of Prof. C. A. Herter, as given in his book on "Infantilism."[1] He describes here a condition of arrested development, consequent upon the non-absorption of food and its subsequent putrefaction in the lower intestine. The patients excreted practically all the calcium ingested, this accounting for the failure of skeletal growth, and the feces contained neutral fat, fatty acids, and soaps in marked excess, indicating impaired fat absorption.[2]

Herter found that the addition of gelatin (which is a most efficient protective colloid) to the milk fed, caused an improved absorption and recommends its use.[3] He further observes that "in sparing protein small quantities of gelatin appear to have about as much effect as larger amounts," a fact quite in accord with protective action, for only a small percentage of gelatin is needed to accomplish protection.

In conclusion, I would point out that bald chemical analysis cannot express the digestibility and availability of a food any more than it can express or explain the action of the digestive juices, or in fact any other physiological process. In all these processes can be traced the influence of the colloidal substances everywhere present in the body, whose effects are quite out of proportion to their small mass. Striking examples are the enzymes which catalyze and direct chemical and physical changes, and the protective colloids which oppose crystallization, precipitation and coagulation, emulsify fats, and facilitate diffusion and absorption. As soon as we approach the chemistry of living organism, we are confronted with problems of colloid chemistry, and there is no doubt but that a correct understanding and application of its principles will throw much light upon many other obscure problems in biology, physiology and medicine.

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[FROM THE LABORATORY OF BIOLOGICAL CHEMISTRY OF THE HARVARD MEDICAL SCHOOL.]

ON THE ESTIMATION OF THE INTENSITY OF ACIDITY AND ALKALINITY WITH DINITROHYDROQUINONE.[4]

BY LAWRENCE J. HENDERSON AND ALEXANDER FORBES.

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Indicators serve two distinct purposes: The one is to mark sharply an end point in titration, the other to measure the concentration of ionized hydrogen or hydroxyl in solution. The properties which qualify a sub-

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[1] "On Infantilism from Chronic Intestinal Infection," by C. A. Herter, M.D. The Macmillan Co., 1908.

[2] It is now well known that the cream layer or fat of milk contains from 10 to 500 times as many bacteria as the whole milk. See (U. S. Dept. Agr.) *Bull.* 56, 737; also Jacobi, *J. Am. Med. Association, loc. cit.*

[3] Herter, *loc. cit.*, pp. 101, 105.

[4] An investigation aided by a grant from the Elizabeth Thompson Science Fund

stance for these two uses are quite different, and usually opposite. Thus our common indicators, selected for use in titration, possess a very limited value in measuring the reaction of a solution, because they change color sharply at a particular reaction, and otherwise not at all. No other means than the employment of indicators permits the rapid estimation of reaction, and even the valuable method of Salm, wherein of necessity a large number of indicators are employed, is far from fulfilling the requirements of a quick method for approximate estimations.

In seeking a solution of this problem the behavior of 2,5-dinitro-hydroquinone, among other substances, has been studied, its chemical properties being such as to suggest that it might be suitable for the purpose. This substance when dissolved in distilled water, has a bright orange color which changes on addition of moderately strong acid to pale greenish yellow, and on addition of moderately strong alkali to deep purple. Evidently this wide range of colors renders it unusually well suited to the measurement of all intensities of hydrogen ion concentration.

In order more accurately to define the characteristics of the substance and to establish a color standard, the following experiments were performed: A preparation as provided by Schuchardt was twice recrystallized from water to insure purity. From this material a 0.1 molal alcoholic solution was made up and used in all experiments. Six standard solutions of constant reaction were then prepared in the usual manner from simple bases, acids and salts, either pure or mixed, in such proportions that the solutions formed a uniformly graduated scale of acidity on the one side of neutrality, and of alkalinity on the other. The composition of the solutions and their reactions are shown in the accompanying table.

Solution.	Composition.	Reaction.	Color.
A	0.10 *N* Tartaric acid	$(H^+) = 10^{-2}$	Green-yellow
B	0.058 *N* NaH_2PO_4	$(H^+) = 10^{-4}$	Orange-yellow
C	$NaH_2PO_4 : Na_2HPO_4 = 4 : 1$	$(H^+) = 10^{-6}$	Orange
D	$NH_4OH : NH_4Cl = 1 : 27$	$(OH^-) = 10^{-6}$	Brownish cherry-red
E	$NH_4OH : NH_4Cl = 3.7 : 1$	$(OH^-) = 10^{-4}$	Red-purple
F	1.0 *N* Na_2CO_3	$(OH^-) = 10^{-2}$	Purple

Six 250 cc. flasks were filled from these stock solutions. To each flask was then added a small constant quantity (10 drops) of the indicator solution. The colors of the solutions thus tested are given in the table. Between any two of these colors an intermediate color may readily be distinguished from either, while in those mixtures between orange and purple, where the color change is greatest, several intermediate hues may be differentiated. The indicator has been employed with a variety of salts of the alkalies and the alkaline earths and with ammonium salts, but in no case has an effect of the salt upon the color been noted which is sufficient to cause a significant error in such approximate estimations.

If the indicator be left long in a solution which is even slightly alkaline, as indicated by a reddish color in the solution, or the color of more intense alkalinity, decomposition occurs and the color fades. This destruction occurs too slowly however to cause any trouble. It should be noted that the color varies greatly with the temperature, thus the purple color of high alkalinity may be converted into almost pure blue by warming.

With the aid of this substance it is, accordingly, possible to estimate the concentration of ionized hydrogen and hydroxyl in an unknown solution rapidly and accurately to about one power of ten between hundredth normal concentration of ionized hydrogen and hundredth normal concentration of ionized hydroxyl. In all, eleven different intensities of reaction may be thus differentiated. To this end it is necessary to prepare a series of standard solutions corresponding to those above described. The process then consists merely in the matching and interpolation of colors.

[FROM THE RESEARCH LABORATORY OF THE ROOSEVELT HOSPITAL, NEW YORK CITY.]

THE QUANTITATIVE DISTILLATION OF AMMONIA BY AERATION.

(SECOND PAPER.)

BY PHILIP ADOLPH KOBER.

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In a previous paper[1] a new apparatus for the quantitative distillation of ammonia was described, which, since, has been tried in several laboratories. In three cases the results have been published and it is the object of this paper to summarize, and to discuss, the criticisms that have been made, and to record the experience accumulated during the past two years.

In a recent article[2] J. Sebelien in conjunction with A. Brynildsen and O. Haavardsholm developed and reported independently the identical process devised by the writer over a year before.

They proved by many experiments that ammonia can be liberated from solutions quantitatively, even if a slow stream of air and a large volume of solution (200 cc.) are employed. They concluded that the aeration procedure has many advantages over the old method and eliminates several sources of error.

Davis[3] found that the method, using an ordinary water pump, gave good results provided the flask was insulated with asbestos. He also makes the statement that if 30 cc. of acid are digested, the flask would require additional heat during distillation.

[1] THIS JOURNAL, **30**, 1131 (1908).

[2] Read by J. Sebelien in the Section for Analytical Chemistry, Congress of Applied Chemistry, June, 1909, printed in *Chemiker-Zeitung*, **1909**, No. 87.

[3] THIS JOURNAL, **31**, 56 (1909).