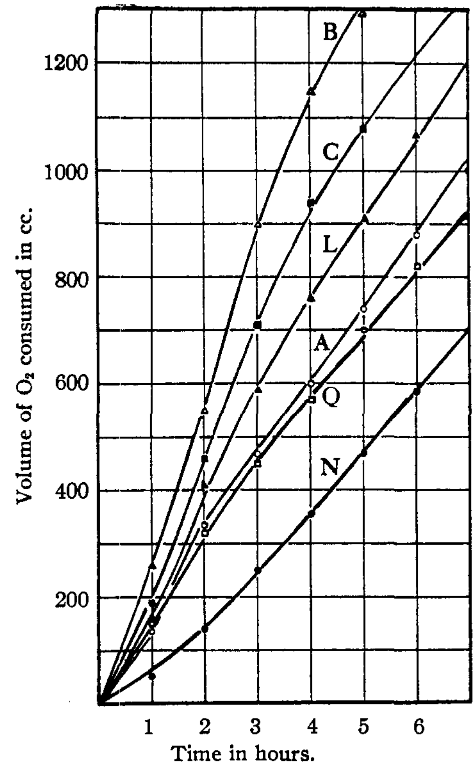


Fig. 1.—Effect of the addition of 0.001 *M* quantities of oxidation catalysts on the oxygen uptake in the oxidative condensation of 2-methoxy-4-nitrotoluene: ●, no catalyst, N; □, anthraquinone, Q; ○, acetone, A; ▲, cyclohexanol, L; ■, cyclohexanone, C; △, benzoquinone, B.

hexanol may be oxidized first to cyclohexanone which then acts as a catalyst. The total oxygen consumption was well above the expected theoretical amount. Therefore, oxygen was also used up in some side reaction.

Atmospheric oxygen as in the original method[1] was then tried in the presence of a catalyst and it proved to be effective. When benzoquinone was used and air bubbled through the reaction mixture for seven hours a 55% yield of the bibenzyl was obtained. When acetone was used and air bubbled through for twenty-two hours, 60% yield of a mixture of two parts of the bibenzyl and one part of the stilbene was obtained.

Experimental

Preparation of 2,2'-Dimethoxy-4,4'-dinitrobibenzyl.—Air is bubbled through a suspension of 10 g. of pure 2-methoxy-4-nitrotoluene in 200 ml. of 33% solution of methanolic potassium hydroxide containing one gram of benzoquinone for seven hours. The product is filtered, washed with dilute hydrochloric acid then with hot methanol and recrystallized from ethyl acetate, m. p. 179°; yield 5.5 g. On bromination it gives α,α'-dibromo-2,2'-dimethoxy-4,4'-dinitrobibenzyl, m. p. 247–249°.

Preparation of 2,2'-Dimethoxy-4,4'-dinitrobibenzyl and 2,2'-Dimethoxy-4,4'-dinitrostilbene.—Forty grams of 2-methoxy-4-nitrotoluene is placed in 800 ml. of 33% methanolic potassium hydroxide and 20 ml. of acetone. The mixture is well stirred and air is bubbled through at room temperature. After twenty-two hours the precipitate is filtered off. The mixture is then taken up in boiling ethyl acetate which dissolves the bibenzyl. The stilbene is filtered off and weighs 7.5 g. On cooling 16 g. of bibenzyl is collected from the ethyl acetate. The stilbene on bromination gives the same α,α'-dibromo-2,2'-dimethoxy-4,4'-dinitrobibenzyl as was obtained by brominating the bibenzyl.

Procedure Followed to Determine the Rate of Oxygen Consumption (see Fig. 1).—Ten grams of 2-methoxy-4-nitrotoluene and 0.001 *M* quantities of the different oxidation catalysts were placed in 200 ml. of 33% methanolic potassium hydroxide and shaken for twenty-two hours in the Adams shaker in the presence of oxygen at atmospheric pressure. The yield of mixture of the bibenzyl and the stilbene for all the catalysts tried varied between 30% and 38%. The mixture contained about equal quantities of the two substances.

RESEARCH LABORATORIES
SCHIEFFELIN & CO.
NEW YORK 3, N. Y.

RECEIVED SEPTEMBER 5, 1947

COMMUNICATIONS TO THE EDITOR

STUDIES ON THE MECHANISM OF THE DEAMINATION OF DIAZONIUM SALTS WITH HYPOPHOSPHOROUS ACID

Sir:

In an attempt to introduce a deuterium atom into the *meta* position of nitrobenzene, a deamination reaction with hypophosphorous acid[1] was carried out. Thus, to a solution of diazotized *m*-nitroaniline containing 30% deuterium oxide was added a solution of hypophosphorous acid in water which contained 50% of deuterium oxide. Before use the hypophosphorous acid solution was allowed to stand at room temperature for twenty-four hours. Much to our surprise *the nitrobenzene which was obtained was found to contain no deuterium.* This was determined for us by Mrs. James L. Johnson and Dr. Foil A. Miller with an infrared spectrometer.[2] Similarly, aniline hydrochloride prepared from this sample of nitrobenzene was found to show no absorption in the region of the C–D stretching frequencies (2270 cm.$^{-1}$).

We are not yet in a position to comment on the mechanism of the reaction, but this result is par-

---

(1) See Kornblum, "Organic Reactions," Vol. II, John Wiley and Sons, Inc., New York, N. Y., 1944, p. 277.

(2) The lower limit of the sensitivity of this method for the detection of deuterium in an aromatic ring has not yet been completely determined. From dilution experiments conducted with deuterobenzene in benzene, however, it appears that if 1–2% of the nitrobenzene molecules contained deuterium, it could be detected.

ticularly interesting in view of the fact that the work of Erlenmeyer, Schoenauer and Schwarzenbach[3] indicates that all three hydrogen atoms of hypophosphorus acid exchange rapidly with deuterium oxide. Apparently in the deamination reaction a hydrogen atom was extracted from another aromatic nucleus, since that is the only source of hydrogen which is not equilibrated with deuterium.

---

[3] Erlenmeyer, Schoenauer and Schwarzenbach, *Helv. Chim. Acta*, **20**, 732 (1937); see also Franke and Mönch, *Ann*, **550**, 1 (1941).

DEPARTMENT OF CHEMISTRY
UNIVERSITY OF ILLINOIS
URBANA, ILLINOIS
ELLIOT R. ALEXANDER
ROBERT E. BURGE

RECEIVED DECEMBER 22, 1947

CAUSE OF EXPLOSIONS OCCASIONALLY OBSERVED DURING EVAPORATION OF SOLUTIONS OF ALUMINUM HYDRIDE AND RELATED COMPOUNDS

Sir:

In experiments involving dimethyl ether solutions of aluminum hydride, as well as of lithium and sodium aluminum hydrides, explosions have occasionally occurred. The details to be described have demonstrated that the explosions were caused by carbon dioxide present as an impurity in some samples of the ether.

Explosions have occurred toward the end of distillations undertaken to remove the solvent from dimethyl ether solutions of lithium aluminum hydride. That the explosions were due to an impurity more volatile than dimethyl ether was indicated by the facts (1) that later samples of the ether taken from a cylinder caused less violent explosions than the first few samples, and (2) ether taken from another cylinder at no time caused these explosions. These facts, as well as chemical tests to prove their absence, excluded aldehydes, alcohols or peroxide as the offending impurity. Considerable carbon dioxide was present in the cylinders whose contents led to explosions; its removal by fractionation *in vacuo* rendered the contents harmless. After carbon dioxide was reintroduced into the purified sample, the explosions recurred. It was then found that diethyl ether solutions of lithium aluminum hydride, which can ordinarily be safely evaporated, may also cause explosions if first treated with carbon dioxide in considerable amount.

Evaporation of dimethyl ether solutions of aluminum hydride and subsequent slow heating of the residue may result in a very rapid reaction in which, at slightly above 40°, relatively large quantities of gaseous material are suddenly released, but which is not accompanied by detonation, unless a large excess of aluminum chloride is present. But carbon dioxide-free ether leads to a residue which begins to decompose *slowly* above 70°, if aluminum chloride is not present in large amount; even in the presence of a large excess of the latter (2.5 moles/l.), the decomposition above 40° is sudden but without detonation. Other experiments have shown that aluminum chloride accelerates the decomposition of aluminum hydride.

On one occasion a violent explosion occurred when the residue resulting from evaporation of a dimethylcellosolve solution of aluminum hydride was warmed. The aluminum hydride was contaminated with aluminum chloride, which is known to decrease the stability of the former, and the dimethylcellosolve was impure. No explosion resulted when the experiment was repeated with purified materials.

Lithium aluminum hydride is considerably more stable than aluminum hydride. Approximately two hundred different reactions using the former reagent have been carried out in this Laboratory without untoward results although no precautions were taken to exclude carbon dioxide. We are, therefore, convinced that it may be used with perfect safety if the precautions suggested by the facts herein reported are followed. In other words, we recommend that in the reduction of organic compounds the normal procedure of hydrolyzing the initial reaction product before evaporation of the solvent be employed. If, as would rarely be the case, it is desired to evaporate the solvent before hydrolysis of the initially produced organic salts, the safety of the procedure should first be tested by use of a small sample.

DEPARTMENT OF CHEMISTRY
THE UNIVERSITY OF CHICAGO
CHICAGO 37, ILL.
GERALDINE BARBARAS
GLEN D. BARBARAS
A. E. FINHOLT
H. I. SCHLESINGER

RECEIVED DECEMBER 15, 1947

CRYSTALLINITY OF HYDRO-CELLULOSES

Sir:

Hydrolytic methods are being used at present for evaluating the intramolecular structure of cellulose. It has been suggested in a paper to be published in THIS JOURNAL that such methods might possibly cause additional crystallization in the initial stages of hydrolysis. We have succeeded in obtaining data which indicate that such changes do occur. Consequently, it would seem necessary to reinterpret accessibility measurements based on hydrolysis rates in terms of two competing processes: hydrolysis and further crystallization.

The data in Table I were obtained on a beechwood pulp that had been digested in boiling 2.5 *N* hydrochloric acid–0.6 *M* ferric chloride solution constantly saturated with air.

TABLE I

Time treated, minutes	Percentage of samples destroyed	Specific vol. of bone-dry hydrocellulose, ml./g.	Percentage increase in crystallinity from Sp *V*. change, %
0	0	0.652	
10	4.8	.647	15
19	6.0	.646	18
28	39.0	.645	19

The specific volume was obtained from density, which was measured by flotation in carbon tetrachloride. The percentage increase in crystallinity was calculated using the values given by H. Mark[1] for the density of crystalline cellulose, 1.59 g./ml. (Sp V = 0.629). The density of amorphous cellulose was taken to be 1.50 g./ml. (Sp. V = 0.667) as estimated by P. H. Hermans.[2] The following relation was used to calculate the increase in crystallinity

$$\frac{\text{Sp. } V \text{ of untreated sample} - \text{Sp. } V \text{ of treated sample}}{\text{Difference in Sp. } V \text{ of amorphous and crystalline cellulose}}$$

A rayon of specific volume 0.653 showed a weight loss of 10% when treated with the reagent for five minutes. The specific volume of the treated sample was 0.647, indicating an increase of 16% in the crystallinity.

If the only process involved in the initial stages of this treatment is the attack and removal of the amorphous portion of the fiber, the density increase should predict a change of crystallinity of the same order of magnitude as the weight of material lost. However, the actual increase in density is much larger. This may indicate that with the rupture of a cellulose chain in an amorphous portion of the fiber a process of crystallization is initiated.

[1] H. Mark, "Physik und Chemie der Cellulose," Berlin, 1932.

[2] P. H. Hermans, "Contribution to the Physics of Cellulose Fibres," Elsevier Publishing Co., Inc., 1946.

INST. OF POLYMER RESEARCH
POLYTECHNIC INST. OF BROOKLYN
BROOKLYN, N. Y.

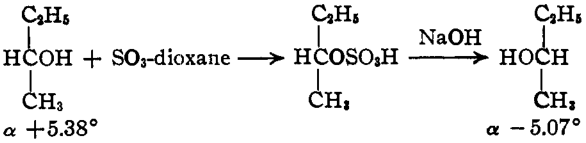
F. C. BRENNER
V. FRILETTE
H. MARK

RECEIVED DECEMBER 9, 1947

A NEW WALDEN INVERSION

Sir:

The following reaction sequence constitutes a new Walden inversion



$$\underset{\alpha\ +5.38^\circ}{\mathrm{H\overset{C_2H_5}{\underset{CH_3}{C}}OH}} + \mathrm{SO_3\text{-}dioxane} \longrightarrow \mathrm{H\overset{C_2H_5}{\underset{CH_3}{C}}OSO_3H} \xrightarrow{\mathrm{NaOH}} \underset{\alpha\ -5.07^\circ}{\mathrm{HO\overset{C_2H_5}{\underset{CH_3}{C}}H}}$$

The second step resembles the displacement of *p*-toluenesulfonate ion from secondary alkyl esters of the sulfonic acid by ethoxide or acetate ions, a reaction which has been shown to invert the configuration of the carbinol carbon.[1]

We previously[2] prepared (+)*s*-butylsulfuric acid by action of Suter's sulfur trioxide–dioxane reagent on (+)*s*-butyl alcohol and found that upon its acid hydrolysis a (+)alcohol was recovered whose rotation was 30% of that of the starting material. We had assumed that the racemization was confined to the hydrolytic step and that the formation of the alkylsulfuric acid proceeded with little if any racemization. This is now confirmed.

Since the action of chlorosulfonic acid and of sulfuric acid upon (+)alcohol gave (+)*s*-butylsulfuric acid, and since it was difficult to see how all these methods could involve breaking the carbon–oxygen bond of the alcohol, it was considered that the alkylsulfuric acid had the same configuration as the alcohol. If this is true, and its plausibility is increased by our finding that the sulfur trioxide–pyridine complex[3] also gives a (+)alkylsulfuric acid, then the saponification involves the displacement of a sulfate ion by a hydroxide ion with inversion of configuration. Since this reaction is one between two ions of like charge it is comparatively slow.

The alkylsulfuric acid was prepared by action of the sulfur trioxide–dioxane complex upon an alcohol of α +5.38. In one saponification, 8 g. of (+)sodium *s*-butylsulfate (from neutralizing the alkylsulfuric acid reaction mixture with sodium hydroxide, evaporating and extracting the sodium alkylsulfate with methanol) and 10 g. of sodium hydroxide were dissolved in 50 cc. of water. At 100° the reaction required two days for substantial completion. Alcohol was recovered in 54% yield with a rotation 6% below that of the starting alcohol and of opposite sign. About 8% of gas, apparently butylene, was evolved. A similar alcohol was obtained from the barium salt in a more concentrated potassium hydroxide solution, but several times as much butylene resulted.

Further investigation now under way at this laboratory should reveal the degree to which this reaction is common to secondary alcohols.

[1] This work of Kenyon, Phillips and co-workers is reviewed by Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., New York, N. Y., 1940, pp. 160–163.

[2] Burwell, THIS JOURNAL, **67**, 220 (1945).

[3] Sobel, Drekter and Natelson, *J. Biol. Chem.*, **115**, 381 (1936).

DEPARTMENT OF CHEMISTRY
NORTHWESTERN UNIVERSITY
EVANSTON, ILL.

ROBERT L. BURWELL, JR.
HOWARD E. HOLMQUIST

RECEIVED DECEMBER 18, 1947

FORMYLFOLIC ACID, A FUNCTIONAL DERIVATIVE OF FOLIC ACID

Sir:

Previous studies from this Laboratory have indicated that a coenzyme containing *p*-aminobenzoic acid is involved in combining a single carbon unit into the pyrimidine ring of purines[1] and that folic acid functions in the biosynthesis of purines.[2] Seeking functional derivatives which could act as "carriers" of formate, we prepared *p*-aminobenzoylhistidine and condensed it with α,β-dibromopropionaldehyde and 2,4,5-triamino-6-hydroxypyrimidine to obtain pteroylhistidine. No pronounced activity was obtained with either of these histidine derivatives. The announcement of the structure of rhizopterin[3] which is *p*-[N-(2-amino-4-hydroxypyramido-[4,5-b]pyrazin-6-ylmethyl)-formamido]-benzoic acid gave a clue as

[1] Shive, *et al.*, THIS JOURNAL, **69**, 725 (1947).

[2] Rogers and Shive, *J. Biol. Chem.*, in press.

[3] Wolf, *et al.*, THIS JOURNAL, **69**, 2753 (1947).

to how formate may be carried by a functional derivative of folic acid. Accordingly, formylfolic acid was prepared by heating 2 cc. of 98% formic acid and 0.5 cc. of acetic anhydride with 26 mg. of folic acid (pteroylglutamic acid) for one hour at 100°. The volatile reactants were removed *in vacuo*, and the light yellow product was precipitated from an alkaline solution by acetic acid. For analysis, a sample was hydrolyzed, and formic acid determined (Calcd.: 9.8%. Found.: 9.5%). The formylfolic acid was as effective as folic acid in stimulating the growth of *Streptococcus faecalis R* and *Lactobacillus casei*.

However, in an inhibition analysis with 7-methylfolic acid,[4] formylfolic acid was approximately thirty times as active as folic acid in preventing the toxicity of the inhibitor for *S. faecalis R;* the antibacterial index was 3,000 for formylfolic acid as compared with 100 for folic acid over a range of inhibitor concentrations from 10 to 100γ per 10 cc. In 10 cc. of medium,[2] a mixture of 0.3 γ of folic acid with 0.03 γ of formylfolic acid was no more effective than 0.03 γ of formylfolic acid alone in preventing the toxicity of the inhibitor. Synthetic rhizopterin[3] was two to three times as effective as folic acid in preventing the toxicity of the inhibitor. The toxicity of methylfolic acid was increased several fold by treating with 98% formic acid at 100° for two hours. For *L. casei* formylfolic acid was only slightly more effective than folic acid in preventing the toxicity of the inhibitor.

These results indicate that formylfolic acid is a functional derivative of folic acid and is competitively inhibited by methylfolic acid. The activity of rhizopterin indicates that it probably is converted directly to formylfolic acid.

These and other unpublished experiments offer additional evidence that the biochemical functions of *p*-aminobenzoic and folic acid derivatives involve the introduction of the single carbon unit into purines, pyrimidines and probably histidine.

The possibility of enhanced activity and decreased toxicity for formylfolic acid as compared with folic acid in treatment of pernicious anemia indicates the desirability for clinical testing.

[4] Crude product from the condensation of α,β-dibromobutyraldehyde, 2,4,5-triamino-6-hydroxypyrimidine and *p*-aminobenzoylglutamic acid obtained from Dr. E. L. R. Stokstad (Franklin, *et al.*, *J. Biol. Chem.*, **169**, 427 (1947)).

THE BIOCHEMICAL INSTITUTE AND
THE DEPARTMENT OF CHEMISTRY
THE UNIVERSITY OF TEXAS, AND
THE CLAYTON FOUNDATION FOR
RESEARCH, AUSTIN, TEXAS

MALCOLM GORDON
JOANNE MACOW RAVEL
ROBERT E. EAKIN
WILLIAM SHIVE

RECEIVED DECEMBER 29, 1947

CONCERNING THE STRUCTURE OF PHTHIOIC ACID

Sir:

Polgar and Robinson[1] have recently proposed, partly on the basis of the highly questionable degradative evidence of Wagner-Jauregg,[2] that phthioic acid,[3] isolated from tubercle bacillus, may be 3,13,19-trimethyltricosanoic acid. This acid was synthesized, but since it contains three asymmetric carbons the *dl*-mixture was of no value for comparison with the natural product. By use of the data which have now been collected concerning the optical rotation of fatty acids containing branching methyl groups, it is possible to estimate with some accuracy the maximum rotation to be expected of an acid such as 3,13,19-trimethyltricosanoic acid. Data of interest to this discussion appear in Table I.

TABLE I

Acid	$[\alpha]_D$	$[M]_D$
(+)-2-Methylhexacosanoic[4]	6.8°[a]	28°
(+)-2-Methyldodecanoic[4]	13.2°	28.3°
(+)-4-Methylhexanoic[5]	8.4°	11.0°
(+)-12-Methyltetradecanoic[6]	4.7°	11.4°
(+)-10-Methyloctadecanoic[7]	0.06°	0.2°
(+)-3-Methylhendecanoic[7]	5.0°	10.1°

[a] Taken in chloroform solution, remaining values obtained homogeneous.

As shown by the above data and elsewhere,[6] the molecular rotation remains reasonably constant for a given relationship of methyl group to the ends of the chain. Variables such as solvent and temperature have appreciable but not large effects in these compounds. For the case of 3,13,19-trimethyltricosanoic acid, it seems safe to assume that (a) the contribution of each asymmetric carbon to the rotation is unaffected by the presence of the other branching groups, (b) the contribution of carbon-19 is less than the rotation of the third or fourth acids listed in Table I, (c) the contribution of carbon-13 is similar to that of the fifth acid in Table I, (d) the contribution of carbon-3 is the same as the rotation of the sixth acid in Table I. It follows that the maximum molecular rotation for 3,13,19-trimethyltricosanoic acid should be approximately 22°, less than half the value of 49.7° observed for phthioic acid. Thus, it seems impossible that phthioic acid could have the structure suggested by Polgar and Robinson.

It is of interest that 2,3,21-trimethyltricosanoic acid has a calculated maximum molecular rotation of approximately 49.5° (ignoring any effect caused by the proximity of the 2- and 3-methyl groups), and all other trimethyltricosanoic acids would have a lower rotation. Thus, if phthioic acid is a long-chain acid with several branching methyl groups (evidence supporting this proposition seems doubtful), three of these groups are probably in the 2-, 3- and anteiso positions. Additional methyl groups between the 3- and anteiso positions should contribute little to the rotation.

[1] Polgar and Robinson, *J. Chem. Soc.*, 389 (1945).

[2] Wagner-Jauregg, *Z. Physiol. Chem.*, **247**, 135 (1937).

[3] Spielman and Anderson, *J. Biol. Chem.*, **112**, 759 (1936).

[4] Ställberg-Stenhagen, *Arkiv Kemi, Mineral. Geol.*, **23**, No. 15 (1946).

[5] Welt, *Compt. rend.*, **119**, 855 (1894).

[6] Weltkamp, THIS JOURNAL, **67**, 447 (1945).

[7] Prout, Cason and Ingersoll, *ibid.*, **69**, 1233 (1947).

It would appear, however, that the absence of rings and quaternary carbons in phthioic acid has hardly been demonstrated.

CHEMICAL LABORATORY
UNIVERSITY OF CALIFORNIA
BERKELEY, CALIFORNIA JAMES CASON
FURMAN CHEMICAL LABORATORY
VANDERBILT UNIVERSITY
NASHVILLE, TENN. FRANKLIN S. PROUT

RECEIVED OCTOBER 14, 1947

THE EXCHANGE REACTION BETWEEN THE TWO OXIDATION STATES OF THALLIUM IN PERCHLORIC ACID SOLUTIONS

Sir:

We have studied the exchange reaction between thallous and thallic ions in aqueous perchloric acid and wish to make a preliminary report of the results. This reaction is of interest in that it involves transfer of two electrons between the reacting species and in that it is found to proceed at a slow and measurable rate, in contrast to the several other electron-transfer exchange reactions which have been reported to be fast.[1] Earlier studies of this reaction[2,3] were handicapped by the short half-life of the radioactive tracer used (ThC$''$, half-life 3.1 min.) and gave results which are difficult to interpret.

Using Tl^{204,206} (half-life *ca.* 3 years), we have observed the exchange reaction and have investigated the dependence of the rate of exchange on the concentrations of thallous and thallic perchlorates, on the acid concentration, and on the temperature. The thallous perchlorate concentration was varied from 0.003 f. to 0.015 f., that of thallic perchlorate from 0.0006 f. to 0.003 f.

The method employed was to mix stock solutions of active thallic perchlorate and inactive thallous perchlorate in a stoppered flask immersed in a constant temperature bath, to withdraw samples at intervals, and to determine the distribution of activity between the two oxidation states of thallium. Thallous chromate was precipitated from the sample by the addition of a mixture containing chromate, cyanide, ethanol and excess ammonia. This method was found to give a reasonably clean separation, and not to induce a significant amount of exchange during precipitation.

The activity in the thallous fraction was found to vary with time in the simple exponential manner expected for an exchange reaction occurring at chemical equilibrium.[4,5] The rate of exchange was found to be proportional to the first power of the concentration of each reactant. The specific rate constant is 2.0 moles^{-1} liter hours^{-1} at 49.5° (perchloric acid 0.4 f.). The experimental activation energy is 12 kcal./mole. The addition of a neutral salt (LiClO$_4$ 0.6 f.) increased the rate, suggesting that ions of like sign participate in the exchange.

When the acid concentration was varied, the rate constant showed a maximum at about 1 f. (k = 2.5) and decreased to 0.9 at 3.9 f. This behavior doubtless results from the combined effects of ionic strength and hydrolysis of thallic ion.

We have also observed the exchange in 0.2 f. hydrochloric acid and find that the rate is markedly greater than in perchloric acid.

We are continuing these studies.

We are grateful to the Los Alamos Laboratory, in particular to Mr. J. W. Starner and Mr. E. L. Bentzen, for neutron irradiation of the thallium. We also are indebted to Dr. Norman Davidson for his interest and helpful advice.

GATES AND CRELLIN LABORATORIES OF CHEMISTRY
CALIFORNIA INSTITUTE OF TECHNOLOGY
PASADENA, CALIFORNIA GARMAN HARBOTTLE[6]
R. W. DODSON[6]

RECEIVED DECEMBER 8, 1947

---

[1] See for example G. T. Seaborg, *Chem. Rev.*, **27**, 199 (1940).

[2] J. Zirkler, *Z. Physik*, **99**, 669 (1936), *et al.*

[3] V. Majer, *Z. physik Chem.*, **A179**, 51 (1937).

[4] H. A. C. McKay, *Nature*, **142**, 997 (1938).

[5] R. B. Duffield and M. Calvin, THIS JOURNAL, **68**, 557 (1946).

[6] Now at Chemistry Department, Columbia University, New York, N. Y.

EXCHANGE REACTION BETWEEN THALLIUM (I) AND THALLIUM (III) IONS IN PERCHLORIC AND NITRIC ACID SOLUTIONS

Sir:

We have measured the rate of the exchange reaction between thallium(I) and thallium(III) ions in aqueous solutions of perchloric and nitric acid and have found it to be slow and measurable. The data from earlier work[1,2] on this reaction are difficult to interpret because the short lived tracer, ThC$''$ (3.1 m.), limited the duration of the experiments.

Part of the Tl^{204,206} (*ca.* 3.5 y) used as tracer in our experiments was prepared by the Tl(d, p) reaction in the Washington University cyclotron, the rest by the Tl(n,γ) reaction in the Oak Ridge pile. Our procedure was to mix acid solutions of thallium (I) perchlorate (nitrate) and active thallium (III) perchlorate (nitrate), remove aliquots at definite intervals of time, separate the two oxidation states, and assay and count the two fractions. Two methods of separation were used. (1) Thallium (III) hydroxide was precipitated with ammonium hydroxide; both fractions being subsequently weighed and counted as thallium (I) chromate. (2) Thallium (I) bromide was precipitated with sodium bromide solution, both fractions being weighed and counted as thallium (I) bromide.

A fast, incomplete, but reproducible exchange was induced at the time of separation. The induced exchange (exchange measured at zero time) could be varied from 45 to 70% for the hydroxide separation and from 8 to 13% for the bromide separation by adding the reagents in a different

---

[1] J. Zirkler, *Z. Physik*, **87**, 410 (1934); **98**, 75 (1935); **99**, 669 (1936); *Z. physik Chem.*, **A187**, 103 (1940).

[2] V. Majer, *ibid.*, **A179**, 51 (1937).

order or at different rates. However, the induced exchange was reproducible when conditions were held constant so we were able to correct for it using the equation

$$\% \text{ exchange} = \frac{\% \text{ exchange (measured)} - \% \text{ exchange (induced)}}{100 - \% \text{ exchange (induced)}} (100)$$

The corrected values (always three or more excluding the value at zero time) obeyed the exponential exchange law.[3,4] The half-times for the exchange rates are summarized in Table I. As expected the exchange rate is not dependent on the method of separation when proper account is taken of the induced exchange.

TABLE I

Tl(I)-Tl(III) EXCHANGE RATES

0.0244 f. Tl(I), 0.0244 f. Tl(III)

Acid	Temperature	Method of separation	Exchange, half-time hr.
1.0 f. HNO_3	*ca.* 25°C.	Bromide	2.5 ± 0.2
1.5 f. HNO_3	24.8 ± 0.2°	Bromide	1.8 ± 0.
1.5 f. HNO_3	24.8 ± 0.2°	Hydroxide	1.6 ± 0.2
1.5 f. $HClO_4$	24.8 ± 0.2°	Hydroxide	36 ± 4
1.5 f. $HClO_4$	24.8 ± 0.2°	Bromide	35 ± 4
1.5 f. $HClO_4$	24.8 ± 0.2°	Bromide	33 ± 4
2.5 f. $HClO_4$	24.8 ± 0.2°	Bromide	45 ± 4
3.5 f. $HClO_4$	24.8 ± 0.2°	Bromide	67 ± 5

We are extending this work to determine the effect of temperature, ionic strength, and concentrations of the reactants on the exchange rate.

[3] H. A. C. McKay, *Nature*, **142**, 997 (1938).

[4] R. B. Duffield and M. Calvin, THIS JOURNAL, **68**, 557 (1946).

DEPARTMENT OF CHEMISTRY
WASHINGTON UNIVERSITY
ST. LOUIS, MISSOURI
RENÉ J. PRESTWOOD
ARTHUR C. WAHL

RECEIVED DECEMBER 10, 1947

THE STRUCTURE OF THE DECABORANE MOLECULE

Sir:

We are studying the structure of crystalline decaborane, $B_{10}H_{14}$, by single crystal X-ray diffraction methods. We have established the approximate positions of the ten boron atoms and four of the hydrogen atoms, and have assigned probable positions to the remaining ten hydrogen atoms. (Hydrogen atoms are well resolved in fourier sections.)

The $B_{10}H_{14}$ molecule has the symmetry C_{2v}–mm2. The bond distances are as follows (see figure): B_1–B_1', B_1–B_4, B_2–B_3, B_2–B_4, B_3–B_4, are all 1.74 ±0.04 kX; B_1–B_2 and B_4–B_4' are 1.96 ± 0.04 kX; B_4–H_4 is 1.34 ± 0.04 kX,[1] and all other B–H distances are assumed the same, except B_4–H_6 which is assumed to be 1.54 ± 0.04 kX. (B_4–B_4''' and B_4'–B_4'' are 2.76 ± 0.04 kX and are not bond distances.) Each hydrogen atom, except H_6 and H_6' is bound to a single boron atom; H_6 and H_6' are each bound to two boron atoms. Each boron atom has three boron neighbors at 1.74 ± 0.04 kX and one hydrogen neighbor at 1.34 ± 0.04 kX. In addition, B_4, B_4', B_4'' and B_4''' each has a boron neighbor at 1.96 ± 0.04 kX and another hydrogen neighbor at 1.54 ± 0.04 kX; B_1, B_1', B_2, B_2', each has two boron neighbors at 1.96 ± 0.04 kX; B_3 and B_3', each has another hydrogen neighbor at 1.34 ± 0.04 kX.

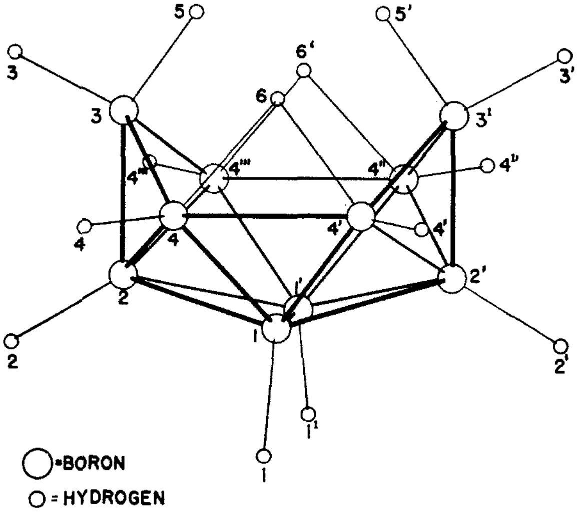


Fig. 1.

Each boron atom is bound to five or six other atoms, but the bonds are not all equivalent. Inasmuch as a bond distance of 1.96 kX has about half the "bond number"[2] of a bond distance of 1.74 kX, one can say that each boron forms five bonds of bond number 0.60. The corresponding radius, R(0.60) = 0.87 kX. Consequently, R(1) = 0.80 kX in agreement with Pauling.[2]

This structure for $B_{10}H_{14}$ gives excellent agreement with the observed X-ray diffraction intensities and also with the electron diffraction observations of S. Bauer.[3]

A detailed discussion of the determination of the structure of crystalline decaborane will be published soon.

[1] H_4, H_4', H_4'' and H_4''' were located on an electron density map; the positions of the other hydrogen atoms are assumed.

[2] L. Pauling, THIS JOURNAL, **69**, 542 (1947).

[3] S. Bauer, *ibid.*, **70**, 115 (1948).

RESEARCH LABORATORY
GENERAL ELECTRIC COMPANY
SCHENECTADY, NEW YORK.
JOHN S. KASPER
C. M. LUCHT
DAVID HARKER

RECEIVED JANUARY 21, 1948

HETEROGENEITY OF CRYSTALLINE BETA-LACTOGLOBULIN

Sir:

That crystalline β-lactoglobulin is not a homogeneous protein was indicated by the solubility measurements of Grönwall[1] and by the electrophoretic results of Li.[2] Our experiments with

[1] Grönwall, *Compt. rend. trav. lab. Carlsberg*, **24**, no. 8–11, 185 (1942).

[2] Li, THIS JOURNAL, **68**, 2746 (1946).

this protein confirm these reports and indicate a relationship between the solubility of the protein and its composition as defined by the electrophoretic patterns obtained at *p*H 4.7.

β-Lactoglobulin was prepared by the method of Palmer[3] which involves the fractionation of milk whey with ammonium sulfate after the removal of casein with acid. The isolated β-lactoglobulin was recrystallized four times by dialyzing away the salt from sodium chloride solutions at the isoelectric point. Electrophoretic studies on a 1% solution of this material in acetate buffer of ionic strength 0.1, *p*H 4.8, showed the presence of a two component system. The same preparation was electrophoretically homogeneous on the alkaline side of the isoelectric point.

Fractionation experiments on the whey proteins from the filtrate after the partial removal of crystalline β-lactoglobulin gave crystalline fractions with a mobility comparable to that of the usual β-lactoglobulin at *p*H 8.3. Electrophoresis *p*H 4.8, however, indicated a variation in the concentration of the components in each of the fractions. These new crystalline fractions also varied from the standard preparation in that they showed a greater solubility in water and in dilute salt solutions.

Crystallization of the standard preparation by dialysis from an acetate buffer solution, varied with respect to *p*H, gave fractions with a partial separation of the components as indicated electrophoretically. Partial separation of β-lactoglobulin has also been obtained by fractionation with alcohol at low temperatures. The fractions obtained by alcohol have properties in agreement with the data reported in the table below for preparations obtained by the other methods. These results will be reported subsequently.

The solubility experiments were made as described by Grönwall.[1] In every case a suspension of the crystalline material containing 7.6 ± 0.1 mg. protein nitrogen per ml. was equilibrated for a twenty-four hour period. The protein nitrogen was then determined on the supernatant liquid after centrifugation. The table illustrates the variations in the electrophoretic components of a number of preparations at *p*H 4.8 and the solubility of these crystalline fractions at *p*H 5.2 ± 0.1.

Preparation	% Composition, $\mu \times 10^5$ 1.4–1.5	2.2–3.2	Solubility (mg. N/cc.) at 25° in H_2O	in 0.02 *M* NaCl
B[a]	28	72	0.08	1.1
A[a]	38	62	.10	1.3
Standard	40	60	.12	1.5
K[b]	46	54	.18	1.7
L[b]	54	46	.13	1.6
M[b]	59	41	.19	1.9

[a] Preparations made from the standard preparation by means of acetate buffer of varying *p*H. [b] From filtrate of standard β-lactoglobulin preparation.

The variation in solubility in salt solutions of preparations with different electrophoretic compositions is an adequate explanation for the divergent solubility data reported by Palmer[3] and Grönwall.[1]

EASTERN REGIONAL RESEARCH LAB. T. L. MCMEEKIN
U. S. DEPARTMENT OF AGRICULTURE B. D. POLIS
AGRICULTURAL RESEARCH ADMINIS. E. S. DELLAMONICA
PHILADELPHIA 18, PA. J. H. CUSTER

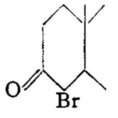
RECEIVED DECEMBER 15, 1947

THE PREPARATION OF 3-KETO-Δ4-STEROIDS

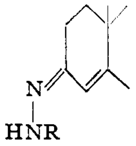
Sir:

For the introduction of a double bond at $C_4 : C_5$ in 3-ketosteroids hydrogen bromide has been eliminated from the 4-bromo derivative by treatment in boiling pyridine.[1] However, the yields have not been good. A new method for dehydrobromination has been found in which the reactivity of the bromine at C_4 is greatly increased through formation of a hydrazone at C_3. For a model compound, methyl 3,11-diketo-12-bromocholanate[2] was brominated to give methyl 3,11-diketo-4,12-dibromocholanate (I), m. p. 157.5–158.5°, $[\alpha]_D$ +13 ± 2° (*c* 0.96 in chloroform). Calcd. for $C_{25}H_{36}O_4Br_2$: C, 53.58; H, 6.48; Br, 28.52. Found: C, 53.56; H, 6.69; Br, 28.5.

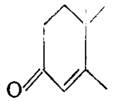
When 2,4-dinitrophenylhydrazine (1.2 equivalents) in the absence of molecular oxygen was added to an acetic acid solution of I which contained 5 equivalents of sodium acetate, a hydrazone was formed and hydrogen bromide was eliminated to give methyl 3,11-diketo-12-bromo-Δ4-cholenate-3-(2,4-dinitrophenylhydrazone) (II). Hydrogen bromide also was eliminated quantitatively without the use of sodium acetate. Red needles, m. p. 238–239°; $\lambda_{max.}$ 387 mμ, log ε 4.48 (chloroform); yield, 82%. Calcd. for $C_{31}H_{39}O_7BrN_4$: C, 56.45; H, 5.96; Br, 12.12. Found: C, 56.21; H, 5.89; Br, 11.93.



I



II



III

The hydrazone (II) was converted into III by treatment in 20 cc. of chloroform, 30 cc. of pyruvic acid and 2.2 cc. of 2.3 *N* hydrogen bromide in acetic acid at 45°. After two and one-half hours the pyruvic acid 2,4-dinitrophenylhydrazone was removed with aqueous sodium bicarbonate, any C_{24}-carboxyl was esterified and III was separated in about 90% yield. The product was identical with a sample of III prepared by dehydrobromination of I in pyridine, m. p. 190–191°, $[\alpha]_D$ +29 ± 2° (*c* 1.00 in chloroform); $\lambda_{max.}$ 238 mμ, log ε 4.22 (methanol). Calcd. for $C_{25}H_{35}O_4Br$: C,

---

[3] Palmer, *J. Biol. Chem.*, **104**, 359 (1934).

[1] Adolf Butenandt and Josef Schmidt, *Ber.*, **67**, 1901 (1934).

[2] R. B. Turner, V. R. Mattox, L. L. Engel, B. F. McKenzie and E. C. Kendall, *J. Biol. Chem.*, **166**, 345 (1946).

62.62; H, 7.36; Br, 16.67. Found: C, 62.44; H, 7.57; Br, 16.77.

A solution of hydrazone II is autoxidizable and should be protected from molecular oxygen with an inert gas. Before autoxidation of the unsaturated hydrazone was recognized dehydrobromination of the 4-bromo intermediates to give the adrenal hormones dehydrocorticosterone acetate (IV) and 17-hydroxydehydrocorticosterone acetate (V) was carried out. The yields of bromide ion, unsaturated hydrazone and IV were 96, 59 and 80% respectively. Slightly lower yields of V were obtained. The preparation of IV and V under more favorable conditions is planned as soon as the essential intermediate compounds again become available.

For the precursors of the adrenal hormones we are indebted to Merck and Co., Inc., Rahway, New Jersey.

DEPARTMENT OF BIOCHEMISTRY
MAYO FOUNDATION VERNON R. MATTOX
ROCHESTER, MINNESOTA EDWARD C. KENDALL

RECEIVED JANUARY 19, 1948

VALENCE INDUCTIVITY

Sir:

Experiments conducted in this Laboratory on supported transition group oxides have led to what appears to be a new general principle in inorganic chemistry.

The effect was first observed in supported oxides of manganese. Thermal decomposition of manganous nitrate is well known to yield manganese dioxide. But if the manganous nitrate is first impregnated on high-area *gamma*-alumina, the oxidation state of the manganese becomes dependent on the concentration of the manganese. On an alumina with area of about 200 sq. m. per g., and with a manganese concentration less than 5%, all the manganese is in the +3 oxidation state. At higher concentrations the manganese reverts increasingly to the +4 state. The oxidation state is established both by direct titration, and by measurement of the magnetic moment.

This observation suggests that the manganese oxide tends to assume the crystal structure of the alumina, even to the extent of changing oxidation state in order to do so. It is well known that the *gamma* form of manganese(III) oxide is isomorphous with *gamma*-alumina.

Confirmation of this view is obtained by supporting the manganese on a high area (94 sq. m. per g.) rutile, which is isomorphous with pyrolusite. In this case the oxidation state of the manganese remains at +4, even at very low concentrations.

The experiments leading to the above results were performed by Marylinn Ellis and Kathryn Wethington. The effect described is strikingly shown by supported nickel in experiments performed by Fred N. Hill.

Magnesia is isomorphous with nickel(II) oxide. Supported nickel prepared by impregnation and ignition of high area magnesia is all in the +2 oxidation state. But dilute nickel oxide supported on high area alumina has the nickel in the +3 oxidation state. This state is established by direct quantitative uptake of hydrogen during reduction, and by magnetic measurements. It has proved possible to obtain up to 10% of nickel, all in the +3 state, by using multiple impregnation. This supported oxide is pale blue in color.

On the other hand, copper supported on alumina shows no tendency to assume the +3 state, nor does supported silver. It appears, therefore, that the effect is shown only when the supported positive ion may fairly readily assume a charge and radius similar to that of the positive ion in the support.

The effect illustrated for manganese and nickel may be described as a leading of the supported ion to take a different oxidation state than it would normally assume. Following a suggestion by Prof. Robert L. Burwell, Jr., we offer "valence inductivity" as a name for this effect. The general principle illustrated may be described as an induced change of valence brought about when a transition group ion is supported on a high area surface with which it may become isomorphous, and in which the ions may become isometric. The effect is somewhat related to the phenomenon of oriented overgrowth. It may be expected to aid in the understanding of several problems in catalysis and promoter action, and to be useful in the fields of mineralogy, crystallography, and coprecipitation.

DEPARTMENT OF CHEMISTRY P. W. SELWOOD
NORTHWESTERN UNIVERSITY
EVANSTON, ILLINOIS

RECEIVED JANUARY 29, 1948

EXCHANGE REACTION BETWEEN FERRIC AND FERROUS IONS IN PERCHLORIC ACID USING A DIFFUSION SEPARATION METHOD

Sir:

Equilibrium exchange was found to occur by Nahinsky in Professor Ruben's laboratory[1] when ferrous and tagged ferric ions mixed in perchloric acid solution were immediately separated by a chemical method. In a chemical separation of the ions, other ionic or molecular species formed through chemical change during the separation may exchange. For the perchlorate system, for example, ether–hydrochloric acid extraction of iron (III) leads to erroneous conclusions, because in 6 *M* hydrochloric acid ferrous and ferric iron exchange rapidly.[2] If separation is achieved by a physical method, without changing the environment of the ions, exchange incidental to chemical changes during separation may be avoided. Since catalytic effects of trace impurities and surfaces cannot be excluded, the results must still be interpreted cautiously.

---

(1) Nahinsky, Ph.D. Thesis, University of California, 1942.

(2) Seaborg, *Chem. Rev.*, **27**, 256 (1940).

We have investigated electronic exchange between ferric and ferrous ions in 3 M perchloric acid using radioferric ion[3] as indicator. The concentration ratio of ferric to ferrous ions (R_0) was consistently close to 0.25 and the total iron concentration was 0.023 M. Exchange was allowed to proceed for periods up to nine days. No measurable oxidation of ferrous to ferric ion by air occurred in this time. Hydrolysis of ferric ion was slight.[4]

Measurement of exchange was dependent on partially separating ferric and ferrous ions by their diffusion from the exchange mixture across a sintered glass membrane[5] into 3 M perchloric acid.

Ferric and ferrous ion concentrations were determined colorimetrically at 480 mμ employing the thiocyanate method.

In thirty-minute diffusion periods at 25° ratios (R) of around 0.54 were obtained for ferric to ferrous ion concentrations in the diffusate. The average separation factor ($S = R_0/R$) of about 0.47 indicates ferric ion diffuses relatively more rapidly than ferrous ion.

If the half time for exchange is much greater than the time required for diffusion separation, it is possible to derive the relation

$$F = \left(\frac{A}{a_0W} - 1\right)\frac{R_0 + 1}{S - 1}$$

where the fraction of equilibrium exchange (F) is related to the radioactivity of the diffusate (A, in counts min.$^{-1}$), the specific activity of ferric ion (a_0, in counts min.$^{-1}$ mmol.$^{-1}$), and the millimoles of ferric ion in the diffusate (W) through the separation terms previously defined.

Initial results, subject to deviations of ± 5% by virtue of the separation technique used, are: for thirty minutes, 2.2 and 3.8% of equilibrium exchange and for one case of poor separation, 9.6%; for two days, 8.7%; for five days, 13.9%; for seven days, 26.2%; and for nine days, 30.7%. These figures give a half time for exchange[6] of 18.5 ± 2.5 days.

We gratefully acknowledge the interest shown by Professor T. de Vries. Experimental results for ferric–ferrous ion exchange under other conditions will be the subject of a later communication.

CHEMISTRY DEPARTMENT L. VAN ALTEN
PURDUE UNIVERSITY C. N. RICE[7]
LAFAYETTE, INDIANA

RECEIVED DECEMBER 13, 1947

---

[3] Supplied by U. S. Atomic Energy Commission.

[4] Rabinowitch and Stockmayer, THIS JOURNAL, **64**, 338 (1942).

[5] Northrop and Anson, *J. Gen. Physiol.*, **12**, 543 (1929).

[6] Mackay, *Nature*, **144**, 997 (1938).

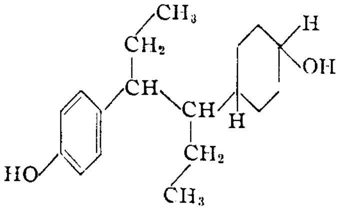
[7] Present address: The Lilly Research Laboratories, Indianapolis 6, Indiana.,

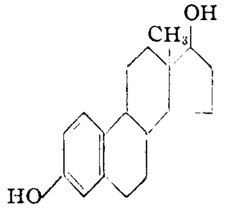
---

THE HEXAHYDRO DERIVATIVES OF *meso*-HEXESTROL

Sir:

After it was discovered that hexestrol is a highly potent estrogen, it became of considerable interest to prepare the hexahydro derivatives I since these bear a closer resemblance to the phenolic alcohol estradiol (II) than does the diphenol hexestrol itself. The activity of I would be of considerable significance in deciding whether the high physiological activities of hexestrol and diethylstilbestrol result from their superficial resemblance to the natural hormone II.


I


II

Our efforts to synthesize I (begun in 1940) have now proved successful, leading to the two stereoisomers corresponding in configuration at the bridge carbon atoms to *meso*-hexestrol. Recently Ungnade and Ludutsky [THIS JOURNAL, **69**, 2629 (1947)] reported the synthesis of the isomers of I related in configuration to the much less potent racemic hexestrol. In our synthesis hexestrol monomethyl ether was hydrogenated in the presence of copper chromium oxide catalyst at 240° and 425 atmospheres pressure, to obtain selective reduction of the phenolic ring. Demethylation of the mixture of methyl ethers by heating with methylmagnesium iodide at 170–190° gave a mixture of the phenolic alcohols I (soluble in Claisen alkali) in 40–50% yield. By recrystallization from dilute alcohol and benzene isomer A, m. p. 183–184° (cor.), was obtained. *Anal.* Calcd. for $C_{18}H_{28}O_2$: C, 78.2; H, 10.2. Found: C, 78.2; H, 9.9. The monobenzoate of A (Schotten–Baumann) melted at 68–69°. *Anal.* Calcd. for $C_{25}H_{32}O_3$: C, 78.9; H, 8.5. Found: C, 79.0; H, 8.6. Isomer B was obtained from the original filtrate and purified through the monobenzoate, m. p. 129.5–130° (cor.). *Anal.* Found: C, 79.0; H, 8.3. Saponification of the pure monobenzoate gave isomer B, which showed a variable melting point behavior. Recrystallization from benzene and from dilute methanol gave samples with the m. p. 133–134°; sublimation at 120° (0.01 mm.) resulted in material of m. p. 128–129.5°. When the 134° material was dried at 60° (0.1 mm.) the m. p. broadened to 134–141.5° (*Anal.* Found: C, 78.2; H, 10.0). Other samples melting as high as 143–145° were obtained. This behavior is indicative of polymorphism. Isomer B may be the same as the compound, m. p. 144–145°, obtained by Hoehn and Ungnade [THIS JOURNAL, **67**, 1617 (1945)] in low yield by hydrogenation of diethylstilbestrol.

Preliminary physiological assays carried out under the direction of Drs. R. K. Meyer and Elva Shipley Meyer of the Department of Zoology, indicate that both isomers are definitely weaker in estrogenic activity in rats than diethylstilbestrol. Further tests are in progress.

DEPARTMENT OF CHEMISTRY
UNIVERSITY OF WISCONSIN A. L. WILDS
MADISON 6, WISCONSIN WILLIAM B. MCCORMACK

RECEIVED JANUARY 19, 1948

STREPTOMYCIN, VII. DEGRADATION OF O-TETRAMETHYLSTREPTAMINE TO D,L-DIMETHOXYSUCCINIC ACID

Sir:

N,N′-Diacetylstreptamine (1,3-diacetamido-2,4,5,6-tetrahydroxycyclohexane)[1] was converted with dimethyl sulfate and sodium hydroxide to O-tetramethyl-N,N′-diacetylstreptamine (m. p. > 300°; *Anal.* Calcd. for $C_{14}H_{26}O_6N_2$: C, 52.81; H, 8.23; N, 8.81; CH_3O, 39.07. Found: C, 52.99; H, 8.28; N, 8.67; CH_3O, 40.1), which on hydrolysis with hydrochloric acid afforded O-tetramethylstreptamine dihydrochloride (m. p. > 300°; *Anal.* Calcd. for $C_{10}H_{22}O_4N_2 \cdot 2HCl$: C, 39.09; H, 7.87; N, 9.12; Cl, 23.08. Found: C, 38.76; H, 7.71; N, 8.98; Cl, 23.4). The free base (m. p. 83–84°; *Anal.* Calcd. for $C_{10}H_{22}O_4N_2$: C, 51.26; H, 9.46; N, 11.96. Found: C, 51.18; H, 9.38; N, 11.71) was oxidized with neutral potassium permanganate at room temperature. The methyl ester mixture formed from the oxidation products with methanolic hydrogen chloride yielded on distillation several fractions which were treated separately with methanolic ammonia or methylamine. From the lower-boiling fractions there were obtained D,L-dimethoxysuccinic acid diamide (m. p. 266–268° (dec.)[2]; *Anal.* Calcd. for $C_6H_{12}O_4N_2$: C, 40.90; H, 6.87; N, 15.90; CH_3O, 35.2. Found: C, 41.14: H, 6.76; N, 16.19; CH_3O, 35.0) and D,L-dimethoxysuccinic acid di-N-methylamide (m. p. 188–189°; *Anal.* Calcd. for $C_8H_{16}O_4N_2$: C, 47.04; H, 7.90; N, 13.72; CH_3O, 30.4. Found: C, 47.46; H, 8.05; N, 13.96; CH_3O, 30.5). Synthetic specimens of these hitherto undescribed amides, prepared from D,L-tartaric acid, showed the same melting points, undepressed by admixture of the degradation products. The diamide[3] and di-N-methylamide[4] of *meso*-dimethoxysuccinic acid melted at 255–257° (dec.) and 210–210.5°, respectively. The *meso*-diamide strongly depressed the melting point of the diamides derived from streptamine and from D,L-tartaric acid.

If, as appears highly probable, streptamine and streptidine are *meso* compounds, it follows from the above results that the 5-hydroxyl group is oriented *trans* with respect to the 4- and 6-hydroxyl groups (*xylo*-configuration), a spatial arrangement also encountered at the corresponding positions in *meso*inositol. This would limit the number of possible *meso* forms for streptamine to four.

Furthermore, there was isolated from high-boiling ester fractions prior to amidation a compound $C_{10}H_{17}O_6N$ (m. p. 109–110°; *Anal.* Calcd C, 48.57; H, 6.93; N, 5.67; $4CH_3O$, 50.1. Found: C, 48.74; H, 6.87; N, 5.65; CH_3O, 49.2) which should be either the 2,6-lactam of a 2-amino-3,4,5-trimethoxyadipic acid-1-methyl ester, or the 3,6-lactam of a 3-amino-2,4,5-trimethoxyadipic-1-methyl ester.

DIVISION OF ORGANIC CHEMISTRY O. WINTERSTEINER
SQUIBB INSTITUTE FOR MEDICAL RESEARCH
NEW BRUNSWICK, NEW JERSEY ANNA KLINGSBERG

RECEIVED JANUARY 19, 1948

---

[1] R. L. Peck, C. E. Hoffhine, E. W. Peel, R. P. Graber, F. W. Holly, R. Mozingo and K. Folkers, THIS JOURNAL, 68, 776 (1946).

[2] All melting points reported are corrected.

[3] W. N. Haworth and E. L. Hirst, *J. Chem. Soc.*, 1858 (1926).

[4] W. N. Haworth and D. I. Jones, *ibid.*, 2349 (1927).

NEW BOOKS

SMALL WONDER—The Story of Colloids. By GESSNER G. HAWLEY. Alfred A. Knopf, New York, N. Y., 1947. 220 pp. Price $3.50.

The author certainly deserves full credit for his courage and, as many parts prove, skill in attempting to explain "to those who have the curiosity but lack the time to study more erudite treatises" what the term "colloid" implies. In his preface he also states, however, that "there is such a thing as having an embarrassment of subject matter" and "it is easy to become bewildered and get lost in it all." That has happened, and it is unfortunate because the reader who takes many of the written words at their face value will not get a more general, but often a wrong, understanding of colloids. This refers specifically to the author's definition of colloids, the history of this branch of science, to the explanation of their electrical properties, to the discussion of how natural rubber is obtained, or synthetic rubber produced.

The discussion of the electron microscope is by far too extensive for a book of this type, so much the more since its use in attempting to pry deeper into the structure of lyophilic colloids is becoming more and more questionable. It would have been far more appropriate to explain to the reader the tremendous developments of ultramicroscopic techniques and what has been achieved by their use.

The author has made a few serious mistakes which must be corrected. Plate XV is a cut-away view of a Sharples *super-* and not ultracentrifuge. The statement that the first periodical given over exclusively to colloid chemistry is the "Journal of Colloid Science," which appeared for the first time in January, 1946, is incorrect; the Germans have had the periodicals "Kolloid Zeitschrift" and "Kolloid-