The History of Oxygen

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TO A CHEMIST, OXYGEN IS ONE of the most interesting elements of the earth because of its remarkable reactivity, its paramagnetism, its participation in living processes, its use as the standard for atomic weights, and, finally, its great abundance in the earth's crust. The origin of oxygen not only as an element, but as a free gas in the earth's atmosphere, and the sequence of events which oxygen now undergoes as it passes through the carbon and other cycles, constitute the history of oxygen, a subject about which we may never know with certainty all of the details, although the broad aspects of some of its history seem at the present time to be clear.

The three stable isotopes of oxygen, masses 16, 17, and 18, presumably were initially formed from the primordial flux of neutrons (called by Alpher the "ylem") which was the precursor of all matter. Alpher, Bethe, and Gamow (1) have recently published a neutron-capture theory of the formation and relative abundance of the elements which makes plausible, from a semiquantitative point of view, the idea that all the elements were initially formed in the same event, by the decay of neutrons to protons and electrons and by the capture of neutrons and protons by each other to form the various known nuclei. Given the formation of oxygen by this mechanism, we now proceed to discuss the history of oxygen on our own planet.

A problem which challenges our imagination is that of explaining the existence of free uncombined oxygen in our atmosphere. D. ter Haar (12) has recently reviewed theories of the origin of our solar system and presented a new modified theory of his own which comes the closest of all extant theories to explaining quantitatively the known facts. He starts with the simplest possible hypothesis, originally proposed by Kant, namely, a sun surrounded by a gaseous envelope. Condensation nuclei formed near the sun will be inorganic compounds of heavy elements (that is, heavier than hydrogen and carbon); hence the inner planets will be small and dense. They will have acquired their size before gravitational capture of matter from the gaseous envelope has become important. The outer planets, however, would be formed in a relatively cold region of the envelope so that organic as

well as inorganic compounds could form and condense. The outer planets, therefore, will be of small density, but will grow to large size because of gravitational capture.
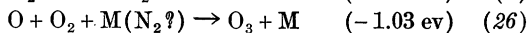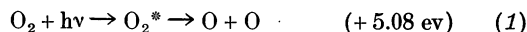
During the process of condensation, the earth seems to have lost practically all of the light elements which could not be bound chemically; for example, as Unsöld (24) has pointed out, in the star τ Scorpii, neon and nitrogen are about equally abundant, and their abundance does not fall much short of that of oxygen, whereas in the earth's atmosphere the abundance of neon is almost negligible. If neon of atomic weight 20 could not be held by the earth, neither could water vapor of mass 18, nor atomic oxygen of mass 16. Argon of mass 40 and molecular oxygen of mass 32 would be much more likely to be captured. However, because of the reactivity of oxygen it is almost inconceivable that the earth's atmosphere when first formed could have contained this element in its uncombined state. Water must have been held in combination with inorganic compounds, most of which are readily hydrated or which adsorb water to some extent. As the matter condensed to create the earth, the potential energy of the gaseous envelope was undoubtedly transformed into heat, raising the temperature of the growing nucleus. Inasmuch as the melting point probably rises more rapidly with the pressure in the interior of the earth than does the temperature, the earth would solidify first within, followed by a gradual consolidation proceeding outward (see discussion in 29). This heat would cause water to be driven out, but if the planet had accumulated enough mass at this point, the water would be held in the gravitational field. In 1924 Tammann (23) calculated that even at as high a temperature as 1,500° K the fraction of water molecules having velocities equal to or greater than the "escape velocity" was only 1.4×10^{-75}, a negligible fraction. Nitrogen and carbon dioxide would collect in the atmosphere on being driven out of the hot magma, in much the same way as water. In addition to being fixed in water and carbon dioxide, we find oxygen in the earth's crust also in the form of oxides and silicates, but we shall not be concerned here with the history of oxygen in these compounds.

We come now to the period of oxygen's history beginning with the earth after it had cooled from its original molten state. The most important fact that

confronts us is the present existence of oxygen in the atmosphere; there are those, such as the Russian biochemist Oparin (*18*), who believe that life was first created on the earth by processes not involving free oxygen, such as the slow formation in the oceans of small organic molecules, then larger polymers, then colloidal particles, such as tactoids or coacervates, and finally living organisms. These living organisms then began to photosynthesize, thus creating the oxygen of the atmosphere by the dehydrogenation of water. The difficulty with this theory of Oparin's is that it cannot be checked by the application of known physicochemical principles, because of our present ignorance concerning not only the mechanism of the origin and production of life, but also of the details of the photosynthesis process. Tammann (*23*) believed that oxygen arose through the thermal dissociation of water followed by escape of hydrogen from the atmosphere, but for this process to have been significant the temperature of the water vapor must have exceeded $1,500°$ K (to produce appreciable dissociation).

In the higher regions of the earth's atmosphere, in the stratosphere, the ultraviolet light from the sun is much more intense than at the earth's surface, producing at a height of 25 km a maximum in the partial pressure of ozone, according to the photochemical reaction

$$O_2 + hv \rightarrow O_2^* \rightarrow O + O \qquad (+5.08 \text{ ev}) \qquad (1)$$
$$O + O_2 + M(N_2?) \rightarrow O_3 + M \qquad (-1.03 \text{ ev}) \qquad (26)$$

(Reaction numbers are those of Table 1.) It is this layer of ozone which prevents practically all of the light of wave length shorter than about 3,000 A from reaching the earth. But above 50 km very little ozone is present, so that we may assume that in this region of the stratosphere and at higher levels in the ionosphere, the gas present is exposed to solar radiation of wave length even as short as 600 A. It is important to consider possible photochemical as well as ionization processes which may have given rise to free atmospheric oxygen as suggested by Poole (*19*). In Table 1 we have listed all of the processes, photochemical or ionic, which are believed to exist or to be capable of existing in the stratosphere or ionosphere. It is a relatively simple matter to write a sequence of reactions which can produce oxygen; for example, water and oxygen on dissociation produce fragments which could react as follows:

$$H_2O + hv \rightarrow H + OH \qquad +5.12 \text{ ev} \qquad (12)$$
$$O_2 + hv \rightarrow O + O \qquad +5.08 \text{ ev} \qquad (1)$$
$$O + OH \rightarrow O_2 + H \qquad -0.69 \text{ ev} \qquad (22)$$
$$OH + OH \rightarrow H_2O + O \qquad -0.61 \text{ ev} \qquad (9)$$

or

$$OH + OH \rightarrow H_2 + O_2 \qquad -0.69 \text{ ev} \qquad (14)$$

If the hydrogen of reactions 3, 17, and 19 has thermal energies greater than 1,000 or $1,500°$ K, it will be

TABLE 1

POSSIBLE PHOTOCHEMICAL AND IONIC PROCESSES OF THE STRATOSPHERE AND IONOSPHERE*

Reaction No.	Reaction	Energy absorbed (+) or liberated (−) in ev	Reference
	Endothermic Photochemical Reactions		
1.	$O_2 + hv \rightarrow O + O$	+ 5.08	(*13*)
2.	$O_2 + hv \rightarrow O \ (^1D) + O$	+ 7.05	(*10*)
3.	$H_2O + hv \rightarrow H + OH$	+ 5.12	(*13*)
4.	$N_2 + hv \rightarrow N + N \ (^2D°)$	+ 9.75	(*17*)
5.	$CO_2 + hv \rightarrow CO + O \ (^1D)$	+ 7.47	(*17*)
6.	$CO + hv \rightarrow C + O$	+ 9.14	(*13*)
	Endothermic Ionizing Reactions		
7.	$O_2 \rightarrow O_2^+ + e$	+ 12.2	(*22*)
8.	$O_2 \rightarrow O + O^+ + e$	+ 18.65	(*22*)
9.	$H_2O \rightarrow H_2O^+ + e$	+ 12.7	(*22*)
10.	$H_2O \rightarrow H + OH^+ + e$	+ 18.4	(*15*)
11.	$H_2O \rightarrow H + H + O^+ + e$	+ 23.1	(*15*)
12.	$H_2O \rightarrow H_2 + O^+ + e$	+ 18.6	(*15*)
13.	$N_2 \rightarrow N_2^+ + e$	+ 15.51	(*22*)
14.	$CO_2 \rightarrow CO_2^+ + e$	+ 14.4	(*22*)
15.	$CO_2 \rightarrow CO + O^+ + e$	+ 19.1	(*22*)
16.	$CO_2 \rightarrow CO^+ + O + e$	+ 19.6	(*22*)
	Exothermic Thermal Reactions		
17.	$OH + O \rightarrow O_2 + H$	− 0.69	(*2*)
18.	$OH + OH \rightarrow H_2O + O$	− 0.61	(*2*)
19.	$OH + OH \rightarrow H_2 + O_2$	− 0.69	(*20*)
20.	$N \ (^2D°) + H_2O \rightarrow NH + OH$	− 0.7	
21.	$N + H_2O \rightarrow NO + H_2$	− 0.31	
22.	$N + OH \rightarrow NO + H$	− 0.95	
23.	$N \ (^2D°) + OH \rightarrow NH + O$	− 1.43	
24.	$N \ (^2D°) + CO_2 \rightarrow NO + CO$	− 2.2	
25.	$N + O_2 \rightarrow NO + O$	− 0.21	
26.	$N + NH \rightarrow N_2 + H$	− 3.98	
27.	$C + H_2O \rightarrow CO + H_2$	− 4.16	
28.	$CO + H_2O \rightarrow CO_2 + H_2$	− 0.52	
	Exothermic Thermal Reactions Requiring a Third Body (Possibly N_2)		
29.	$O + O_2 \rightarrow O_3$	− 1.03	(*22*)
30.	$O + H \rightarrow OH$	− 4.34	(*8*)
31.	$OH + OH \rightarrow H_2O_2$	− 2.2	(*27*)
32.	$N + N \rightarrow N_2$	− 7.38	(*13*)
33.	$N + O \rightarrow NO$	− 5.29	(*13*)
34.	$N + H \rightarrow NH$	− 3.4	(*13*)
35.	$N + C \rightarrow CN$	− 5.96	(*13*)
36.	$C + H \rightarrow CH$	− 3.47	(*13*)
	(Also the reverse of reactions 1, 3, 5 and 6)		
	Exothermic Ionic Processes		
37.	$O_2^+ + e \rightarrow O + O$	− 7.11	
38.	$N_2^+ + e \rightarrow N + N$	− 8.12	
39.	$H_2O^+ + e \rightarrow H + OH$	− 7.6	
40.	$OH^+ + e \rightarrow H + O$	− 9.0	

* All atoms and molecules listed are gaseous and in their ground states unless otherwise indicated. All atoms and molecules are composed of the most common isotope (slight differences will exist if hydrogen, H^1, is replaced by deuterium, for example).

quickly lost, leaving an excess of free oxygen in the atmosphere. Poole (*19*) has estimated that, at a tem-
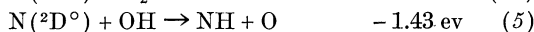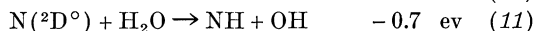
perature of 1,000° K, the ionosphere will lose all of its hydrogen in five years; if the hydrogen exists as atomic rather than molecular hydrogen, the time of escape is reduced from five years to two days. For every two atoms of hydrogen lost, the atmosphere will show a net increase of one oxygen atom gained.

It will be noted that reactions 17, 18, and 19 are all bimolecular and give rise to two products, and thus do not require a "wall" or a third body for conservation of momentum; in other words, they are homogeneous reactions. However, reactions 18 and 19 can probably be ruled out as being of minor importance because the relative amount of water present is smaller than that of oxygen. As a consequence, the partial pressure of OH will be smaller than that of atomic oxygen, and the probability of a collision between OH and O will be greater than the probability of a collision between two OH radicals. Dobson, Brewer, and Cwilong (4), who have made careful measurements of the frost point in the stratosphere up to heights of about 10 km, point out that the humidity content of the stratosphere, which drops off rapidly when the stratosphere is reached, is less than expected. The water content of the upper stratosphere and ionosphere is unknown. Elvey (9) states, however, that the spectrum of the nonpolar aurora (light of the night sky) demonstrates in the upper atmosphere the presence of water molecules in certain excited states along with oxygen, and sodium atoms, N_2, NO, and O_2 molecules.

For hydrogen atoms or molecules to have velocities equal to the escape velocity as calculated by Jeans (14) their velocities must be equivalent to temperatures of several thousands of degrees (at 1,500° K the fraction of hydrogen molecules, assuming a Maxwellian distribution, which have velocities equal to or greater than the escape velocity, is 0.00032, (23). Is there any mechanism by which hydrogen atoms or molecules can acquire such high energies? In reaction 17 which we favor as the oxygen-producing reaction, the energy liberated is 0.69 ev; if this excess energy is dissipated as translational energy the fraction 32/33, or 0.67 ev, will be carried by the hydrogen atom. This is equivalent to a "temperature" of about 5,000° K. Such a "hot" molecule will soon reach thermal equilibrium by collisions with the gases at the level of the stratosphere at which it is formed because the molecular mean free path is negligible in comparison to stratospheric distances at the pressures of 10^{-3} to 10^{-5} mm Hg existing in the stratosphere. As greater heights are reached, the pressure drops and the temperature rises. Thus, at a height of 285 km Godfrey and Price (11) estimate the temperature to be 1,350° K and the pressure, 6×10^{-8} mm Hg. At this temperature the hydrogen would be rapidly lost as mentioned above. At the present time the rate of vertical mixing between layers of the stratosphere and the ionosphere is unknown, hence it is impossible to estimate the rate at which the hydrogen escapes from the stratosphere. Certainly there is a negligible abundance of hydrogen in the upper atmosphere, in contrast to earlier ideas. Wildt (30) points out that only about 1 per cent of the helium formed by radioactive decay in the earth's sedimentary rocks is present in our atmosphere; the rest "must have escaped from the atmosphere during geological time." Hydrogen presumably would escape even more rapidly than helium.

In addition to reactions 3 and 17 as a source of oxygen, the following series of reactions could conceivably result in the presence of free atmospheric oxygen:
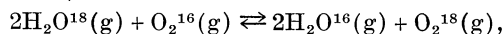
$CO_2 + h\nu \rightarrow CO + O$	+ 5.5 ev	(23)
$CO + h\nu \rightarrow C + O$	+ 9.14 ev	(18)
$N(^2D°) + H_2O \rightarrow NH + OH$	− 0.7 ev	(11)
$N(^2D°) + OH \rightarrow NH + O$	− 1.43 ev	(5)

These reactions do not require atomic oxygen for the production of O_2 as does reaction 17. Although Dhar (3) states that formaldehyde has been found in rain water, reactions producing it have not been listed in Table 1 because formaldehyde is rapidly decomposed photochemically, particularly in the presence of oxygen (17). Table 1, however, does contain an amazingly long list of reactions and demonstrates the richness of the chemical processes in the upper atmosphere. The energy for the ionization of molecules can be thought of as coming from cosmic rays; evidence for ionization is, of course, the free electron and ion content of the ionosphere.
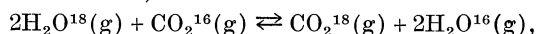
So much for the past history of the origin of oxygen, as far as we can infer it from modern knowledge. What about the present oxygen of our atmosphere, does it have a history? The universally accepted belief is that all of the oxygen of the atmosphere has been produced by the photosynthesis process in land and marine plants, mostly from photosynthesis in the oceans of the world. There is one small difficulty with this theory—the isotopic composition of atmospheric oxygen is not that of oxygen in the water of the oceans as discovered by the author (5). Ruben, Randall, Kamen, and Hyde (21) demonstrated that the oxygen produced by photosynthesis came entirely from the water and not from the carbon dioxide of the culture medium, while Dole and Jenks (6) showed that the photosynthetic oxygen had the expected isotopic composition for isotopic exchange equilibrium between oxygen and water, a composition which is definitely lower in O^{18} content than that of atmospheric oxygen. The present oxygen of the atmosphere

must have an origin different from photosynthetic oxygen or the photosynthetic oxygen must have changed its composition by isotopic exchange after being produced photosynthetically. Before we give up the century-old belief in the importance of the photosynthetic origin of oxygen, let us investigate the possibility of O^{18} enrichment of atmospheric oxygen by isotopic exchange.
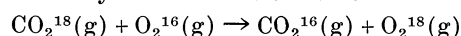
There are two possible exchange reactions which could enhance the O^{18} content of atmospheric oxygen; the first,

$$2H_2O^{18}(g) + O_2^{16}(g) \rightleftarrows 2H_2O^{16}(g) + O_2^{18}(g),$$

and the second,

$$2H_2O^{18}(g) + CO_2^{16}(g) \rightleftarrows CO_2^{18}(g) + 2H_2O^{16}(g),$$

followed by the irreversible reaction

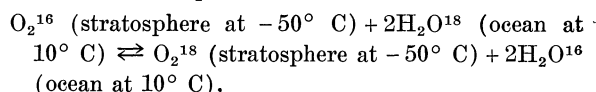$$CO_2^{18}(g) + O_2^{16}(g) \rightarrow CO_2^{16}(g) + O_2^{18}(g)$$

Urey (25) has recently recalculated, at 0° C and above, the equilibrium constants of these reactions first published by Urey and Greiff (28). Values for the fractionation factor, which is the ratio of O^{18} to O^{16} in the oxygen or carbon dioxide divided by the similar ratio for the water in the two equilibria at different temperatures, are shown in Table 2 for the oxygen-water and oxygen-carbon dioxide equilibria. We have extended Urey's calculations to −50° C in the case of the oxygen-water vapor equilibrium.

TABLE 2

FRACTIONATION FACTORS FOR THE WATER-OXYGEN AND WATER-CARBON DIOXIDE ISOTOPIC EXCHANGE EQUILIBRIA

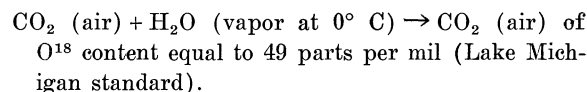| Reaction	223°K	273.1°K	298.1°K
Water vapor-oxygen	1.030	1.017	1.014
Water vapor-carbon dioxide ..		1.055	1.047
Liquid water-oxygen		1.006	1.006
Liquid water-carbon dioxide ..		1.044	1.039

Note that the fractionation factors for the oxygen-water equilibria are much smaller than those for the carbon dioxide-water equilibria and have a practically zero temperature coefficient in the case of the equilibrium between oxygen gas and liquid water. Letting α represent the fractionation factor, the percentage increase in the O^{18} abundance of the oxygen when brought into equilibrium with liquid water (assuming that both the free oxygen and the oxygen in water had the same isotopic composition before being brought into isotopic exchange equilibrium with each other and assuming an infinite volume of water) is $100(\alpha - 1) = 0.6\%$ or 6 parts per mil at 273° K. Atmospheric oxygen has an O^{18} content 30 parts per mil greater than the oxygen of Lake Michigan water. Even assuming that the exchange equilibrium is established at 221° K and that the water vapor participating in the

equilibrium at this low temperature eventually established an equilibrium with ocean water at 10° C so that the over-all process can be written

$$O_2^{16} \text{ (stratosphere at } -50° \text{ C)} + 2H_2O^{18} \text{ (ocean at } 10° \text{ C)} \rightleftarrows O_2^{18} \text{ (stratosphere at } -50° \text{ C)} + 2H_2O^{16} \text{ (ocean at } 10° \text{ C),}$$

the O^{18} content of free oxygen will be enhanced 20 parts per mil with reference to ocean water or 25 parts per mil with respect to Lake Michigan water, still less than 30 parts per mil.

Dole and Roake (7) have shown, furthermore, that oxygen isotopes do not exchange between oxygen gas and water vapor when passed through a glow discharge at low temperatures and pressures. An exchange of oxygen isotopes could be demonstrated, however, between carbon dioxide and oxygen, so that the enhanced O^{18} content of the atmosphere can be fully and quantitatively explained by the following sequence of reactions:

$$CO_2 \text{ (air)} + H_2O \text{ (vapor at } 0° \text{ C)} \rightarrow CO_2 \text{ (air) of } O^{18} \text{ content equal to 49 parts per mil (Lake Michigan standard).}$$

This CO_2, having an O^{18} content enhanced to 44 parts per mil with respect to ocean water, then undergoes an irreversible and random exchange of its oxygen isotopes with the oxygen of the stratosphere by virtue of reactions 1, 3, 5, 6, 17, and 19 of Table 1 or their reverse, producing in the stratosphere an O^{18} content enhanced to the maximum extent of 44 parts per mil with respect to ocean water. The observed value is 30 parts per mil with respect to Lake Michigan water or 25 parts per mil, to ocean water (26). The assumptions in this calculation are: (1) the irreversibility of the photochemically produced exchange of isotopes and (2) the temperature at which equilibrium between the CO_2 and H_2O is established.

Despite this success in accounting for the O^{18} content of the atmosphere by assuming a stratospheric irreversible exchange of O^{18} between O_2 and CO_2, there is still the problem of demonstrating a transfer of the O^{18} enhanced oxygen from the stratosphere to the troposphere. As Dobson, Brewer, and Cwilong have pointed out (4), a "general world circulation between the equator and the poles which causes the air in the stratosphere to rise slowly near the equator and to subside slowly near the poles" has been postulated to explain differences in temperature in the stratosphere at low and high altitudes. These authors find it unnecessary to invoke this hypothesis, preferring to base the temperature differences on the differences in radiation due to different ozone content. Thus there seems to be no proof at the present time of vertical mixing between the low and high levels of the atmosphere, although horizontal winds of high velocity

are known to exist in the stratosphere. It would be extremely interesting to investigate the isotopic composition of oxygen in the stratosphere as well as from various geographical locations such as the arctic and the tropics. It might be possible to follow large-scale air currents in this way because, from all we know now, the photosynthesis reaction of the oceans is delivering to the atmosphere oxygen having a different isotopic composition from that already present. Similarly, the stratospheric oxygen may have an isotopic composition different from the oxygen in contact with the tropical oceans. The maximum expected difference between stratospheric and photosynthetic oxygen is 33 parts per mil of which the error of measurement is estimated as 1 part per mil.

We come finally to the thought that if oxygen exchange in the stratosphere is significant, then synthesis of oxygen in the stratosphere may also be significant, and that the total abundance of free oxygen in the atmosphere may be slowly increasing. In other words, it would appear that oxygen has a future as well as a past and present history.

On the Origin of the Chemical Elements

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IN RECENT YEARS MANY ATTEMPTS have been made to explain the observed abundance of the chemical elements and their isotopes. A survey of papers on this subject will be presented in a forthcoming paper (*4*). In this article, an attempt will be made to connect these theories with the general cosmogony developed by von Weizsäcker (*9*), not by offering a complete and quantitative theory, but rather by sketching a possible way of combining ideas expressed by other authors into a theory which may, perhaps, be more satisfactory than the existing ones.[2] The limitations of earlier theories are considered in the survey paper (*4*) mentioned earlier.

The idea is to eliminate one of von Weizsäcker's initial conditions, on the one hand, and to suggest possible initial steps for the theories of van Albada (*1*), Beskow and Treffenberg (*2*), and Mayer and Teller (*7*). In von Weizsäcker's cosmogony the universe is, at the beginning of the present epoch, filled with a turbulent gas. The origin of this gas and its turbulence is assumed to be a problem lying outside the scope of this cosmogony. Von Weizsäcker

further assumes, however, that the composition of this gas is the same as the present composition of interstellar gas, or stellar material. It may be sufficient, however, to assume this gas to consist of hydrogen only, thus simplifying the initial conditions of von Weizsäcker's cosmogony. Expressed differently, the solution of the problem of the origin of the chemical elements, other than hydrogen, is thus brought into the present epoch. By the "present epoch," we shall understand the period often referred to as the age of the universe. Following present ideas about the age of the universe, we shall put it at between 10^9 and 10^{10} years (*3, 4, 8*).

As was shown by von Weizsäcker, in the gas filling the universe, concentrations of mass and finally gaseous bodies with masses of the order of stellar masses will be formed. Due to the conservation of angular momentum, these bodies will be rotating fast. In the beginning the energy of these stars will be provided by gravitational energy, and later the deuterium reaction $(H + H \rightarrow D)$ will set in. When the densities and central temperatures have become greater, the von Weizsäcker-Bethe carbon-nitrogen cycle will start. (The necessary carbon nuclei will have been formed by nuclear reactions in the way, for instance, discussed by Bethe.) When all the hydrogen has been used, the star will collapse. The development described by Hoyle (*5*) will follow. Of particular interest are those stars which have not yet lost their

and supplemental nitrogen fertilization with adapted hybrids, are ignored. Modern methods for weed control with chemical sprays are mentioned but not explained.

Apparently the authors have tried to cover too much too fast. Thus, they avoid specificity, yet specificity is essential in farm management. The result is a mixture of old concepts and comparatively old practices, some good and some mediocre. It is far behind the modern successful farmer.

Farm soils is a more useful book. First of all, the author has confined his attention largely to soil management and thus not attempted to cover so much ground. Yet, relatively, it is less up to date than the earlier editions. Too much old material has been retained and not enough of the new results and new methods explained.

The author's concept of soil management is largely one of maintaining the soil. He writes that ''farm soils tend to become acid.'' More often ''farm soils'' that need lime were acid at the start. Our problem with lime, organic matter, plant nutrients, and soil structure is, generally, as much or more one of improving the natural soils as it is of maintaining them.

The discussion of soil classification as a basis for predicting yields under alternative practices is outdated by 15 years or more. And even the system presented is not integrated with the recommendations in a way to permit these to be applied to specific conditions.

The moldboard plow is defended almost too much. Under dry-farming in regions of low rainfall he writes: ''Therefore, plow to a depth of 7 or 8 inches promptly after crops are harvested.'' The greatly improved tillage methods now used by good farmers for the Chernozem, Chestnut, and Brown soils are inadequately explained.

This book is much better than *Crop management and soil conservation* on lime and fertilizer. Worthen stresses the savings to be made from home mixing of fertilizers and explains how it is done. He emphasizes the great savings resulting from the use of high-analysis materials rather than low-analysis ones. Yet, one wishes he had gone further and not written that ''the 5-10-5, with a total of 20 percent of nitrogen, phosphoric acid, and potash, would class as a high-analysis or a high-grade fertilizer. . . .'' But he shows the advantage of really high-analysis kinds. There are little errors such as ''complete'' for ''mixed'' fertilizer and ''alkali'' for ''saline'' soil. Generalizations are used where published data, if reviewed, would have permitted specificity.

Like the earlier editions, this book has a place for vocational agricultural courses in high schools and junior colleges, particularly in the humid-temperate part of the country. But it will need careful correction and supplementing by the teacher, using local soil surveys and recent bulletins of state agricultural colleges and the U. S. Department of Agriculture.

Two books are badly needed—one in agronomy and one in soil management—that can compare with *Farm management* by Black, *et al.* and *Feeds and feeding* by Morrison. Neither of these comes close to filling this need.

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