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Chemistry of Isotopes

JACOB BIGEISEN

MASTER

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FOREWORD

The Brookhaven Lectures, held by and for the Brookhaven staff, are meant to provide an intellectual meeting ground for all scientists of the Laboratory. In this role they serve a double purpose: they are to acquaint the listeners with new developments and ideas not only in their own field, but also in other important fields of science, and to give them a heightened awareness of the aims and potentialities of Brookhaven National Laboratory.

Before describing some recent research or the novel design and possible uses of a machine or apparatus, the lecturers attempt to familiarize the audience with the background of the topic to be treated and to define unfamiliar terms as far as possible.

Of course we are fully conscious of the numerous hurdles and pitfalls which necessarily beset such a venture. In particular, the difference in outlook and method between physical and biological sciences presents formidable difficulties. However, if we wish to be aware of progress in other fields of science, we have to consider each obstacle as a challenge which can be met.

The lectures are found to yield some incidental rewards which heighten their spell: In order to organize his talk the lecturer has to look at his work with a new, wider perspective, which provides a satisfying contrast to the often very specialized point of view from which he usually approaches his theoretical or experimental research. Conversely, during the discussion period after his talk, he may derive valuable stimulation from searching questions or technical advice received from listeners with different scientific backgrounds. The audience, on the other hand, has an opportunity to see a colleague who may have long been a friend or acquaintance in a new and interesting light.

The lectures are being organized by a committee which consists of representatives of all departments of the Laboratory. A list of the lectures that have been given and of those which are now scheduled appears on the back of this report.

Gertrude Scharff-Goldhaber

The drawing on the cover is taken from a 5th Century B.C. relief on the Acropolis in Athens, the "Dreaming Athena," by an unknown sculptor.

Chemistry of Isotopes

JACOB BIGELEISEN

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INTRODUCTION

Dr. Jacob Bigeleisen was [REDACTED]
After receiving his A.B. degree from New York University, he studied for his M.S. degree at the State College of Washington. His thesis research there, which was guided by Otto Redlich, dealt with the thermodynamic properties of electrolytic solutions. He then proceeded to the University of California, Berkeley, where he carried out his Ph.D. research on photochemical studies under the direction of Gilbert N. Lewis. Dr. Bigeleisen spent the war years on the Manhattan Project. After short periods at Ohio State University and the University of Chicago, he came to Brookhaven in 1948. He has been here since then except for a stint as Visiting Professor at Cornell and a year at the Eidgenossische Technische Hochschule, from which he has just returned.

Dr. Bigeleisen's paper on equilibrium isotope effects, published jointly with Maria Goeppert Mayer in 1947, is considered a classic in the field today. This paper was followed, shortly after his arrival at Brookhaven, by a general study of isotope effects on chemical rate processes which contains the well-known "Bigeleisen formulation." Over the years, Dr. Bigeleisen has studied, both theoretically and experimentally, isotope effects on rates, on chemical equilibria, and recently on vapor pressures. A characteristic of his approach has been the ability to carry out simultaneously both experimental and theoretical study. His contributions won him the American Chemical Society Award for Nuclear Applications to Chemistry in 1958.

It is a great pleasure to introduce Dr. Bigeleisen. The title of his talk tonight is, appropriately enough, Chemistry of Isotopes.

MAX WOLFSBERG

Chemistry of Isotopes

The subject of this lecture is the origin of the differences in physical and chemical properties of isotopes. Phenomena arising from differences in nuclear decay and reaction properties will not be considered. The differences in physical and chemical properties can be used in applied and pure science. In the applied field the obvious application is to isotope separation. In pure science, the differences in physical and chemical properties of isotopes can be used to study such diverse problems as molecular structure, quantum effects, and chemical kinetics in the field of chemical physics; paleotemperature and biogenesis in geochemical studies; and the growth and morphology of plants and animals in the biochemical and medical sciences. After an introduction to the principles of isotope chemistry, some examples of the application of isotope effects to these problems in pure science will be discussed.

What is meant by "isotope effects" or the "chemistry of isotopes" can be illustrated by three experiments which can be done with water. Natural water is a mixture of light and heavy water. If ordinary water is equilibrated with hydrogen gas, say by means of a platinum catalyst, then at equilibrium the ratio of deuterium (heavy hydrogen) atoms to protium (light hydrogen) atoms in the water is found to be 3.7 times the ratio of deuterium to protium atoms in the hydrogen gas. This ratio, 3.7, is found if the equilibration is done at room temperature, and it becomes progressively smaller, approaching unity, as the equilibration temperature is increased. The deuterium atoms behave differently from the protium atoms in these two compounds, water and hydrogen. This is ordinary chemistry. If water is electrolyzed, hydrogen is obtained at the negative pole (cathode) and oxygen at the positive pole (anode). Comparison of the deuterium-to-protium ratio in the water being electrolyzed with that in the evolved hydrogen shows the water to be richer in deuterium by a factor of 3.7 to 9. The precise difference depends on the electrode materials and the conditions of electrolysis. Finally, if the deuterium abundance in liquid water is compared with that in vapor in equilibrium with the liquid, the liquid is again

found to be richer in deuterium, but this time by a factor of only 1.07. The discrimination between deuterium and protium in water depends on the process involved.

In classical theory there are no differences in physical or chemical processes of systems at equilibrium. In processes which depend on the velocities of particles, protium may react more rapidly than deuterium by a factor of $\sqrt{2}$. This is a consequence of the kinetic theory, from which it is deduced that the kinetic energy, $\frac{1}{2}mv^2$, is proportional to the absolute temperature. Thus at a given temperature the velocity, v , is proportional to $(1/m)^{1/2}$, or $v_H = \sqrt{2}v_D$. The classical theory is seen to be inadequate to explain the observations on the differences in chemical properties of protium and deuterium in both equilibrium and rate processes. One must, therefore, turn to the quantum theory to look for an explanation of the observed phenomena. In fact, it is just for this reason that isotope effects are so useful in the study of quantum effects in chemistry.

BASIS OF CHEMICAL ISOTOPE EFFECTS

The energy of a molecule or a collection of molecules is described in terms of the electronic energy plus the translation, rotation, and vibration of the molecules. To this must be added the energies associated with the mutual interactions of these simple motions. At room temperature the translation and rotation of all gases is adequately described by classical mechanics. Anything which is classical contributes nothing toward the difference in chemical properties of isotopes. From observations in atomic and molecular spectra, it is concluded that the differences in electronic energies between protium and deuterium in their various compounds cannot result in an effect greater than 0.5%. This leaves as the sole potential large origin of "isotope effects" the molecular vibrations.

Figure 1 shows schematically the energy of a diatomic molecule as a function of the distance between the two atoms. When the atoms are far apart, they move independently of one another. As they approach one another there is an attrac-

Figure 1. Schematic potential energy curve for the interaction of two atoms in a stable molecule or between two molecules in a liquid or solid.

tive force. If the atoms get closer to one another than the distance corresponding to the stable molecule, they repel one another and the energy increases rapidly. Another way of looking at this repulsion is simply that the atoms have finite size and both cannot occupy the same place without one or both being squeezed. This compression takes work. According to the quantum theory, the molecule cannot assume any energy on the continuous curve shown in Figure 1 but is restricted to certain discrete energy levels. The lowest level is not at the minimum of the energy curve but is $\frac{1}{2} h\nu$ above it, where h is Planck's constant and ν is the frequency with which the atoms in the molecule vibrate with respect to one another. The vibrational frequency of a molecule depends inversely on the masses of the atoms in the molecule. It also depends directly on the forces holding the atoms together in the molecule. The forces in a molecule depend on the electron arrangements, the nuclear charges, and the positions of the atoms in the molecule. For our purposes it suffices to consider them identical for isotopes. However, since isotopes differ in mass, two molecules of the same chemical formula but of different isotopic species, e.g., H_2 and D_2 , will have different vibrational frequencies. They will have different zero-point energies; the molecule with a heavy isotope will have a lower zero-point energy than the same chemical species with a light isotope. This is shown schematically in Figure 1, where the upper horizontal line represents the zero-point energy of the light molecule and the lower line that of the heavy one.

Now consider the energy necessary to dissociate the diatomic molecule when all the molecules are in the ground vibrational state. This energy is E_L for the light molecule but E_H for the heavy molecule. There is a difference in the energies necessary to dissociate the two kinds of isotopic molecules.

This energy difference for dissociation affects the stability of the molecule. This principle can be illustrated with a familiar process, the evaporation and condensation of water, which is represented as the chemical reaction

The process involves both an energy and an entropy change. It takes energy to evaporate the liquid, and conversely heat is liberated when steam condenses. For simplicity the entropy change will be neglected, although it is by no means inconsequential. At equilibrium in a closed system the ratio of concentration (i.e., the number per unit volume) of water molecules as steam, $(H_2O)_g$, divided by the concentration of water molecules as liquid, $(H_2O)_l$, is constant at any particular temperature:

$$\frac{(H_2O)_g}{(H_2O)_l} = K = \text{pressure}.$$

If the concentration of the liquid is used as the unit of concentration, then the constant, which is a function of temperature, is the vapor pressure of the liquid. If the closed system initially at equilibrium contains a movable piston-cylinder and the system is suddenly compressed by increased pressure on the piston, then steam will condense and liberate heat. The converse process will cause water to evaporate and cool the water. This is how refrigerators operate. If the system is subjected to some external force or strain, it readjusts itself to remove the strain. For the system liquid water and steam there is only one equilibrium state at any given temperature; namely, the steam must have the vapor pressure of water at that temperature. An increase in the temperature results in an increase in the vapor pressure. It takes energy to vaporize the water, and the chance that a molecule will have enough energy to get away from the other molecules in the liquid increases with increasing temperature.

These same principles govern the course of chemical reactions. Consider the chemical reaction

in which radioactive bromine atoms (Br^*) exchange positions with ordinary bromine atoms (Br) between two chemical species. The species to be discussed are lithium bromide dissolved in water and benzyl bromide dissolved in ethylene diacetate. The ethylene diacetate (abbreviated as org) and water are sparingly soluble in one another. The lithium bromide is sparingly soluble in org, and the benzyl bromide is sparingly soluble in water. If initially the radioactivity is in the lithium bromide and none in the benzyl bromide, the radioactivity will increase in the benzyl bromide and decrease in the lithium bromide.¹ These results can be summarized by the equation

At equilibrium, when the radioactivity in the benzyl bromide is no longer increasing with time and the activity of the bromide dissolved in water is no longer decreasing, the following relationship is found between the concentrations of the different molecular species:

$$\begin{aligned} \frac{(\phi\text{CH}_2\text{Br}^*)(\text{Br}^-)}{(\phi\text{CH}_2\text{Br})(\text{Br}^*)} &= \frac{(\phi\text{CH}_2\text{Br}^*/\phi\text{CH}_2\text{Br})}{(\text{Br}^*/\text{Br}^-)} = K, \\ K &= \frac{(\text{Br}^*/\text{Br})_{\text{org}}}{(\text{Br}^*/\text{Br})_{\text{H}_2\text{O}}} = 1.00 . \end{aligned}$$

The specific activities of the water and organic phases are equal at equilibrium. The radioactive bromine acts as a random tracer. The random distribution results from the fact that the zero-point energy difference between the various bromine isotopes is small with respect to their thermal energy. The vibrations of the molecules in which the bromine atoms participate behave as though they were classical, and there is no discrimination between isotopes.

Now consider the isotopic exchange reaction involving protium and deuterium between water and hydrogen sulfide:

The chemical binding of hydrogen to oxygen is much stronger than that of hydrogen to sulfur. The H_2O and H_2S molecules have similar structures. Each molecule has three vibrational frequencies, and there is a one-to-one correspondence

of the types of vibrations in the two molecules. The important difference is that each vibration in the water molecule has a higher frequency than that in the hydrogen sulfide molecule. For such a molecule the zero-point energy, E^0 , is

$$E^0 = \sum_{i=1}^3 \frac{1}{2} h\nu_i .$$

The zero-point energies can be depicted by the following level scheme:

The energy change in the chemical reaction is

$$\begin{aligned} \Delta E &= E^0_{\text{H}_2\text{S}} + E^0_{\text{HDO}} - E^0_{\text{HDS}} - E^0_{\text{H}_2\text{O}} , \\ &= (E^0_{\text{H}_2\text{S}} - E^0_{\text{HDS}}) - (E^0_{\text{H}_2\text{O}} - E^0_{\text{HDO}}) . \end{aligned}$$

The relationship of the relative concentrations of different molecules at equilibrium is

$$K = \frac{(\text{HDO})/(\text{H}_2\text{O})}{(\text{HDS})/(\text{H}_2\text{S})} = \frac{(\text{D}/\text{H})_{\text{H}_2\text{O}}}{(\text{D}/\text{H})_{\text{H}_2\text{S}}} = 2.35 \quad \text{at } 25^\circ\text{C} ;$$

$\ln K \cong -\Delta E/RT$ where R is the gas constant;

$$\ln K \cong \frac{(E^0_{\text{H}_2\text{O}} - E^0_{\text{HDO}}) - (E^0_{\text{H}_2\text{S}} - E^0_{\text{HDS}})}{RT} .$$

The deuterium atoms concentrate preferentially in the water.* *The heavy isotope goes preferentially to the chemical compound in which the element is bound most strongly.* This has been shown to hold when the energy change for the exchange reaction can be calculated by assuming all the molecules to be in the

*This process is a very efficient method for producing heavy water and is used on a large scale at Savannah River. Isotopes of other elements such as boron and nitrogen are separated in a similar manner.

lowest vibrational energy state and the entropy change to be negligible.

What happens to the isotope distribution when the molecules get distributed among the higher vibrational states? At very high temperatures the rule is definitely for the heavy isotope to prefer the chemical species in which the element has the strongest binding. This will generally, but not necessarily, be the behavior over the entire temperature range. In some special cases deviations will occur from this rule. These are matters of interest for the expert. Figure 2 shows the equilibrium constants for the exchange of tritium and protium between the hydrogen halides and water above 1000°K. The isotope fractionation faithfully reproduces the chemical affinity of the halogens, $F > Cl > Br > I$. The tritium favors the HF most and the HI least. The isotope distribution rule is valid at high temperatures, where the higher vibrational states become appreciably populated. In fact, it can be shown through the methods of statistical mechanics that this high temperature behavior is the fundamental rule.²

The quantitative treatment of isotopic exchange equilibria is not much more complicated than the zero-point energy approximation. For the isotopes of all elements other than hydrogen, the zero-point

Figure 2. Fractionation of tritium and protium between the hydrogen halides and water above 1000°K. (Exchange equilibrium: $TX + H_2O = HX + HTO$.)

Figure 3. Bigeleisen-Mayer isotope effect function, $G(u)$ as a function of u .

Table 1

Protium-Deuterium Equilibrium Fractionation Factors* for Some Compounds Compared With Water at 25°C

O—H Bonds		C—H Bonds	
Acetic acid	1.03	Benzene	0.87
Methanol	1.0	Naphthalene	0.87
Ethanol	1.11	Toluene (all)	0.9
<i>n</i> -Amyl alcohol	1.09	Acetone	0.85
Benzyl alcohol	1.10	Pentene	0.85
Phenol	1.07	Cyclohexane	0.88
N—H Bonds		S—H Bonds	
Ammonia	0.954	H ₂ S	0.455
Formamide	1.04	EtSH	0.43
Acetamide	1.02	<i>i</i> -C ₃ H ₇ SH	0.495
Aniline	1.11	<i>n</i> -C ₃ H ₇ SH	0.493

*Fractionation factor is defined as $(D/H)_{\text{compound}} / (D/H)_{\text{water}}$.

Figure 4. Experimental and calculated equilibrium constants for the exchange reaction $HT + H_2O = H_2 + HTO$.

energy factor $\frac{1}{2}$ is found to be replaced by a function of the vibrational frequencies $G(u)$ such that

$$G(u) = \frac{1}{2} - \frac{1}{u} + \frac{1}{e^u - 1}$$

where $u = hv/kT$. The quantities h and k are respectively Planck's and Boltzmann's constants. The function is plotted in Figure 3. At low temperatures or large u , $G(u)$ approaches the zero-point energy value of $\frac{1}{2}$ asymptotically. At high temperatures $G(u)$ goes to zero, and the equilibrium differences in chemical properties of isotopes vanish. Some mathematical approximations in the derivation make the $G(u)$ function not applicable to the isotopes of hydrogen. Nevertheless, the same principles hold. The hydrogen isotope effect at room temperature is primarily a function of the type of chemical bond. Table 1 lists the protium-deuterium equilibrium fractionation factors for a variety of compounds compared with water at 25°C. It is clear that compounds with similar types of bonds have similar fractionation factors or isotope effects. To find appreciable isotope effects one has to compare chemical compounds with significant differences in chemical bonding. The largest effects are to be observed when the isotopes are intercompared in a compound having strong chemical bonding with the free monatomic element, which has no bonding. (The function of Bigeleisen and Mayer shows that the free element is a logical basis from which to intercompare isotope effects.)

QUANTITATIVE ASPECTS OF ISOTOPE EFFECTS

It is important to demonstrate the degree to which our ideas about the origin of equilibrium isotope effects agree with experiment. For this pur-

pose detailed knowledge is required about the vibrational frequencies of polyatomic molecules and their change with isotopic substitution. Some of this information can be obtained from spectroscopic measurements; some depends on calculations based on well established principles of mechanics.³ For the isotopes of hydrogen, the theory of small vibrations is not adequate, and the large amplitudes of the vibrations associated with small masses (0.1 Å) complicate the analysis of the energy levels. The analyses of the spectra of the H_2 and H_2O molecules have reached a high degree of refinement. The equilibrium constants for the reaction

calculated from the spectra without any arbitrary or adjustable constants, are shown as the solid line in Figure 4. The circles are the experimental data obtained some 20 years ago by J.F. Black and H.S. Taylor⁴ in the temperature range 15° to 300°C. The calculations are in satisfactory agreement with experiment. Table 2 gives a comparison of experimental and theoretical equilibrium constants for some typical isotope exchange reactions of the light elements.

The theory of equilibrium isotope effects for gaseous molecules has reached a rather advanced stage of development. Theorems and corollaries have been derived which relate the isotope effect directly to the structure and chemical bonding in the molecule. For complex molecules procedures are now available for direct computation with high speed digital computers.⁶ The necessary input data are the structure and forces in the molecule. Conversely, experimentally determined isotope effects can be used to evaluate such quantities.

Table 2

Comparison of Experimental and Theoretical Equilibrium Constants for Isotope Exchange Reactions*

Reaction	$T, ^\circ K$	K_{exp}	K_{calc}
$H_2 + DI = D_2 + HI$	600	1.21 ± 0.02	1.207
$HT + H_2O = H_2 + HTO$	298	6.26 ± 0.1	6.35
$\frac{1}{2} C^{16}O_2 + H_2^{18}O(liq) = \frac{1}{2} C^{18}O_2 + H_2^{16}O(liq)$	273	1.046	1.044
$^{15}NH_3 + ^{14}NH_4^+ = ^{14}NH_3 + ^{15}NH_4^+$	298	1.034	1.035
$H^{12}CN + ^{13}CN = H^{13}CN + ^{12}CN$	295	1.026	1.030
$^{12}CO_2 + ^{13}CO_2 = ^{13}CO_2 + ^{12}CO_2$	273	1.017	1.016
$\frac{1}{2} C^{16}O_3 + H_2^{18}O = \frac{1}{2} C^{18}O_3 + H_2^{16}O$	273	1.036	1.033

*Taken from the compilation by Urey.⁵

The questions naturally arise as to how large isotope effects can get, and how they depend on the mass of the element and its chemistry. Theory can provide unequivocal answers to these questions. The logarithm of the chemical isotope effect in a given compound is proportional to the mass difference of the isotopes divided by the product of the masses of the isotopic atoms (which is a characteristic constant for the isotopes of a given element) multiplied by the force constant acting to restore this atom when it is displaced from its equilibrium position in the molecule. The force constant is a measure of the strength of the chemical binding. Mathematically,

$$\log(\text{isotopic effect}) \sim (\Delta m/mm') \cdot \text{force constant}.$$

The maximum isotope effect between two compounds occurs when the force constant in one of them is zero. This compound is the free gaseous atom. A histogram of maximum isotope effects for the first half of the periodic table is given in Figure 5. A number of features are noteworthy. The effects range from a factor of near 100 for protium-tritium to 0.01 for the isotopes of phosphorus and iodine. Some salient chemical facts become manifest if the effects in S, N, O, C, and F are considered. The strengths of the chemical bonds are in the order written, and these dominate the features of the isotope effects for these elements, which have comparable values of $\Delta m/mm'$.

This brief exposition of the principles of isotope chemistry can serve as a background for a discus-

Figure 5. Maximum isotope effects for elements in the first half of the periodic table.

sion of some applications of isotope effects in chemical physics, geochemistry, and biology. The choice of examples is representative of some of the types of problems which have been attacked both here at Brookhaven and elsewhere. The application of isotope effects in the area of chemical kinetics will be explicitly omitted; this topic is worthy of an entire lecture by itself, since the study of isotope effects has had a profound impact on this area of chemistry.

ISOTOPE EFFECTS IN THE STUDY OF INTERMOLECULAR FORCES

The subject of isotope effects in the interaction of molecules with one another is one in which major progress has been made in the last few years. A number of new phenomena have been discovered through this unique tool, and some of the results are reviewed here. The experiments consist in the measurements of the vapor pressures of pure isotopic compounds or alternatively the equilibrium fractionation of an isotopic mixture between two phases, e.g., vapor and liquid. The behavior of the isotopes of helium will not be discussed because, although fascinating, it constitutes a subject in its own right.

Figure 6. Vapor pressure ratio $^{20}\text{Ne}/^{22}\text{Ne}$ (corrected to ideal gas reference state) as a function of temperature. Upper curve, solid neon; lower curve, liquid neon.

The simplest systems are naturally the monatomic rare gases. Figure 6 shows the results obtained for neon.⁷ These experiments were undertaken to check on older experiments described in the literature whose results were in disagreement with a theory for this effect worked out in 1919 by Lord Cherwell. The new data confirm Cherwell's theory, but also provide new information. There is a striking difference between the magnitudes of the vapor pressure ratios at the melting point. The behavior of the solid and liquid can each be represented by a curve, which can be described by a characteristic temperature or energy quantity. These characteristic temperatures are in reasonable agreement with other types of measurements, which should give the same quantity. The discrepancy that exists has provided evidence for the fact that the lattice oscillations have large amplitude and cannot be described as idealized springs. The difference between the liquid and solid is a manifestation of the change in structure that is responsible for the melting process.

Figure 7. Distribution coefficient for HT in H_2 between liquid and vapor as a function of temperature. Dashed curve is that calculated for a solution of D_2 in H_2 .

Neither Lord Cherwell's theory nor subsequent generalizations of it are capable of describing the behavior of polyatomic molecules. This required the development of a new theory, which predicted a number of new phenomena. Experiments have confirmed these predictions and then led to the discovery of a number of new phenomena connected with rotation in liquids and solids. These in turn have led to further amplification of the theory. When two molecules interact with one another, not only is the motion of the molecule through space different from that of the free molecule, but also the rotation of the molecule and the vibrations of the molecule are affected. Further, each of these motions becomes coupled to the other types of motions. A striking example of this is shown in Figure 7, which gives experimental data on the fractionation of HT in a dilute solution in H_2 between liquid and vapor. Between 20° and $28^\circ K$ the HT molecule is 22 to 24% more volatile than D_2 , which has the same molecular weight. The difference in volatility between HT and D_2 arises from the fact that the center of grav-

Figure 8. Separation of the isotopic N_2O molecules in a distillation column as a function of the number of plates in the column.

ity of the HT molecule is displaced from the geometric center of the molecule. This leads to a coupling of the translation and rotation in this molecule, which is absent in D_2 . This is a new type of quantum effect discovered by means of isotope effects.⁸ The theory of the effect has been developed⁹⁻¹² and is in good agreement with experiment.

A simpler type of rotational effect was demonstrated for nitrous oxide. This molecule is linear, with the structure NNO. The data in Figure 8 show the sequence of the vapor pressures to be $^{14}N^{14}N^{16}O > ^{14}N^{15}N^{16}O > ^{15}N^{14}N^{16}O > ^{14}N^{14}N^{18}O$. The difference between the two species with ^{15}N arises from hindered rotation of the molecule in the liquid.

A third type of rotational effect was discovered in the study of the vapor pressures of the deuterated ethylenes.¹³ Figure 9 shows the difference between the vapor pressures of *trans*- and *cis*-dideuteroethylene. This difference is in part due to hindered rotation, just as in the case of N_2O , but is principally due to the coupling of the rotation of the molecule about the $C=C$ axis with the out-of-plane bending of the molecule in the liquid. These

are different for the different ethylenes because of symmetry considerations.^{13,14}

APPLICATIONS TO GEOCHEMISTRY

Some 25 to 30 years ago both Dole¹⁵ and Nier¹⁶ showed that the isotopic abundances of the light elements are not constant in nature. Since that time a branch of science, isotope geology, has matured which applies these variations to the investigation of geochemical and geological processes.

The first systematic studies of the variation of sulfur isotope abundances in nature were carried out by Thode and co-workers.¹⁷ A summary of their findings, which have since been confirmed and extended by others, is shown in Figure 10. Several patterns emerge from the analyses. First the heavy isotope, ^{34}S , is more abundant in sulfates than in native sulfur or sedimentary sulfides. The spread in meteoritic samples is small and, further, the latter are almost identical with high temperature ores. The average composition of all terrestrial samples (weighted by their relative natural abundances) is close to that of meteoritic and igneous sulfur. These results indicate that meteoritic

Figure 9. The three mercury levels shown from left to right measure the vapor pressures of ordinary ethylene (I), *trans*- d_2 -ethylene (II), and *cis*- d_2 -ethylene (III). The level for *gem*- d_2 -ethylene (IV) is very similar to that for the *cis* form and is not shown. Note that the *trans* isomer is the most volatile.

and igneous sulfur have isotope abundances close to their primordial composition. The other terrestrial sources have been fractionated by chemical and physical processes. High temperature processes show little, if any, isotope discrimination. This explains the close similarity of igneous terrestrial sulfides and meteoritic sulfur. The observed terrestrial fractionation agrees with the theory of

isotope effects. The heavy isotope, ^{34}S , concentrates in the chemical species with the strongest chemical binding, namely sulfate, and is depleted where the binding is weakest, in elemental sulfur, sedimentary sulfides, and hydrogen sulfide. More detailed calculations support these qualitative arguments.

Sulfur is turned over in nature by the cycle shown in Figure 11. Note that cyclic, natural proc-

Figure 10. Isotope distribution in terrestrial and meteoritic sulfur

Figure 12. The distribution of ^{34}S in sedimentary sulfides and sulfates of different ages.

Figure 11. Sulfur cycle in nature.

Figure 13. $^{18}\text{O}/^{16}\text{O}$ ratios of calcium carbonate samples equilibrated with water at known temperatures (controlled laboratory experiments) compared to a standard reference sample.

Figure 14. Temperature at time of formation of a Jurassic belemnite as a function of radius along a line from center of growth.

esses exist for the interconversion of sulfur among the various different chemical species. Over a long time scale, the various species will reach isotope exchange equilibrium. The extent to which bacteria are involved in establishing the cycles is noteworthy. In another set of experiments Thode and co-workers¹⁷ measured the $^{32}\text{S}/^{34}\text{S}$ ratios of sedimentary sulfides and sulfates incorporated into limestones, in samples that had been geochemically dated. Their findings are shown in Figure 12. Note that samples older than 800 million years shown no difference between sulfate and sulfide. Their $^{32}\text{S}/^{34}\text{S}$ ratio, 22.1, is identical with that of igneous sulfides and meteoritic sulfur. Recent samples show ^{34}S concentration in sulfate and ^{32}S concentration in sulfide. These results strongly suggest a time scale of this order of magnitude for the onset of the turnover of natural sulfur by the biological cycle.

Early in 1947 Urey⁵ was led to predict that the temperature coefficient of the oxygen isotope exchange equilibrium between carbonate and water might serve as a thermometer. Calculations and experiments showed (cf. Table 2) the $^{18}\text{O}/^{16}\text{O}$ ratio in carbonates to be some 3% larger than in liquid water. The temperature coefficient of this isotope exchange equilibrium was calculated to be $1.8 \times 10^{-4}/^\circ\text{C}$. If carbonates exchange oxygen with water and if the equilibrium is "frozen in," then a determination of the isotope ratio with a precision of 0.2‰, or an absolute variation of the rare ^{18}O isotope by 1 part in 2.5 million, corresponds to a temperature difference of 1°C . This was the goal that Urey and his associates set for themselves. Figure 13 shows that they did, in fact, overcome numerous difficulties and establish the $\text{CO}_3^{2-}\text{-H}_2\text{O}$ exchange as a thermometer. The observed tem-

Figure 15. Analyses of cores from (top to bottom) Mediterranean, Caribbean, equatorial Atlantic, and North Atlantic. (Scales are adjusted to correspond to equal geological times.)

perature coefficient is remarkably close to the predicted value.

Figures 14 to 16 show some geochemical applications that have been made of the ^{18}O paleotemperature method. In one of the first tests of the method Urey and co-workers¹⁸ analyzed a Jurassic belemnite from the Isle of Skye. These extinct fossils are 150 million years old and have a backbone of pure calcium carbonate. Upon sectioning they show growth rings analogous to tree rings. In one such belemnite, 2.5 cm in diameter, 24 rings were analyzed and the results plotted as a function of distance from the center. The mean temperature was measured to be 17.6°C . The seasonal variation is clearly indicated in Figure 14,¹⁸ which shows a life cycle of four winters and three summers with a maximum temperature variation of 6°C . The belemnite certainly retained the temperature record.

Emiliani¹⁹ has found fossils of the tiny one-celled marine animals Foraminifera particularly useful as samples. From them he has measured ocean temperatures covering the last 200,000 years (Figure 15). The variations in the different oceans are similar but differ in magnitude. The largest variations are found for the closed body, the Mediterranean. The absolute difference between the oceans needs no further explanation. Emiliani has measured the variations of ocean temperatures

Figure 16.

during the Pleistocene and has compared them with solar radiation fluctuations calculated by Milankovitch (Figure 16). The calculations consider the change in the summer and winter solar radiation at a given latitude due to perturbations in the earth's orbit. Some schools of geochemists correlate these with the periods of the glaciers during the Pleistocene. Although the correlation shown in Figure 16 is striking, the evidence is not conclusive that glaciation is due solely to astronomical perturbations.

BIOLOGICAL EFFECTS OF DEUTERIUM

Shortly after the discovery of deuterium both Urey and G.N. Lewis recognized that the chemical and physical properties of protium and deuterium would be sufficiently different from one another to lead to marked differences in the biological activities of these isotopes. These predictions have been fully verified, and biological studies of deuterated systems have become established as important new tools in molecular biology, although the field is still in an early stage of development.

A model experiment was performed recently by Calvin, Hermans, and Scheraga.²⁰ These authors studied the effect of deuteration on the helix-random coil transition in polypeptides, proteins, and enzymes. These transformations, which are of fundamental biological importance, involve the formation and rupture of hydrogen bonds from one amino acid group in the molecule to another. If the strength of the hydrogen bonds is changed, as by the substitution of deuterium for protium, then a percentage shift can be expected in the transition temperature of the order of the percentage

Figure 17. Transition in polybenzyl glutamate.

shift in the strength of the hydrogen bonds. The formation and rupture of hydrogen bonds is just what is involved in the vaporization of water. For this process the over-all effect per hydrogen atom between protium and deuterium is 7%. This vapor pressure difference would lead to the expectation of a 21°C shift in transition temperature for a transition occurring at about room temperature. A shift of 12°C in transition temperature has been found between deuterated and ordinary polybenzyl glutamate²⁰ (cf. Figure 17). At any temperature the helical form is more stable for the deuterated compound than for the protium compound. This is a consequence of the greater stability of the deuterium bonds than of the protium bonds in hydrogen-bonded species. Similar results have been found with the enzyme ribonuclease. If enzymatic activity is dependent on the structure of the enzyme, deuteration can be expected to have profound effects on enzyme activity in some cases. The small shift in transition temperature can manifest itself in a qualitative difference in enzymatic behavior.

Superimposed on this effect of catalyst structure, differences can be expected in rates and equilibria in deuterated media compared to those in ordinary water. Mechanical properties such as viscosity are different in light and in heavy water. All these effects will lead to differences in biological activity.

The general findings have been that growth rates are retarded in D₂O compared with those in H₂O. It has been found possible to adapt lower animals from H₂O media to D₂O and vice versa. The change must be made slowly, and nevertheless the morphology is considerably altered²¹ (cf. Figure 18).

Figure 18. Phase contrast photomicrographs of blue-green algae grown in H_2O (left) and D_2O (right).

In higher animals, for instance rats, tolerance to 30% D_2O has been established. Even at this level the animals become quite sick after a month's exposure. The symptoms include considerable impairment of kidney function, anemia, disturbed metabolism, and altered adrenal function.²²

SUMMARY

With the exception of the field of chemical kinetics, a brief survey has been presented of the principles of isotope chemistry and their utility to the ever unfolding panorama of scientific research. We have come a long way since Soddy and Fajans were led to the concept of chemical twins, isotopes. We still have places to go.

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(Continued inside back cover)