[Contribution from the Department of Chemistry of the University of Delaware and the Ballistics Research Laboratory of the Aberdeen Proving Grounds]

The Rates and Equilibria of Hydrogen–Deuterium Exchange in Hydroxylic Compounds

By Harold Kwart, Lester P. Kuhn and Ernest L. Bannister

Received December 21, 1953

Studies are reported on the rate and equilibria of exchange in systems comprised of varying proportions of D_2O, HOD, H_2O, CH_3OH and CH_3OD both in liquid and gas phase reaction. Calculation of data obtained by application of infrared spectroscopic techniques as well as the bearing of these results on the mechanism of exchange are discussed.

Introduction

The exchange of hydrogen for deuterium in the O–H bond of an alcohol, ROH, has been studied previously.[1–3] Orr has reported measurements on both the rate of exchange and the position of equilibrium in the instance supposedly corresponding to the reaction of C_2H_5OH in equation 2 (see below). He found that the equilibrium constant for this reaction, $K_2 = 1.1$ and that the rate of exchange is exceedingly slow, equilibrium being established only after approximately 16 hours at room temperatures. The other workers,[2,3] however, obtained results indicating that equilibrium is established too rapidly for measurement. Indeed, the experimental value obtained by Orr for the equilibrium constant K_2 seems questionable as well. Application of ordinary statistical considerations[4] to the interlocked equilibria depicted in the equation

$$CH_3OH + D_2O = CH_3OD + HOD,\ K_1 = 2.00 \quad (1)$$
$$CH_3OH + HOD = CH_3OD + HOH,\ K_2 = 0.50 \quad (2)$$
$$HOD + HOD = DOD + HOH,\ K_3 = 0.25 \quad (3)$$

leads to the listed values of the corresponding equilibrium constants, on the assumption that hydrogen and deuterium distribution would be random. It is known, furthermore, that the assumption made is a good first approximation. Urey[5] reports a value of $K_3 = 0.254$ at 298°K. in close agreement with the value derived above. It is to be expected that the experimental values of the remaining equilibrium constants must lie fairly close to the statistical values and, thus, Orr's experimental K_2 value, differing by almost a factor of two from the statistical value, must be seriously in error.

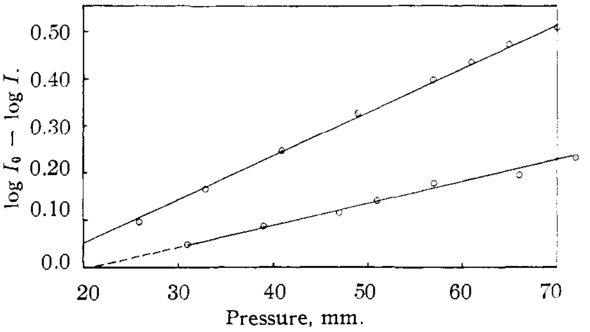


Fig. 1.—Dependence of vapor phase absorption on pressure: top curve, $CH_3OH + D_2O$; bottom curve, $CH_3OH + H_2O$.

Each[2,3] of these workers employed a technique of measurement which involved separation of the reaction components, water and ethanol, by chemical means, *i.e.*, by adding desiccants such as calcium sulfate, and by physical means, *i.e.*, distillation. The possibility of catalysis by the chemical agents or acceleration of the reaction by prolonged distillation existed in these experiments. We were, therefore, interested in arriving at a determination of the rate and equilibrium constants both in gas phase and liquid phase reaction without involving the possibility of catalysis. Our interest in this problem extends to obtaining some understanding of the mechanism of the exchange reaction and its bearing in turn on reaction mechanisms subject to acid–base catalysis.[6]

Experimental Procedures

In our experiments we chose methanol as the exchanging molecule rather than ethanol used by Orr. Later in the course of our studies we verified that no detectable difference in the results occurs with this substitution.

The procedure for following the liquid phase reaction of methanol and water is most clearly delineated with reference to the apparatus employed. This consisted of a gas-tight optical cell connected across its width by means of a stopcock to a mixing bulb. The appropriate quantities of reagents were rapidly weighed by difference and transferred to the mixing bulb. After the contents were briefly shaken and the bulb fitted into place on the cell, one drop of liquid was admitted to the evacuated cell chamber for analysis. This phase of the operation required less than two minutes from the time of mixing to the first reading on the spectrophotometer.

The cell consisted of a ten centimeter glass annulus equipped with rock salt windows thinly coated on the inside with petroleum jelly to protect them from the aqueous components of the product. Bands at 11.95 and 7.45 μ were used for the determination of MeOD and MeOH, respectively. HOH, HOD, DOD and the petroleum jelly absorption did not interfere. The intensity of absorption was found to be a linear function of the pressure in the range in which we were working and calibration diagrams of known proportions of the constituents readily were obtained.

The accompanying calibration curve data were obtained in the following manner. A drop of an equimolar mixture of water and methanol was introduced into the evacuated absorption cell, the pressure noted and the reading obtained at the specific wave length under consideration. The cell was then pumped successively to lower pressures and the appropriate spectrophotometer readings repeated for each

---

[1] W. J. C. Orr, *Trans. Faraday Soc.*, **32**, 1033 (1936).

[2] J. C. Jungers and K. F. Bonhoeffer, *Z. physik. Chem.*, **A177**, 460 (1936).

[3] While this work was proceeding we were informed by Dr. J. Hine of the Georgia Institute of Technology that related rate studies using C_2H_5OH were in progress in his laboratories. Since then the report of these workers has appeared, J. Hine and C. H. Thomas, This Journal, **75**, 739 (1953).

[4] This may be demonstrated in a variety of ways. Perhaps the briefest among these utilizes a principle of statistics discussed by W. Feller, "Probability Theory and its Applications," Vol. I, Chapt. V, John Wiley and Sons, Inc., New York, N. Y., 1950.

[5] H. C. Urey, *J. Chem. Soc.*, 569 (1947).

[6] R. P. Bell, "Acid-Base Catalysis," Oxford University Press, London, 1941, Chapt. VII; L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., Chapt. VIII.

new condition. Within the range of water–alcohol compositions which we used the extinction coefficients, *k*, at the above mentioned wave lengths were independent of this variable. It will be noted that the respective curves do not intersect at the zero pressure axis. This calibration, however, must be regarded as purely empirical in nature and the theoretical points of intersection of the two curves are not to be expected because of such factors as zero absorption of the cell, residual water vapor pressure and limited obedience to Beer's law. Again, it is to be noted that the measured values were independent of the size of the drops we introduced, obviating any concern about the completeness of vaporization of the components.

In using these calibration curves to obtain *R*, the ratio of the concentrations of CH_3OD and CH_3OH in a sample of unknown composition, all that we required was a reading of the cell pressure, the absorptions at the two specific wave lengths and calculation from the relationship

$$R = \frac{C_{\mathrm{CH_3OD}}}{C_{\mathrm{CH_3OH}}} = \frac{(\log I_0/I)_{\mathrm{CH_3OD}}k_{\mathrm{CH_3OH}}}{(\log I_0/I)_{\mathrm{CH_3OH}}k_{\mathrm{CH_3OD}}}$$

The procedure for following the exchange reaction of methanol and water in which there existed no liquid phase mixture of the reagents differed only in the method of preparing a reaction sample and introducing it into the cell for analysis. The apparatus consisted of two flasks, B and C, equipped with stopcocks on a manifold and the manifold in twin connected through a stopcock to a gas-tight optical cell A similar to the one used above. B contained approximately two ml. of D_2O and C a similar amount of CH_3OH. The flasks were frozen down in liquid nitrogen and the contents thoroughly degassed in several cycles of pumping down and thawing to room temperature under a diminished pressure of approximately 0.1 mm. In a typical run, A was introduced into the light path of the infrared spectrophotometer with the wave length scale set at the CH_3OD band used for analysis and exhausted to about 0.1 mm. before shutting down the line leading to the pumping system. Flask B was warmed slightly, D_2O vapor admitted and the pressure in the cell recorded. Vapors of CH_3OH were then admitted with a quick twist of the stopcock plug from flask C, the new pressure recorded and the absorption intensity noted.

A series of liquid phase exchange reactions were run in which the attempt was made to estimate the rate of hydrogen–deuterium exchange between alcoholic groups. The -OH bands of methyl -OH and tertiary butyl -OH and the -OD bands of methyl -OD and tertiary butyl -OD (each reagent in dilute carbon tetrachloride solution), can be distinguished with lithium fluoride optics in our Baird spectrophotometer. The tertiary butyl -OH band appears as a shoulder on the long wave length side of the methyl -OH band. The tertiary butyl -OD band appears as a shoulder on the methyl -OD band. One hundredth molar solutions of the alcohols were used. The exchange of the deuterium in methyl -OD for the hydrogen in *t*-butyl -OH and the converse experiment was observed by mixing aliquots of the proper solutions and rapidly transferring the mixture to a 2-cm. rock-salt window liquid cell.

Calculations and Results

The extent to which equilibria 1, 2 and 3 have been attained in the isotope exchange may be readily calculated from the measured *R* value and the initial concentrations of the reagent materials. Thus, since

$$K_1 = \frac{(\mathrm{CH_3OD})(\mathrm{HOD})}{(\mathrm{CH_3OH})(\mathrm{D_2O})},\ K_1 = \frac{\mathrm{R(HOD)}}{(\mathrm{D_2O})}$$

$$K_2 = \frac{(\mathrm{CH_3OD})(\mathrm{H_2O})}{(\mathrm{CH_3OH})(\mathrm{HOD})},$$

Thus

$$(\mathrm{D_2O}) = \frac{R}{K_1}(\mathrm{HDO}) = \frac{R^2}{K_1K_2}(\mathrm{H_2O})$$

$$K_3 = \frac{(\mathrm{D_2O})(\mathrm{H_2O})}{(\mathrm{HOD})^2},\ \mathrm{HOD} = \frac{R}{K_2}(\mathrm{H_2O})$$

it will be seen that the product of $1/K_1$ and K_2 gives K_3. The same result will thus be obtained from considering equil. 3 and 2 or equil. 3 and 1 and therefore a general derivation based on either pair of the above equilibria will apply to both cases. The total number of equivalents of exchangeable hydrogen is related to the total number of equivalents of deuterium by the equation

$$(\mathrm{CH_3OH}) + (\mathrm{HOD}) + 2(\mathrm{H_2O}) = M[(\mathrm{CH_3OD}) + (\mathrm{HOD}) + 2(\mathrm{D_2O})] \quad (4)$$

where

$$M = \frac{\text{no. of equiv. of exchangeable hydrogen}}{\text{no. of equiv. of exchangeable deuterium}}$$

The number of equivalents of methoxyl radical (containing exchangeable hydrogen or deuterium) is related to the total number of equivalents of exchangeable deuterium at equilibrium by the equation

$$(\mathrm{CH_3OH}) + (\mathrm{CH_3OD}) = N[(\mathrm{CH_3OD}) + (\mathrm{HOD}) + 2(\mathrm{D_2O})] \quad (5)$$

where

$$N = \frac{\text{no. of equiv. of } (\mathrm{CH_3O})}{\text{no. of equiv. of exchangeable deuterium}}$$

On making appropriate substitution into equation 5 and simplification and rearrangement of the terms we obtain the relationship

$$\frac{1 - MR}{(1 - N)R + 1} = \frac{M}{N}\left[\frac{\frac{2K_3}{K_2}\mathrm{R} - \frac{2K_2}{MR} + 1 - \frac{1}{M}}{1 + \frac{2K_3}{K_2}R}\right] \quad (6)$$

When the value of K_3 obtained by Urey is substituted into equation 6 the resulting expression may be used to calculate K_2, and thereby K_1 as well, from data on the variation of *R* with *N* and *M*.

With regard to the rate of attainment of a time-stationary *R* value our results were very clear cut. In each case complete reaction appeared to take place almost instantaneously; whether the reagents were mixed in the liquid phase and *R* determined from the vapor composition, or whether the reagents existed (in admixture) only in the vapor phase, or whether the reagents were mixed in dilute solution and *R* determined directly from the solution composition the results showed a constant value of *R* was attained in considerably less than the time required for a full measurement, less than two minutes from the time of mixing. The constancy of *R* was examined in each case over a period of six hours.

The data in Table I represent typical results obtained by analysis (in the vapor state) of the liquid mixtures listed. The wide range (0.001–0.092) of uncertainty associated with the values of *R* is due to the variation in sensitivity of the spectrophotometer over the range of absorption intensities encountered. The uncertainties in the K_2 values in turn depend upon the uncertainties in the *R* value from which they have been computed by means of the differential equation that relates the two quantities. The uncertainty in the mean value of K_2 has been computed from the individual uncertainties in the four independent observations of *R*, allowing statistically for the propagation of errors through calculation.[7]

---

(7) T. K. Sherwood and C. E. Reed, "Applied Mathematics in Chemical Engineering," McGraw-Hill Book Co., New York, N. Y., 1939, p. 377.

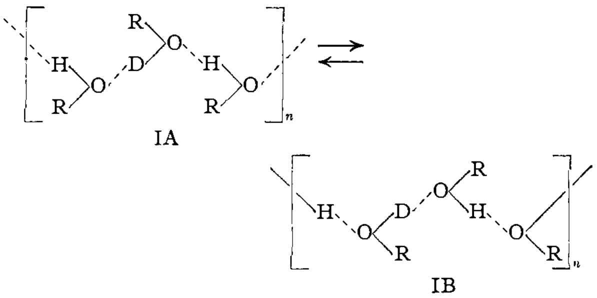
TABLE I

DATA FOR EQUILIBRIUM CONSTANT CALCULATIONS OBTAINED FROM LIQUID PHASE MIXTURES

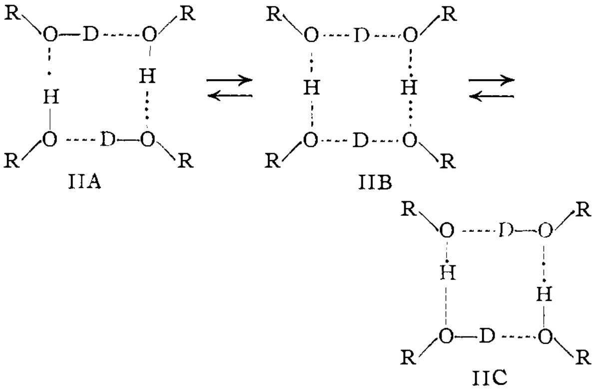
Run no.	Initial moles					Calcd.	
	CH_3OD	CH_3OH	HOH	D_2O	*R*, obs.	K_1	K_2
1	....	0.0481	0.0275	0.0275	0.500 ± 0.092	1.82	0.46 ± 0.08
2	....	.0481	.0275	.0275	.482 ± .022	1.74	.44 ± .03
3	0.0245	....	.0220	....	.467 ± .021	1.59	.40 ± .03
4	....	0.121	.0825	.0275	.199 ± .001	2.10	.53 ± .004
					Av.	1.81 ± 0.08	0.46 ± 0.02

Interpretation of Results.—It will be noted that the experimental values of the equilibrium constants are close to the statistical values and contrary to the findings of Orr. Furthermore, though no quantitative measure of the rate of exchange could be obtained by our procedure, it is clear from the estimable upper limit of this quantity that the reaction half-life was less than 20 seconds. Orr has nicely justified a very much slower reaction velocity on the assumption of an ionic mechanism[8] for exchange in which the rate he reported was close to the maximum possible rate consonant with the postulated mechanism. Our results would seem to indicate that a net proton transfer is probably not the essential step in the mechanism of any of the reactions 1, 2 or 3. The similarly high velocity of reaction we observed in the gas phase exchange, where ions could not be involved, demonstrates further the probability that the reaction in the liquid phase does not proceed solely through the agency of ions.

The observed high rates of exchange may be readily accounted for by the assumption that shifting of hydrogen and deuterium atoms and establishment of bonds to neighboring atoms occurs readily within the highly flexible polymeric association structures (IA and IB) that characterize water and alcohols.[9] It is entirely conceivable that this equilibrium is attained through a four-center transition state as suggested by Hine,[3] with high probability, however, that this state is formed within the polymeric association structures. In the vapor state association as a cyclic tetramer affords a mechanism of exchange as depicted by the mobile equilibrium between IIA and IIC (*cf.* Weltner and Pitzer, ref. 9), proceeding through a transition state IIB, in which each hydrogen (or deuterium) is partially bonded to two oxygen atoms. It must be assumed, however, that breakdown and reformation of the tetrameric structure occurs very readily on collision.


IA

IB


IIA

IIB

IIC

The conclusion we have reached that hydrogen–deuterium exchange reactions in hydroxylic compounds proceed largely through the agency of the unshared pairs of oxygen is borne out also by the results of others working with compounds of nitrogen and hydrogen. Turner[10] has shown that the exchange of the hydrogen of ethylenimine with D_2O in the gas phase is instantaneous for all practical purposes. Brodskii and Sulima[11] have confirmed this result with ammonia in solution and have shown, in addition, that in ammonium ion-exchange of the hydrogen on nitrogen is a comparatively slow reaction. Apparently, in the latter case, tying up of the unshared pairs on nitrogen has vitiated the mechanism involving hydrogen bonding by which the rapid exchange was possible

---

(8) Thus he assumed in the initial step

$$HOD + ROH \underset{K_b}{\overset{K_a}{\rightleftarrows}} OH^- + ROHD^+$$

and that (OH^-) and $(ROHD^+) \cong 10^{-7}$, (HOD) = 55.5 and ROH = 17.1. Then $K = k_a/k_b = 1.05 \times 10^{-17}$. If recombination occurs at every collision of the ions, *i.e.*, $k_b \cong 2.77 \times 10^{-11}$, then the rate of the forward dissociation reaction $K_a \cong 2.8 \times 10^{-6}$ l./g. mole-sec., a value close to the one he reported.

(9) F. A. Smith and E. C. Creitz (*J. Res. Natl. Bur. Standards*, **46**, [2] 145 (1951)) have only recently considered the types of aggregation obtaining in the liquid state of hydroxylic compounds and have concluded that association in alcohols may be represented by structures similar to I. The mean state of aggregation passes successively from the polymeric, through the tetrameric, trimeric and dimeric states with increasing dilution by inert solvents. A similar picture has been previously discussed for the associated state of liquid water; J. D. Bernal and R. H. Fowler, *J. Chem. Phys.*, **1**, 520 (1933); G. S. Forbes, *J. Chem. Educ.*, **18**, 18 (1941). In the vapor state W. Weltner and K. Pitzer, THIS JOURNAL, **75**, 2606 (1953) have shown that the data can best be accounted for by assuming an aggregate of alcohol molecules as represented by the structure II. This tetrameric configuration with one hydrogen (or deuterium) bond per alcohol unit appears to have just the right steric requirements for a stable polymer. The Grotthus mechanism which has been invoked in consideration of the extremely high mobility of hydronium ions in water [A. Gierer and K. Wirtz, *J. Phys. Chem.*, **56**, 914 (1952)] suggests that even protons depend for their transfer on such polymeric association of the solvent.

(10) Private communication of unpublished results from Dr. T. Turner of the Ballistics Research Labs. of the Aberdeen Proving Ground, Maryland.

(11) A. I. Brodskii and L. V. Sulima, *C. A.*, **45**, 424*a* (1951).

in ammonia in solution and in ethylenimine in the gas phase.

Acknowledgment.—We wish to express our gratitude to our colleagues including Drs. Turner, Pigford and Sager for many helpful discussions during the progress of our work.

[CONTRIBUTION FROM THE RICHARDSON CHEMICAL LABORATORY, TULANE UNIVERSITY]

The Heats of Combustion of Cyclopentyl and Cyclohexyl Azides

BY THOMAS F. FAGLEY AND HULON W. MYERS

RECEIVED JUNE 18, 1954

The ΔH's of combustion at 298.1°K. of cyclopentyl and cyclohexyl azides were found to be 820.48 ± 0.33 and 965.96 ± 0.21 kcal. per mole, respectively. From these the enthalpies of formation from the elements and from the gaseous atoms have been calculated. The resonance energy of the azido group has been estimated.

In continuation of a study of the physical properties of organic azides in this Laboratory, the heats of combustion of cyclopentyl and cyclohexyl azides have been measured. The only other recorded heats of combustion of organic azides have been those of 2-triazoethanol,[1] hydrazoic acid,[2] triazoethyl acetate and phenyl azide.[3] Since the structural properties of the azido group are of interest the resonance energy has been estimated.

Experimental

Preparation and Purification of Materials.—Both azides were prepared by the method of replacing the bromine of the mono-substituted cyclic bromide with the azido group by reaction with sodium azide. The procedure used in the preparation was that of Boyer[4] and is essentially as follows. To a solution of the bromide (0.24 mole) in 200 ml. of ethanol was added a solution of sodium azide (0.48 mole) and 10 g. of sodium acetate in 100 ml. of water, and the resulting mixture heated at reflux temperature for 12 hours. Addition of 500 ml. of water caused two layers to form; these were separated and the bottom layer was extracted with ether. The ether extracts were added to the top oily layer and the solution dried over sodium sulfate, which was then removed by filtration. Distillation of the ether left an oily residue. Distillation at reduced pressure, in an atmosphere of nitrogen, of the cyclopentyl and cyclohexyl azides gave excellent yields at 51–52° (20 mm.) and 64–65° (21 mm.), respectively.

The samples used in this investigation had been purified by triple distillation at reduced pressure. A nearly constant refractive index (±0.0001) for successive constant boiling fractions was considered evidence for the purity of the compounds. The refractive index for the sodium-D line at 25° was 1.4690 for cyclohexyl azide and 1.4615 for cyclopentyl azide. The densities at 25° were 0.98546 g./ml. and 0.9789 g./ml. for cyclohexyl azide and cyclopentyl azide, respectively. The heat capacities of the liquids, as measured by the method of mixtures, were found to be 19.4 abs. j./g. for cyclopentyl azide and 19.2 abs. j./g. for cyclohexyl azide, with an estimated error of ±1.2 abs. j./g. The heat of combustion was a more sensitive criterion of purity of the compounds than the boiling points or refractive indices.

The constant temperature jacket calorimeter used in this work consisted of a Parr Adiabatic Oxygen Bomb Calorimeter, which has been previously described,[1] with appropriate modifications for isothermal measurements. The major change involved the installation of a thermal regulator and heating unit, and insulation between the bucket well and the walls by filling with vermiculite. A coil of twenty-gage nichrome wire, wound on a glass tube and coated with glyptal, comprised the electrical heater. The temperature was controlled by an American Instrument Company Supersensitive Mercury Relay and an Aminco Rectran Rectifier–Transformer. With these changes, the jacket was maintained at constant temperature to ±0.002°.

Temperature changes were measured with a calorimetric-type platinum resistance thermometer constructed and calibrated by the Leeds and Northrup Company, and checked by comparison with standard research platinum thermometers certified by the National Bureau of Standards. The thermometer, of the four-lead cable type, was used in conjunction with a Leeds and Northrup G-2 Mueller Bridge and a high sensitivity galvanometer.

The rating-period procedure of White[5] with modifications recommended by Prosen[6] was employed in order to obtain the correct resistance change.

The spherical glass bulbs into which the samples were placed for ignition were prepared from 4 mm. soft glass tubing according to the procedure of Coops.[7] The bulbs averaged 11 mm. in diameter and, with a 20-mm. capillary neck, weighed between 75 and 150 mg. With both liquids it was found that carbon-free combustion was obtained with completely filled bulbs at a pressure of 35 atmospheres of oxygen.

An ordinary 2-cc. hypodermic syringe with a number 22 needle into which was fitted a fine glass capillary was used for filling the bulbs. After the bulb was filled with the freshly distilled azide it was placed in crushed Dry Ice and the vapors in the neck removed by placing a small warmed coil of copper around it. The tip of the neck was then sealed with a micro-burner flame.

The preparation for ignition and the determination of the nitric acid formed were performed as described previously.[1] Ten milliliters of water were placed in the bomb with the sample.

Calibration.—The energy equivalent of the calorimeter was determined by burning Bureau of Standards benzoic acid (Standard sample 39 g.). For its isothermal heat of combustion per gram under standard conditions at 25° the value of 26,433.8 abs. j./g. mass (weight *in vacuo*) was reported with an estimated uncertainty of 2.6 j./g. This value was converted[8] to the bomb conditions used throughout this investigation and found to be 26,434.6 abs. j./g.

In a series of calibration determinations, the mean energy equivalent for the system was 19,705.5 ± 3.5 joules per ohm. Since the amount of water in the bucket (m_w) varied from one experiment to another, it was not included in the standard calorimeter system.

Results and Calculations

All weights were corrected to "*in vacuo*." The data are referred to a standard temperature of 25°.

---

[1] T. F. Fagley, J. F. Albrecht and E. Klein, THIS JOURNAL, **75**, 3104 (1953).

[2] P. Gunther and R. Meyer, *Z. physik. Chem.*, **A175**, 154 (1935).

[3] W. A. Roth and F. Muller, *Chem. Ber.*, **62**, 1190 (1929).

[4] Private communication from Dr. J. H. Boyer of this Laboratory.

[5] W. P. White, "The Modern Calorimeter," American Chemical Society Monograph Series (No. 42). The Chemical Catalog Company, New York, N. Y., 1928.

[6] E. J. Prosen, "Determination of Heats of Combustion Using a Bomb Calorimeter," National Bureau of Standards Report No. 1119, August 6, 1951; U. S. Department of Commerce.

[7] J. Coops, D. Mulder, J. W. Dienske and J. Smittenberg, *Rec. trav. chim.*, **66**, 153 (1947).

[8] National Bureau of Standards Certificates for Standard Sample 39g Benzoic Acid (16-6506-1).