electron, we would expect this effect to increase in the order  $B_e-B > Mg-Al > Zn-Ga > Cd-In > Hg-Tl$ , for penetration effects are intensified as the number of electron shells underlying the valence shell increases and more shells are thus made available through which the valence electron can penetrate. That this expectation is, in a general way, borne out may be observed from Figure 2.

*Ionization Potential Data and the Number of Valence Electrons.* The data in Table 1 provide excellent experimental evidence for the number of valence electrons in an atom. In the case of lithium, the first ionization potential is very low (5.37 volts) but the second is very much higher (75.28 volts), indicating that one electron is much more easily removed than the rest; in other words, a lithium atom has one valence electron. There is a marked increase between the second and third ionization potentials of beryllium, indicating two valence electrons. Similarly, sharp increases occur between the third and fourth ionization potentials of boron, the fourth and fifth ionization potentials of carbon, the fifth and sixth of nitrogen, the sixth and seventh of oxygen, and the seventh and eighth of fluorine, indicating three, four, five, six, and seven valence electrons, respectively. Analogous conclusions may be drawn for the elements from sodium to chlorine, respectively, and for the heavier elements for which sufficient ionization potential data are available.

*Variation within a Given Group.* A comparison of the

ionization potentials of the elements in a given group in the periodic system is of value to the student in the study of the gradation of properties within the given group. As is indicated by the data in Table 3, with certain minor exceptions which are discussed below, the ionization potentials of the elements in a given family decrease as one goes from the top of the group to the bottom.

TABLE 3

IONIZATION POTENTIALS REQUIRED TO REMOVE ALL VALENCE ELECTRONS

Group Ia	Group IIa	Group VIA
Li 5.37	Be 18.14	O 137.48
Na 5.12	Mg 14.97	S 87.67
K 4.32	Ca 11.82	Se 81.4
Rb 4.16	Sr 10.98	Te 72.0
Cs 3.88	Ba 9.95	

This fact may be explained to the elementary student as follows:

It is pointed out that, in a given group in the periodic system, the effect of higher nuclear charge would tend to cause the lower elements in the group (*i.e.*, those of higher atomic number) to have higher ionization potentials than the higher members (those of low atomic number). This effect is more than counterbalanced, however, by the fact that in the atoms of higher atomic number the valence electrons are farther from the nucleus and are shielded by more intervening shells of electrons than in the atoms of lower atomic number of that same family. Of course, the increased shielding effect is reduced somewhat by penetration of the valence electrons within these inner shells, but the over-all effect is that the ionization potentials decrease from the upper members of the family to the lower.

This fact provides the explanation for the rule that, in general, metallic properties become more prominent within a given periodic family as the atomic number is increased, while nonmetallic characteristics become less prominent in the same direction. Combined with the statement above concerning trends from left to right in the periodic system, this gives a basis for understanding the fact that in the new "long" forms of the periodic chart the most active metal and nonmetal are at the lower left and upper right of the chart, respectively. It accounts, furthermore, for the fact that the nonmetals may be roughly separated from the metals by means of a diagonal line on the periodic chart. This line separates beryllium from boron, aluminum from silicon, germanium from arsenic, antimony from tellurium, and polonium from element No. 85, as is shown in Figure 3.

There is, of course, almost no limit to the number of applications of the above principles to specific chemical situations. The increased tendency for the elements in the alkaline earth group to react with water is an example. Beryllium fails to react even with boiling water, magnesium reacts with boiling water but is unaffected by cold water, calcium, strontium, and barium react even with cold water with vigor increasing in that order. This order corresponds exactly to that of de-

creasing ionization potentials, and may thus be explained. Likewise the order of the activities of the halogens affords another example of the application of the above principles. Thus, fluorine displaces chlorine from solutions of chloride ion, chlorine displaces bromine from bromides, and bromine displaces iodine from iodides, which is in agreement with the fact that ionization potentials of these elements decrease in the order  $F > Cl > Br > I$ , and their electron affinities decrease in the same order. The decreasing acidity of the hydroxides of trivalent nitrogen, phosphorous, arsenic, antimony, and bismuth, as well as similar trends in other groups, is readily interpreted in terms of decrease in ionization potentials with increasing atomic number in a given family.

*Other Applications.* For those who still use the Mendeleeff form of the periodic chart, the explanation for the marked difference in chemical and physical properties of the elements of the two subgroups belonging to a single main group is a difficult pedagogical problem. Of course, in the "long" form of the chart these subgroups are in different columns and are considered, in fact, to be separate groups, and the questions as to reasons for their differences do not arise so commonly. By the application of the principles discussed above the explanation of these differences, *e. g.*, of the lesser activity of the copper family elements than of the corresponding alkali metals, is not particularly difficult. The electronic configurations, and atomic numbers of corresponding members of the two families are listed in Table 4. The ionization potentials of the

TABLE 4

Element	At. No.	Configuration			
K	19	2	2,6	2,6	1
Cu	29	2	2,6	2,6,10	1
Rb	37	2	2,6	2,6,10	2,6
Ag	47	2	2,6	2,6,10	2,6,10
Cs	55	2	2,6	2,6,10	2,6,10
Au	79	2	2,6	2,6,10	2,6,10,14
					2,6,10 1

copper family elements are much higher than those of the corresponding alkali metals, which is in accordance with the much greater reactivity of the latter. If we were to assume that the single valence electron in each atom of corresponding members of the two families is completely shielded from the nucleus by the intervening shells of electrons, *i. e.*, that the electron did not penetrate these shells, we would expect that the ionization potentials and reactivities of the corresponding members of the two families would be approximately equal except for effects due to the somewhat smaller radius of the copper family elements which results from their greater nuclear charge. However, in the normal state of the atoms of these two families the valence electron is an *s* electron which, therefore, penetrates to a considerable extent the inner electron shells. It is readily apparent that the penetration of the valence electron beneath an inner shell of 18 electrons in an atom of the copper family will produce a much greater lowering of its energy than the penetration of the

valence electron beneath an inner shell of 8 electrons in an alkali metal atom. The copper family elements, therefore, release their valence electrons much less readily than the alkali metal atoms and are, hence, less active.

The elements (At. Nos. 72 through 82) immediately following the rare earth elements in the periodic system provide an interesting series of exceptions to the general rule that in a given family ionization potentials decrease and metallic activity increases. These elements are enclosed by a dotted line in Figure 3. For each of these elements for which ionization potential data are available the ionization potential is higher than that of the preceding element in that family. The chemical properties are in accord with this fact. It is interesting to note, also, how many of the very "noble" metals, such as iridium, platinum, gold, and mercury, four of our least active metals, are included in this series. Furthermore, the normal situation where the atomic volumes of the elements in a given family of the periodic system increase with increasing atomic number is also altered for these elements for they have atomic volumes, which are only very slightly, if at all, larger than the preceding elements of their respective families. It is desirable that the student be given an explanation for this apparently anomalous situation. Such an explanation follows readily from the general principles set up in the beginning of this discussion. The principal factor involved is the extraordinary increase in nuclear charge which takes place between these elements and the preceding elements in their respective families.

The increase in nuclear charge between rubidium and cesium, and between strontium and barium is equal to 18 units in each case. However, all the elements of atomic numbers 72 through 86 possess nuclear charges 32 units larger than the preceding elements in their respective families, as a result of the series of rare earth elements (At. Nos. 57 through 71) in which the 4*f* subshell is filled with 14 electrons giving the fourth shell a total of 32 electrons. It might be supposed that the

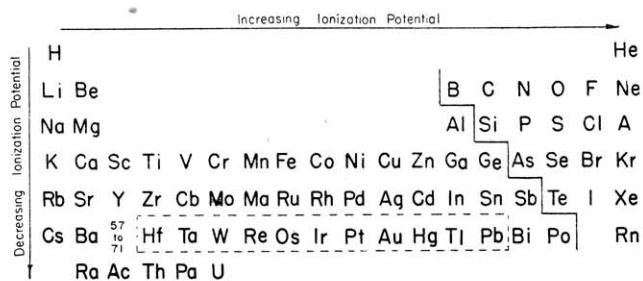


FIGURE 3

effect of this extraordinarily large increase in charge would be balanced by the shielding effect of the additional electrons in the fourth shell. However, this is not true, since the outer electrons penetrate this shell to some extent; the ionization potentials of these elements are, therefore, much greater and the atomic

(Continued on page 401)

### IONIZATION POTENTIALS IN ELEMENTARY CHEMISTRY

(Continued from page 395)

volumes are smaller than would have been the case had the rare earth interlude not occurred. This contraction in volume occurring in the rare earth series is commonly called the lanthanide contraction.

Students who have been trained to utilize ionization potential data in predicting the nature of the bonds formed between atoms begin their study of organic chemistry with an appreciation of the fact that the valence bonds of the carbon atom must almost invariably be covalent in nature. This truth, coupled with the fact that only carbon, of all the known elements, possesses in its normal covalent state a completely shared octet of electrons which cannot be expanded, accounts for the low reactivity of the carbon-to-carbon bond, and therefore the unique position of that element in the organic world, not as a whim of nature, but as a logical consequence of the electronic configuration of carbon.

Students who understand the relationship between the properties of a molecule and the nature of the bonds existing between its atoms are quick to perceive that

many of the so-called "differences between typical organic and typical inorganic compounds" are actually merely the differences between nonpolar covalent compounds, on the one hand, and ionic compounds on the other. This knowledge is particularly useful in that it simplifies the problem of accounting for those many cases in which organic compounds depart from the "typically organic" behavior. The ability to interpret ionization potential data likewise provides the student in organic chemistry with a starting point in an approach to the understanding of the electron displacements ( $I$  and  $T$ ) which are all important in determining both the physical and chemical properties.

The points which have been made in this discussion illustrate but do not exhaust the application of ionization potential data in the teaching of inorganic chemistry. In conclusion, the authors would reiterate their conviction of the desirability of incorporating developments in theory of matter into the elementary chemistry course. It is only by so doing that the teacher can justify his claim to be teaching a science.