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S. P. L. Sørensen.

THE DETERMINATION OF HYDROGEN IONS

An elementary treatise on the hydrogen electrode, indicator and supplementary methods with an indexed bibliography on applications

BY

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To

Fellow Workers in the Biological Sciences,

Architects of Progress,

Who Hew the Stone to Build Where Unseen Spires Shall Stand

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PREFACE TO THE FIRST EDITION

Poincaré in *The Foundations of Science* remarks, "There are facts common to several sciences, which seem the common source of streams diverging in all directions and which are comparable to that knoll of Saint Gothard whence spring waters which fertilize four different valleys."

Such are the essential facts of electrolytic dissociation.

Among the numerous developments of the theory announced by Arrhenius in 1887 none is of more general practical importance than the resolution of "acidity" into two components—the concentration of the hydrogen ions, and the quantity of acid capable of furnishing this ionized hydrogen. For two reasons the hydrogen ion occupies a unique place in the estimation of students of ionization. First, it is a dissociation product of the great majority of compounds of biochemical importance. Second, it is the ion for which methods of determination have been best developed. Its importance and its mensurability have thus conspired to make it a center of interest. The consequent grouping of phenomena about the activity of the hydrogen ion is unfortunate when it confers undue weight upon a subordinate aspect of a problem or when it tends to obscure possibilities of broader generalization. Nevertheless, such grouping is often convenient, often of immediate value and frequently illuminating. Especially in the field of biochemistry it has coördinated a vast amount of material. It has placed us at a point of vantage from which we must look with admiration upon the intuition of men like Pasteur, who, without the aid of the precise conceptions which guide us, handled "acidity" with so few mistakes.

In the charming descriptions of his experimental work Pasteur has given us glimpses of his discernment of some of the effects of "acidity" in biochemical processes. In the opening chapter of *Studies on Fermentation* he noted that the relatively high acidity of must favors a natural alcoholic fermentation in wine, while the low acidity of wort induces difficulties in the brewing of beer. He recognized the importance of acidity for the cultivation of the bacteria which he discovered and was quick to see the lack of

such an appreciation in his opponents. In describing that process which has come to bear his name Pasteur remarks, "It is easy to show that these differences in temperature which are required to secure organic liquids from ultimate change depend exclusively upon the state of the liquids, their nature and above all upon the conditions which affect their *neutrality whether towards acids or bases.*" The italics, which are ours, emphasize language which indicates that Pasteur was aware of difficulties which were not removed till recently. Had Pasteur, and doubtless others of like discernment, relied exclusively upon volumetric determination of acidity they would certainly have fallen into the pitfalls which at a later date injured the faith of the bacteriologist in the methods of the chemist. Was it reliance upon litmus which aided him? Perhaps the time factor involved in the use of litmus *paper,* which is now held as a grave objection, enabled Pasteur to judge between extremes of reaction which the range of litmus as an indicator in equilibrium does not cover. At all events he recognized distinctions which we now attribute to hydrogen ion concentrations. Over half a century later we find some of Pasteur's suggestions correlated with a marvelous development in biochemistry. The strongest stimulus to this development can doubtless be traced to the work of Sørensen at the Carlsberg Laboratory in Copenhagen and not so much to his admirable exposition of the effect of the hydrogen ion upon the activity of enzymes as to his development of methods. At about the same time Henderson of Harvard, by setting forth clearly the equilibria among the acids and bases of the blood, indicated what could be done in the realm of physiology and stimulated those researches which have become one of the most beautiful chapters in this science.

Today we find new indicators or improved hydrogen electrode methods in the physiological laboratory, in the media room of the bacteriologist, serving the analyst in niceties of separation and the manufacturer in the control of processes. The material which was admirably summarized by Michaelis in 1914, and to which Michaelis himself had contributed very extensively, presents a picture whose significance he who runs may read. There is a vast field of usefulness for methods of determining the hydrogen ion. There is real significance in the fruits so far won.

There remain many territories to explore and to cultivate. We are only at the frontier.

In the meantime it will not be forgotten that our knowledge of the hydrogen ion is an integral part of a conception which has been under academic study for many years and that the time has come when the limitations as well as certain defects are plainly apparent. While there is now no tendency nor any good ground to discredit the theory of electrolytic dissociation in its essential aspects, there is dissatisfaction over some of the quantitative relationships and a demand for broader conceptions. It requires no divination to perceive that while we remain without a clear conception of why an electrolyte should in the first instance dissociate, we have not reached a generalization which can cover all the points now in doubt. Perhaps the new developments in physics will furnish the key. When and how the door will open cannot be foreseen; but it is well to be aware of the imminence of new developments that we may keep our data as pure as is convenient and emphasize the experimental material of permanent value. We may look forward to continued accumulation of important data under the guidance of present conceptions, to distinguished services which these conceptions can render to various sciences and to the critical examination of the material gathered under the present régime for the elements of permanent value. These elements will be found in the data of direct experimentation, in those incontrovertible measurements which, though they be but approximations, have immediate pragmatic value and promise to furnish the bone and sinew of future theory. In the gathering of such data guiding hypotheses and coördinating theories are necessary but experimental methods are vital.

The time seems to have come when little of importance is to be accomplished by assembling under one title the details of the manifold applications of hydrogen electrode and indicator methods. It would be pleasing to have in English a work comparable in scope with Michaelis' *Die Wasserstoffionenkonzentration;* but even in the short years since the publication of this monograph the developments in special subjects have reached such detail that they must be redispersed among the several sciences, and made an integral part of these rather than an uncoordinated treatise by themselves. There remains the need for a

detailed exposition, under one cover, of the two *methods* which are in use daily by workers in several distinct branches of biological science. It is not because the author feels especially qualified to make such an exposition that this book is written, but rather because, after waiting in vain for such a book to appear, he has responded sympathetically to appeals, knowing full well from his own experience how widely scattered is the information under daily requisition by scores of fellow workers.

For the benefit of those to whom the subject may be new there is given in the last chapter a running summary of some of the principal applications of the methods. This is written in the form of an index to the bibliography, a bibliography which is admittedly incomplete for several topics and unbalanced in others, but which, it is believed, contains numerous nuclei for the assembling of literature on various topics.

The author welcomes this opportunity to express his appreciation of the broad policy of research established in the Dairy Division Laboratories of the Department of Agriculture under the immediate administration of Mr. Rawl and Mr. Rogers. Their kindness and encouragement have made possible studies which extend beyond the range of the specialized problems to which research might have been confined and it is hoped that the bread upon the waters may return. To Dr. H. A. Lubs is due the credit for studies on the synthesis of sulfonphthalein indicators which made possible their immediate application in bacteriological researches which have emanated from this laboratory. Acknowledgment is hereby made of the free use of quotations taken from the paper *The Colorimetric Determination of Hydrogen Ion Concentration and Its Applications in Bacteriology* published in the *Journal of Bacteriology* under the joint authorship of Clark and Lubs.

The author thanks his wife, his mother, Dr. H. W. Fowle and Dr. H. Connet for aid in the correction of manuscript and proof, and Dr. Paul Klopsteg for valuable suggestions.

It is a pleasure to know that the publication of the photograph of Professor S. P. L. Sørensen of the Carlsberg Laboratory in Copenhagen will be welcomed by American biochemists all of whom admire his work.

Chevy Chase, Maryland
March 17, 1920

PREFACE TO THE SECOND EDITION

The first edition of this book was offered to fellow workers for the reasons stated in the preface. The rapid exhaustion of two printings has revealed the extent of the demand for information upon the topics discussed; but it has also brought to the author a disquieting realization of the responsibility assumed at the first venture, and regret that his preoccupation in a distinctive although allied realm of research has prevented investigations which might have contributed data for a more complete second edition. This same preoccupation may be offered as an excuse for the deficiencies in the bibliography and its classification. Over 900 new references have been added to the eleven hundred odd said to be in the first edition; but, when it is realized that much of the newer information is contained in papers neither the title nor general subject of which would indicate that hydrogen ion concentrations have been considered, it will be appreciated that the task of the bibliographer requires more time than an investigator can afford. Indeed it will not be long before it will be as difficult to trace this information as it has become to trace all the effects of temperature. In certain fields of investigation "pH" is becoming almost as common as "°C." Were it not that the introduction of a new symbol would introduce confusion we would wish that the special interpretation of pH given in Chapter XVII of the first edition (Chapter XIX, this edition) could be symbolized by °S (degrees Sørensen).

Certain chapters of the first edition have been rewritten and all have been expanded to bring the book up to date and to meet the very helpful suggestions given in the generous reviews of the first edition, or by personal correspondence. It has been advisable, however, either to balance one suggestion against another or to rely upon one's own judgment to maintain a balance in the general treatment.

The question of a change of treatment to conform throughout to the "activity" concept has been given serious consideration. The author has been counseled by experienced teachers not to attempt such a change, but his chief reason for definitely rejecting

the proposal is simply that most of the data in use are still in terms of the older conceptions. In the recasting of this data a great deal of new experimental material must be collected and the newer conceptions must be stabilized. Anything short of a thorough revision of existing data would be but to cover the subject with a thin veneer giving the appearance rather than the substance of an up-to-date treatment.

The author is indebted to so many people for helpful suggestions that it would appear ungracious to mention but a few. However, due credit must be given to Dr. Barnett Cohen for painstaking correction of proof, to Miss Florence Lansdale for clerical assistance and to the publishers for their unfailing and courteous coöperation.

Chevy Chase, Maryland
May 22, 1922

CHAPTER I

Introduction—Some General Relations Among Acids and Bases

In a country rich in gold observant wayfarers may find nuggets on their path, but only systematic mining can provide the currency of nations.—F. Gowland Hopkins.

Why certain solvents such as water should cause or permit the splitting of a compound into electrically charged bodies, called ions, has not yet been very clearly explained. That they do has been demonstrated with reasonable certainty. The evidences are described in texts of physical chemistry and will not be reviewed here, except as they are revealed in the verification of the laws of chemical equilibria among electrolytes.

That aspect of electrolytic dissociation which is of special interest to us may be conveniently *pictured* as follows.

A chemical element is conceived to be an aggregate of unit, negative, electrical charges (electrons) grouped at relatively enormous distances about a central, neutralizing nucleus of positive electricity. The numerical value of this nucleus, in terms of the number of electrons required for neutralization, and the geometrical configuration of the positions of the surrounding electrons are supposed to distinguish the several elements.

Certain of the electrons are but weakly incorporated in the planet-like system of certain elements. When such an electron has escaped, the element is left with a unit excess of positive electricity. It is then a positive ion, a cation, having distinctive properties.

If an element is so constituted that it can hold an extra electron, the extra charge gives it new characteristics. The negatively charged element is called an anion.

Certain compounds such as HCl are made up of elements of the two types mentioned above. On electrolytic dissociation HCl breaks up in such a way that the hydrogen atom loses an electron and this is taken up by the chlorine atom. HCl, thus, dissociates

into the positively charged hydrogen ion and the negatively charged chlorine ion. The process may be represented as follows:

$$HCl \rightleftarrows H^+ + Cl^-$$

In the case of complex compounds such as acetic acid a similar exchange of an electron occurs. The group CH_3COO acts as a unit and when negatively charged becomes the acetate anion.

Frequently an element or group can lose or acquire several electrons. For instance Ca^{++} is the divalent cation of calcium and SO_4^{--} is the divalent anion of the sulfate group—called divalent because there are concerned two of those electrons which are supposed to be intimately connected with the phenomenon of valency.

In passing it is interesting to note that the hydrogen ion is unique. The element hydrogen is supposed to have but one electron to the atom. When this is lost there is left the hydrogen ion, a lone unit, positive charge.

Now this pictorial conception of the structure of elements, while pregnant with possibilities, must not be considered vital to the subject at hand. The one aspect which is vital is that there occur dissociations whereby an element or group becomes electrically charged—positively or negatively, as the case may be. It is the electrical charge which turns an element or group into a virtually new body and at the same time furnishes a handle, as it were, with which we may lay hold on it by electrical devices.

On the other hand the electrical charge does not prevent a limited application to ions of the laws of chemical equilibria. Indeed it is among dilute solutions of certain electrolytically dissociated compounds that there have been found the most exact data supporting the laws of chemical equilibria.

It is with these laws of chemical equilibria that we are chiefly concerned when dealing with the measurement of and the effects of hydrogen ion concentration. Therefore, if electrolytic ionization be granted as a fact, it is only necessary to sketch the concept of chemical equilibrium before coming to the simple, if somewhat detailed account of the special manner in which the concept is applied to acid-base equilibria.

Consider an acid of the type HA dissociating into the cation H^+ (hydrogen ion) and the anion A^-. The process may be expressed as follows:

$$HA \rightleftarrows H^+ + A^- \qquad (1)$$

Arrows are used to indicate that the process is reversible,—that among the large number of anions and cations present in a given volume some are recombining to form HA the while a portion of the HA molecules are dissociating.

This concept of a "reaction" as labile, continuous, reversible is of profound importance. So long as analysts are content to balance the two sides of a written reaction with regard only to the stoichiometrical relations, it is convenient to use the equation sign and to forget the reality implied in the use of arrows. Reactions do not go to completion and only approach completion when by design or chance the proper conditions are supplied. This reversibility of chemical reactions displays a world in flux. From it the "everlasting hills" cannot escape; but upon it life balances its intricate organization. Often this is done so nicely that the life of certain organisms is almost immortal.

In this interminable interplay of chemical reactions there occur situations when on the *statistical average* a given reaction is proceeding no faster in one direction than in the other. In such circumstances a chemical equilibrium is said to occur. Let us formulate in as simple a way as possible the condition of a chemical equilibrium.

Let brackets placed about a symbol indicate *concentration* of the bracketed "species." Thus [HA] represents the concentration of the residual, undissociated acid HA. Throughout the following discussions we shall always let it be implied that by "concentration" is meant molar concentration. A molar solution is one containing in one litre of solution that number of grams of the indicated substance which is equal to its formula weight.

In equation (1) the rate at which the concentration [HA] is being diminished because of the ionization may depend upon several physical conditions. To know these is unnecessary for the purpose at hand if we may assume that their effect on the individual molecules of HA is constant on the statistical average. Then, obviously, the rate at which reaction (1) proceeds from left to right will depend upon the concentration of HA and some constant factor which will be called k_1.

$$\text{Velocity left to right} = k_1 [HA] \qquad (2)$$

The velocity of the reverse reaction wherein the ions recombine to form HA might be supposed to be dependent only upon the

rate at which the ions in their thermal agitation collide. But it is difficult to say what degree of approach is necessary for combination or what other conditions must be fulfilled before the combination can be considered to have taken place. It is much safer then to assume only that some degree of meeting is necessary, that some average state is to be considered virtual combination and that the physical factors bringing about this state are, on the statistical average, constant. Here again then we ascribe the velocity of the reaction first to a factor dependent solely upon the numbers of ions concerned [concentration] and second another factor embracing all the known and unknown influences, exclusive of concentration. Suppose then that we start with equal numbers of H^+ ions and A^- ions and double the concentration of H^+. Evidently the number of collisions of H^+ ions with A^- ions will double. Likewise, if $[A^-]$ is doubled, the number of collisions of A^- with H^+ ions will be doubled. If both are doubled, the collisions are quadrupled. Consequently the velocity of association, in so far as it is dependent upon the concentrations of the reactants, is proportional to the *product* of these concentrations. Introducing the unknown proportionality factor representing the constant effect of all physical influences, we have:

$$\text{Velocity right to left} = k_2 [H^+] [A^-]. \tag{3}$$

We have already said that the state of equilibrium occurs when the velocity of the reaction in one direction equals the velocity in the reverse direction. Then at once by combining (2) and (3) we have:

$$\frac{[H^+] [A^-]}{[HA]} = \frac{k_1}{k_2} = K_a \tag{4}$$

For the ratio of two constants there is substituted in (4) another constant, K_a, known as the equilibrium constant. This equilibrium constant when applied to electrolytes is known as the ionization or dissociation constant.[1]

[1] It should be particularly noted that in equation (4) the brackets symbolize the concentrations occurring at the equilibrium state. Whenever numerical values are to be introduced it is to be assumed that there will be employed the same unit of concentration that was used in the experimental derivation of K_a, and also the conventional form of the ratio with the ions in the numerator.

Since equation (4) deals with the active masses of the reactants it is a special application of the so-called law of mass action which states that the velocity of a reaction is proportional to the product of the *concentrations* of the reactants.

Using equation (4) for a particular acid it will be seen by inspection of the equation that if $[H^+]$ is increased, as by the addition of another acid, there must be a readjustment of either $[A^-]$ or [HA] or both to keep K_a constant. Likewise if $[A^-]$ should be increased by the addition of a highly dissociating salt of the acid in question, there would be a readjustment of either $[H^+]$ or [HA] or both to keep K_a constant. Thus the independent alteration of the concentration of any one of the species included in the equilibrium equation causes a displacement of the equilibrium to a new position. This illustrates how difficult it is to keep track of the affair unless use is made of the simple algebraic relations.

If the acid alone be present, $[H^+] = [A^-]$. Substituting $[H^+]$ for $[A^-]$ and solving equation (4) for $[H^+]$ we have

$$[H^+] = \sqrt{K_a\,[HA]}$$

If the acid is so weak that practically all is in the undissociated form, no great error is made in putting [HA] equal to the concentration [S] of the total acid. Then $[H^+] = \sqrt{K_a[S]}$.

In general it can be shown that for any reaction such as

$$A + B + C + \ldots\ldots \rightleftharpoons A' + B' + C' + \ldots\ldots$$

the equilibrium condition is:

$$\frac{[A]\,[B]\,[C]\ \ldots\ldots}{[A']\,[B']\,[C']\ \ldots\ldots} = k$$

From the assumptions introduced in the argument it is evident that the equilibrium constant will hold good only so long as there are maintained constant those physical conditions which affect the velocity of a reaction in one direction or the reverse. A change in temperature will alter the "constant," but not to such an extent as will a change in solvent. With due regard for such matters we may regard the equilibrium constant as a number characteristic of a given reaction at the equilibrium state.

In the derivation of the equilibrium equation we have employed as an example the electrolytic dissociation of an acid. We may

now state that all substances capable of yielding hydrogen ions must be considered as having an acidic nature and their conduct in solution must be governed by the equilibrium equation.

With the ionization constant defined we are prepared to give quantitative significance to comparative "strengths" among acids. Inspection of equation (4) shows at once that if K_a is large the numerator of the left hand side must be large in relation to the denominator. In other words an acid having a relatively high K_a value will, if left to itself in solution, tend toward a high degree of dissociation. A given over-all concentration of an acid with high dissociation constant will furnish a higher concentration of hydrogen ions than will the same over-all concentration of an acid with low dissociation constant. Thus the value of K_a at once indicates the "strength" of an acid so far as "strength" is measurable in terms of ionization.

In the following table are given a few dissociation constants of acids and also of bases.

TABLE 1

Showing acidic and basic dissociation constants and their relation to a rough classification of acids and bases

CLASS	COMPOUND	DISSOCIATION CONSTANT
Strong acid	Hydrochloric	Not well defined
Moderately strong acid	Oxalic (first H)	1.1×10^{-1}
Weak acid	Acetic	1.8×10^{-5}
Very weak acid	Boric	6.5×10^{-10}
Strong base	Sodium hydroxid	Not well defined
Weak base	Ammonium hydroxid	1.8×10^{-5}
Very weak base	Aniline	4.6×10^{-10}

The dissociation of bases will now be considered. Just as a substance ionizing to give hydrogen ions is called an acid so a substance which ionizes to give hydroxyl ions (OH^-) is called a base.

The reversible reaction $NaOH \rightleftarrows Na^+ + OH^-$ may be written as $BOH \rightleftarrows B^+ + OH^-$ where B represents any monovalent metal. This reaction may be treated in precisely the same way that reaction (1) was treated. The equilibrium condition is:—

$$\frac{[B^+]\,[OH^-]}{[BOH]} = K_b \qquad (5)$$

Just as the value of K_a is characteristic of a given acid so is the value of K_b characteristic of a given base.

A very important relationship between acids and bases in aqueous solution is brought about by the conduct of water. It dissociates into the hydrogen ion (H^+) characteristic of acids and the ion characteristic of bases, OH^-, called the hydroxyl ion. The equilibrium of the reversible reaction $HOH \rightleftarrows H^+ + OH^-$ is represented by

$$\frac{[H^+]\,[OH^-]}{[HOH]} = k$$

Because the concentration of the undissociated water is so large in relation to the dissociation products, [HOH] will not be changed appreciably by the slight dissociation. [HOH] may therefore be considered a constant and combined with k. Then the above equation becomes:

$$[H^+]\ [OH^-] = K_w. \tag{6}$$

It follows from this equation that, no matter how concentrated the hydroxyl ions may be, there must remain sufficient hydrogen ions to satisfy the above relation.[2] This permits us to speak of the hydrogen ion concentration of alkaline solutions and, as will be shown presently, to construct a scale of acidity-alkalinity in which we do not discriminate between hydrogen and hydroxyl ion concentration.

Starting from equations (4), (5) and (6), applying certain approximations and then using graphic methods of presentation we can present a generalized picture of the conduct of acids and bases similar to that first used by Henderson (1908). The final simplicity of the picture warrants what may at first appear to be a complicated reconstruction of the above equations.

In order to emphasize the hydrogen ion concentration as the quantity in equation (4) with which the other species keep in adjustment, let us rewrite equation (4) as follows:

$$\frac{1}{[H^+]} = \frac{[A^-]}{K_a[HA]}$$

[2] $K_w = 10^{-14}$. If in an alkaline solution the concentration of hydroxyl ions is 0.01 normal (10^{-2}), $[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{10^{-2}} = 10^{-12}$ N.

We choose the form which will give the reciprocal of $[H^+]$ because we shall have to make use of the logarithm of this value under the symbol pH for reasons which will appear later. For the present let it be granted that it will be found convenient to use $\log \frac{1}{[H^+]}$ rather than $[H^+]$. Taking the logarithm of each side of the above equation we have

$$\log \frac{1}{[H^+]} = \log \frac{1}{K_a} + \log \frac{[A^-]}{[HA]} \qquad (7)$$

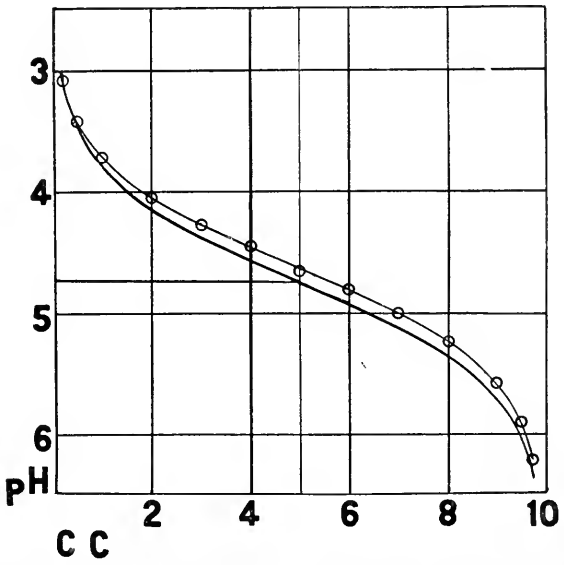


FIG. 1. COMPARISON OF EXPERIMENTAL TITRATION CURVE OF ACETIC ACID WITH THEORETICAL APPROXIMATION

With the use of this equation we can chart some important relationships. Let it first be applied to what may be called "titration curves."

Suppose we titrate 10 cc. of 0.2N acetic acid with 0.2N sodium hydroxid. Ordinarily no attention would be given to the state of the solution until the so called "end point" of the titration were reached. In the present instance we shall follow the course of the titration from the beginning by determining after each addition of alkali the hydrogen ion concentration.

The experimental curve is plotted in figure 1. Let us compare it with the values obtained by the use of equation (7).

In the first place acetic acid is classed among the moderately weak acids. Its dissociation constant as given in Landolt-Börnstein is 1.82×10^{-5} at 18°C. Hence $\log \frac{1}{K_a} = 4.74$. Be-

TABLE 2

Comparison of log $1/[H^+]$ *for acetic acid-sodium acetate calculated by means of the approximation formulated in equation (8) and determined experimentally by Walpole*

N/5 NaOH	RATIO $\frac{[\text{SALT}]}{[\text{ACID}]}$	LOG RATIO	LOG 1/Ka	LOG $1/[H^+]$ CALCULATED	LOG $1/[H^+]$ WALPOLE
cc.					
0.20	9.00	−1.69	4.74	3.05	3.08
0.25	0.020	−1.59	4.74	3.15	3.15
0.30	0.026	−1.51	4.74	3.23	3.20
0.40	0.031	−1.38	4.74	3.36	3.32
0.50	0.042	−1.28	4.74	3.46	3.42
0.75	0.053	−1.09	4.74	3.65	3.59
1.0	0.081	−0.95	4.74	3.79	3.72
2.0	0.111	−0.60	4.74	4.14	4.05
3.0	0.250	−0.37	4.74	4.37	4.27
4.0	0.429	−0.18	4.74	4.56	4.45
5.0	0.667	0.00	4.74	4.74	4.63
6.0	1.000	+0.18	4.74	4.92	4.80
7.0	1.500	+0.37	4.74	5.11	4.99
7.5	2.33	+0.48	4.74	5.22	5.09
8.0	3.00	+0.60	4.74	5.34	5.23
8.5	4.00	+0.75	4.74	5.49	5.37
9.0	5.67	+0.95	4.74	5.69	5.57
9.5	19.00	+1.28	4.74	6.02	5.89
9.625	25.67	+1.41	4.74	6.15	6.02
9.75	39.00	+1.59	4.74	6.33	6.21
9.875	79.00	+1.90	4.74	6.64	6.52

cause of the small dissociation of acetic acid (less than 2 per cent in 0.2N solution even with no acetate present) the concentration of the undissociated residue [HAc] is *approximately* equal to the concentration of the total acetic acid. It is characteristic of the alkali salts of acids that they are very highly dissociated. Therefore, when sodium hydroxid is added to the acetic acid solution, the resulting sodium acetate furnishes the greater amount of the

total acetate (Ac^-) ions. As an approximation therefore we may substitute for the ratio $\frac{[A^-]}{[HA]}$ in equation (7) the ratio $\frac{[\text{salt}]}{[\text{acid}]}$. Equation (7) then becomes:

$$\log \frac{1}{[H^+]} = \log \frac{1}{K_a} + \log \frac{[\text{salt}]}{[\text{acid}]}. \qquad (8)$$

In table 2 are given the ratios $\frac{[\text{salt}]}{[\text{acid}]}$ calculated from the number of cubic centimeters of 0.2N alkali added to 10 cc. of 0.2N acetic acid. Then follow the logarithms of these ratios, the value of $\log \frac{1}{K_a}$ for acetic acid, and $\log \frac{1}{[H^+]}$ calculated from these data by means of equation (8). Finally in the last column are given the values of $\log \frac{1}{[H^+]}$ calculated by Walpole (1914) from his hydrogen electrode measurements. The experimental values $pH = \log \frac{1}{[H^+]}$ are plotted in figure 1 as circles while the values calculated by means of the approximation equation (8) are on the unmarked line. There is evidently a substantial agreement with a more or less regular discrepancy which remains to be explained. The discrepancy may be ascribed in part to the assumption that the salt is wholly dissociated and that it is entirely responsible for the anions of equation (7). If there be applied a correction for the partial dissociation of the acetate, there is obtained a much closer agreement.

But even this correction does not take into consideration certain minor points, and it leaves untouched both the accuracy with which K_a has been determined and the comparability of data obtained by widely different methods which are often applied (sometimes uncritically) in making such calculations as those indicated above.

We shall proceed with the approximate treatment to bring out certain more general relations, and shall leave to Chapter XXI their further application to ordinary titrations.

In equation (8) when the ratio $\frac{[\text{salt}]}{[\text{acid}]}$ equals one, $\log \frac{1}{[H^+]} = \log \frac{1}{K_a}$. Then $[H^+] = K_a$.

In other words the middle portion of the titration curve of a particular acid lies at ("near" if we are to be strict) a point where the hydrogen ion concentration is numerically equal to the dissociation constant.[3]

Thus if one wishes a solution of $[H^+] = 1 \times 10^{-5}$, an acid with dissociation constant close to this value is selected and mixed with the proper amount of its alkali salt.

Or to look at the matter from another point of view, if we determine the half transformation point in the titration of a weak acid, we know approximately the dissociation constant of the acid.

A similar set of relationships can be constructed for bases.

Instead of putting the fundamental equation (4) into the form which we have utilized in following titration curves it is sometimes advantageous to use the following development.

Transforming (4) we have:

$$\frac{[A^-]}{[HA]} = \frac{K_a}{[H^+]}$$

Now let us represent the concentration of the total acid by [S]. Then the concentration of [HA] will be:

$$[HA] = [S] - [A^-]$$

$$\frac{[A^-]}{[S]-[A^-]} = \frac{K_a}{[H^+]}$$

or

$$\frac{[A^-]}{[S]} = \frac{K_a}{K_a + [H^+]}$$

The ratio $\frac{[A^-]}{[S]}$ is the ratio of the dissociated acid to the total acid present in the solution. This ratio may be represented by α. Hence,

$$\alpha = \frac{K_a}{K_a + [H^+]} \qquad (9)$$

[3] There is implied in this the maintenance of the customary unit of concentration. Cf. page 18.

Since we are interested in $\log \frac{1}{[H^+]}$ or pH rather than $[H^+]$, because of the resultant simplification of chart representations and because of other reasons which will appear later, we may recast equation (9) and taking the logarithm of each side we have:

$$\log \frac{1}{[H^+]} = \log \frac{1}{K_a} + \log \frac{\alpha}{(1 - \alpha)} \tag{10}$$

Plotting $\log \frac{1}{[H^+]}$, which is pH, against α, and expressing α as *percentage* dissociation, there is obtained a curve such as A or B in figure 2. Such curves are identical in form, the form being determined by the ratio $\frac{\alpha}{(1 - \alpha)}$. Their position on the pH axis is determined by the value of the dissociation constant in the expression $\log \frac{1}{K_a}$.

Since (10) is useful in plotting type curves a table of values for $\log \frac{\alpha}{1 - \alpha}$ is given in the appendix (p. 460).

In a similar way we arrive at the relation for bases:

$$\alpha = \frac{K_b}{K_b + [OH^-]} \tag{11}$$

or

$$\log [OH^-] = \log \frac{K_b (1 - \alpha)}{\alpha}. \tag{12}$$

But since we wish to deal uniformly with $\log \frac{1}{[H^+]}$, which is pH, rather than with the hydroxyl ion concentration or any direct function thereof, we shall introduce the water equilibrium, equation (6). Then (12) becomes

$$\log \frac{K_w}{[H^+]} = \log \frac{K_b (1 - \alpha)}{\alpha}$$

or

$$pH = \log \frac{1}{[H^+]} = \log \frac{K_b}{K_w} + \log \frac{(1 - \alpha)}{\alpha} \tag{13}$$

With the introduction of K_w, the dissociation constant of water, into our equations it becomes advisable to consider its numerical value. K_w has been determined in a variety of ways of which the following are examples. Kohlrausch and Heydweiller (1894) determined the electrical conductivity of extremely pure water. Assuming that the conductance is proportional to the mobility of

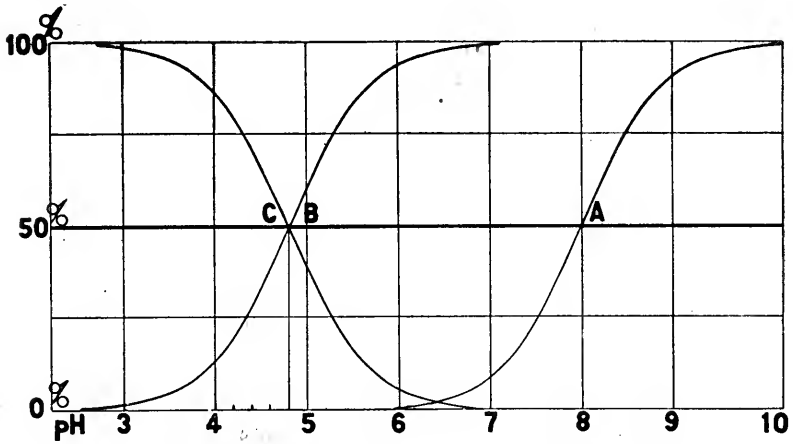


FIG. 2. DISSOCIATION CURVES AND DISSOCIATION-RESIDUE CURVES

A. Dissociation curve of acid, $\log \frac{1}{K_a} = 8.0$.

B. Dissociation curve of acid, $\log \frac{1}{K_a} = 4.8$.

C. Dissociation-residue curve of acid, $\log \frac{1}{K_a} = 4.8$, or dissociation curve of a base $\log \frac{1}{K_b} = \log \frac{1}{K_w} - 4.8$.

the hydrogen and the hydroxyl ions, and that these are present in equal concentrations, their product is found to be 1.1×10^{-14}. The hydrolysis of methyl acetate having been found to be proportional to the concentration of hydroxyl ions, Wijs (1893) determined the hydrolysis by water and found $K_w = 1.44 \times 10^{-14}$.

By determining the hydrogen ion concentration with the hydrogen electrode in solutions of known hydroxyl ion concentration (as determined by conductance measurements), K_w is obtained from the product of the concentrations of the two ions.

By this method Lewis, Brighton and Sebastian (1917) found the value 1.012×10^{-14} at 25°C.

Kolthoff (1921) has compiled the following table showing the dissociation constant of water at different temperatures as given by different authors and methods:

TEMPERATURE	I	II	III	IV
0°	0.12×10^{-14}	0.14×10^{-14}		0.089×10^{-14}
18°	0.59×10^{-14}	0.72×10^{-14}	0.74×10^{-14}	0.46×10^{-14}
25°	1.04×10^{-14}	1.22×10^{-14}	1.27×10^{-14}	0.82×10^{-14}
50°	5.66×10^{-14}	8.7×10^{-14}		
100°	58.2×10^{-14}	74.0×10^{-14}		48.0×10^{-14}

I. Kohlrausch and Heydweiller recalculated by Heydweiller (1909).
II. Lorenz and Böhi (1909).
III. Michaelis (1914).
IV. Noyes and coworkers (1907).

The following values of $\log \frac{1}{K_w}$ given by Michaelis (1914) were obtained on a somewhat different basis from that used by Lewis, Brighton and Sebastian (1917).

Since in pure water $[H^+] = [OH^-]$, $[H^+]$ or $[OH^-] = \sqrt{K_w}$. Hence from the datum of Lewis, Brighton and Sebastian the normality of H^+ or OH^- in pure water at 25°C. is $\sqrt{K_w} = 1.006 \times 10^{-7}$ (practically pH = 7.0).

In the following pages wherever we have occasion for purposes of illustration to use a numerical value for K_w we shall employ the rounded value 10^{-14}.

Introducing the numerical value of K_w into equation (13) we have the convenient form:

$$pH = 14 - \log \frac{1}{K_b} + \log \frac{(1 - \alpha)}{\alpha} \qquad (14)$$

In figure 2 we have plotted α as percentage dissociation. It is obvious that the percentage dissociation residue will give the complement of the dissociation curve and will cross any particular one of these at the fifty per cent dissociation point. See, for example, the curve C of figure 2.

Now by comparing equation (10) with equation (14) it is found that the curve for the dissociation-residue of an acid is identical with the curve for the dissociation of a base when K_a of the acid is related to K_b of the base as $\log \frac{1}{K_a} = 14 - \log \frac{1}{K_b}$. In other

TABLE 3

TEMPERATURE	$\log \frac{1}{K_w}$	pH OF NEUTRAL POINT
16	14.200	7.10
17	14.165	7.08
18	14.130	7.07
19	14.100	7.05
20	14.065	7.03
21	14.030	7.02
22	13.995	7.00
23	13.960	6.98
24	13.925	6.96
25	13.895	6.95
26	13.860	6.93
27	13.825	6.91
28	13.790	6.90
29	13.755	6.88
30	13.725	6.86
31	13.690	6.85
32	13.660	6.83
33	13.630	6.82
34	13.600	6.80
35	13.567	6.78
36	13.535	6.77
37	13.505	6.75
38	13.475	6.74
39	13.445	6.72
40	13.420	6.71

words curve C (fig. 2) is either the dissociation-residue curve of an acid for which $\log \frac{1}{K_a} = 4.8$ or the dissociation curve of a base for which $\log \frac{1}{K_b} = 9.2$ (since $14 - 9.2 = 4.8$).

The importance of this relation lies in the fact that a determination of the effect of hydrogen ion concentration on some process may not reveal whether the phenomenon has to do with

an acid or a base, unless an independent method reveals the nature of the active substance.

The student will find it interesting to plot dissociation curves for acids with percentage dissociation as one coördinate and pH as the other, and then dissociation curves for bases with $\log \frac{1}{[OH^-]}$ (which may be called pOH) as one of the coördinates plotted inversely as pH. At a given temperature and given value for K_w there is a fixed value for pOH at each value for pH. This follows directly from equation (6); and it is particularly to be noted that in deriving this relation we need not fix the position of the pOH scale in its relation to the pH scale by confining our attention to the special case where $[H^+] = [OH^-]$, occurring roughly at pH 7.0. Indeed the so-called neutral point (pH 7.0) may be considered only as a convenient, mental reference point having comparatively little physical significance. It is not the point to which titrations are led, except under the rare condition that the acid and the base are of exactly equal strength; and it is of far less importance for amphoteric electrolytes than is the isoelectric point of the given ampholyte.

Having plotted the two systems mentioned above the student will find it interesting to assume that for moderate variations of temperature the dissociation constants of acids and bases do not change seriously, and then to note the shift in the two systems relative to one another when K_w is altered with temperature.

The treatment accorded simple acids and bases may be extended to poly-acidic acids and poly-basic bases as well as to those compounds containing both acidic and basic groups which are called amphoteric electrolytes. It seems to be true very often for such compounds that they dissociate in steps as is illustrated in the titration curve of the tri-acidic phosphoric acid shown on page 41. In this, as in many other cases, the several dissociation constants are of such widely different magnitudes that, when we plot the dissociation curves as if of separate acids possessing these dissociation constants, the curves do not seriously overlap.

Such acids may therefore be treated as if composed of two or more independent acids. The effect produced when two dissociation constants lie closer together is illustrated by the titration curve of o-phthalic acid shown on page 273. If in this case the formal dissociation curve of a simple acid be plotted over the main position of each section of the phthalate curve, it will be found (as shown by Acree) that the experimental curve follows very closely the interpolated resultant of the two formal single curves.

For amphoteric electrolytes (i.e., electrolytes containing acidic *and* basic groups) a relation of great importance to protein chemistry may be illustrated by the conduct of the simple ampholyte, p-amino benzoic acid. The acid dissociation constant K_a is 6.8×10^{-6} and the basic dissociation constant K_b is 2.3×10^{-12} (Scudder). Translating these into the corresponding pH values we have 5.17 and 2.36. If we regard the compound as if it were made up of an acid and a base with the above dissociation con-

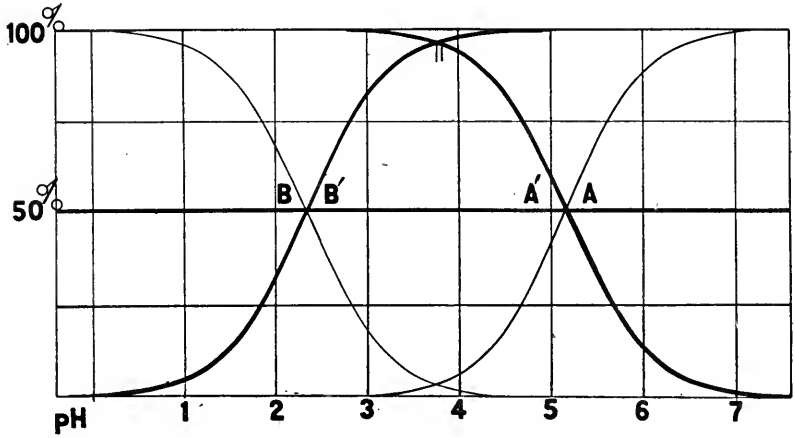


FIG. 3. DISSOCIATION AND DISSOCIATION-RESIDUE CURVES OF p-AMINO-BENZOIC ACID

Treated as if the amphoteric electrolyte were composed of an acid of $\log \frac{1}{K_a} = 5.17$ and a base of $\log \frac{1}{K_b} = \log \frac{1}{K_w} - 2.36$.

stants (in terms of pH) and each independent of the other, we can plot the dissociation curves of each with the aid of equations (10 and 14). In each case the dissociation-residue curves are the complements. These are plotted in figure 3 with heavy lines. It is seen that they cross at pH = 3.77. This means that at pH = 3.77 there is a maximum of undissociated residue. Now if the salts are more soluble than the free compound itself there should be a minimum solubility at pH 3.77. Michaelis and Davidsohn (1910) found a minimum solubility at pH 3.80.

Turning again to the light lines A and B of figure 3, we see that their intersection is at a point where the percentage of the com-

pound ionized as an anion is equal to the percentage ionized as a cation. In other words the amount carrying a negative charge is equal to the amount carrying a positive charge. Because of this equality the point where it occurs is called the *isoelectric point.*

If we still maintain the simple conditions postulated in this elementary treatment, we can calculate the isoelectric point from the dissociation constants of an amphoteric electrolyte.

Consider an amphoteric electrolyte of the type HROH for which we have the following equilibrium equations:

$$\frac{[HR^+]\,[OH^-]}{[HROH]} = K_b \tag{15}$$

$$\frac{[R\bar{O}H]\,[H^+]}{[HROH]} = K_a \tag{16}$$

When $[HR^+] = [R\bar{O}H]$ (isoelectric condition)

$$K_b \frac{[HROH]}{[OH^-]} = K_a \frac{[HROH]}{[H^+]}$$

Hence

$$[H^+] = \sqrt{\frac{K_a}{K_b} K} \tag{17}$$

In the case cited above $[H^+] = \sqrt{\frac{6.8 \times 10^{-6}}{2.3 \times 10^{-12}} 10^{-14}}$

or $pH = \log \frac{1}{[H^+]} = 3.77$

Furthermore from equations (15) and (16)

$$[HR^+] + [ROH^-] = K_b \frac{[HROH]\,[H^+]}{K_w} + K_a \frac{[HROH]}{[H^+]}$$

If we let $[HR^+] + [ROH^-] = X$, X becomes a minimum when

$\frac{dX}{d[H^+]} = 0$, a condition fulfilled when $[H^+] = \sqrt{\frac{K_a}{K_b} K_w}$

In other words the sum of the anion and cation concentrations is a minimum at the isoelectric point.

Only in case $K_a = K_b$ will the isoelectric point correspond with the "neutral point."

It is at once evident that the isoelectric point of an amphoteric electrolyte is a point at or near which there should tend to occur maximal or minimal properties of its solution. Indeed at such points have been found to occur minimum solubilities, minimum viscosities, minimum swelling, optimum agglutinations, etc.

It should be emphasized that the foregoing relationships have been developed from very simple conditions. When these conditions have been approached experimental verification has been found. The insight thus gained has led to a better understanding of complex ampholytes, the complete equilibria of which can be seen only in broad outline. In attempting to formulate more precisely the equilibrium equations which hold under more complex conditions than those postulated above, Michaelis (1920) has started with the influence of uni-univalent salts upon a simple ampholyte and has then extended his propositions to cover the influence of divalent ions and the influence of micelle formation. It is of special interest to note that he can account for the displacement of the precipitation optimum from the isoelectric point by the influence of salts and that he finds it necessary to caution against considering the isoelectric point to be always identical with the point of maximum dissociation residue. He also outlines the direction in which various relations will be modified by the aggregation of the undissociated ampholyte into micelles.

SUPPLEMENTARY REFERENCES

Texts on the principles of electrolytic dissociation: LeBlanc, Jones, Nernst, Ostwald, Stieglitz (1917).

Generalized relations among acids and bases: Henderson (1908), Michaelis (1914, 1922), Sørensen (1912).

CHAPTER II

Some Special Aspects of Acid-Base Equilibria

Words are the footsteps of reason.—Francis Bacon.

In the foregoing chapter we have outlined the chief aspects of acid-base equilibria. We now have to discuss in more detail some of the terminology of special use in acid-base studies and also certain important matters which are continually met in dealing with that class of electrolytes called the "strongly dissociating" acids, bases and salts.

the pH scale

When "acidity" was resolved into its two components the normality unit was retained for each. As a normal solution of an acid had been defined as one containing in 1 litre of solution the equivalent of 1 gram atom of acidic hydrogen, so the normal solution of the hydrogen ion was defined to be one containing in 1 litre of solution 1 gram atom of hydrogen ions.[1]

To distinguish between these two components with their common unit it has been suggested that we call "normality" in its older sense the *quantity* factor of "acidity" and the hydrogen ion concentration the *intensity* factor. This may serve to emphasize a distinction, but the suggested analogy with the quantity and intensity factors of energy is confusing when we retain for each a unit of the same category. Nevertheless the two components remain in a restricted sense the quantity and intensity factors of "acidity." The one is the total quantity of available acid. The second, the concentration of the hydrogen ions, represents the real intensity of "acidity" whenever it is the hydrogen ion which is the more directly active participant in a reaction. This is admirably expressed when we use for hydrogen ion concentrations a mode of expression which links it with the *potential* of a hydrogen electrode. It so happens that in determining the hydrogen

[1] It makes little difference whether the atomic weight of hydrogen be taken as 1.008 or as 1.0 in calculating $[H^+]$.

ion concentration with the hydrogen electrode the potentials of this electrode are put into an equation which reduces to the form:

$$\frac{\text{Potential}}{\text{Numerical factor}} = \log \frac{1}{[H^+]}$$

Thus $\log \frac{1}{[H^+]}$ is at once obtained by the most simple of calculations. Sørensen (1909) saw that this value serves to define a hydrogen ion concentration quite as well as $[H^+]$ itself and in his *Enzyme Studies II*, he used this mode of expression and gave to $\log \frac{1}{[H^+]}$ the symbol P_H^+.

As a matter of typographical convenience[2] we shall adopt pH in place of P_H^+. Since this is coming into wide usage its uniform adoption is recommended in place of the bothersome variations[3] which have made their way into the literature.

Although Sørensen has not revealed the considerations which led to the choice of the letter P in his symbol, we might regard P as suggesting the potential (intensity) factor of acidity in the sense described above.

Writing the potential equation given on page 154 as

$$W = EF = RT\ln \frac{1}{[H^+]}$$

where

it will be seen that E is the intensity factor in the work required to carry a gram atom of hydrogen ions from concentration $[H^+]$ to concentration 1 normal; and pH is a linear function of E.

pH is sometimes called the Sørensen value or Sørensen unit and following Sørensen's original suggestion it is named the hydrogen ion exponent. The last mentioned name must be used with some caution because of a difference in sign between a given pH value and the exponent occuring when the normality of the corresponding hydrogen ion concentration is written. For

[2] As is the custom of the *Journal of Biological Chemistry*.

[3] Certain punctilious authors have insisted that the original symbol should be retained but have made the mistake of assuming it to be P_H. The following variations are found in the literature:

ph,pH,Ph,PH,P_h,P^H,P_h^+,P_H^+, also each case italicised.

examples -7 is the exponent in 10^{-7}, but the pH value corresponding to $[H^+] = 10^{-7}$N is $+7$.

The convenience of pH over $[H^+]$ is manifest when we compare the numerical values encountered in chemical and physiological studies. For instance, one enzyme may operate most actively at a hydrogen ion concentration of 0.01 normal while another is most active at 0.000,000,001 normal. While convenient abbreviations of such unwieldy values are 1×10^{-2} and 1×10^{-9}, there remains the difficulty of plotting such values on ordinary cross-section paper. If the difference between 0.000,000,001 and 0.000,000,002 is given a length of one millimeter, the difference 0.01 to 0.02 when plotted on the same scale would be ten kilometers, ten kilometers distant. Evidently the logarithmic spacing should be followed and fortunately it is the logarithmic plotting of hydrogen ion concentration (in terms of pH) which correctly depicts the fact that the difference between 1×10^{-9} and 2×10^{-9} may be as important for one set of equilibria as the enormously greater difference between 1×10^{-2} and 2×10^{-2} is for another set of equilibria. This is revealed in the charts on previous and subsequent pages.

Thus both convenience and the nature of the physical facts compel us directly or indirectly to operate with some logarithmic function of $[H^+]$.

It is unfortunate that a mode of expression so well adapted to the treatment of various relations should conflict with a mental habit. $[H^+]$ represents the hydrogen ion concentration, the quantity usually thought of in conversation when we speak of increases or decreases in acidity. pH varies inversely as $[H^+]$. This is confusing.

The normality mode of expression has historical priority and consequently conventional force. Since there is a hydrogen ion concentration for each hydroxyl ion concentration it became the custom, following Friedenthal (1904), to express both acidities and alkalinities in terms of $[H^+]$. This gave a scale of one denomination and the meaning of "higher" and of "lower" became firmly fixed. Now we meet the new scale with its direction reversed. The inconvenience is unquestionable and very largely because of it the pH scale has been criticized.

See the discussion in the Journal of the Washington Academy of Sciences by Wherry and Adams (1921) and by Clark (1921). Wherry's (1919) chief object is to establish a scale of convenient direction but in doing so he gains a superficial advantage at the expense of several simple and very important experimental and theoretical relations which he has not taken into consideration.

In Chapter XIX there will be advanced a reason for adhering to the use of the pH introduced by Sørensen; but at this point it may be well to say that in both of the two chief methods of determining hydrogen ion concentration we encounter physical relations which make the errors proportional to pH rather than to $[H^+]$. Furthermore, pH is the more directly related to certain electrode phenomena which are partially dependent upon hydrogen ion concentration and therefore pH is useful in dealing with subjects outside the strict limits of hydrogen electrode measurements.

The gross relation of $[H^+]$ to pH is shown in the following table. See also table B appendix.

$[H^+]$	pH	$[H^+]$	pH
10^{-0}	0	10^{-8}	8
10^{-1}	1	10^{-9}	9
10^{-2}	2	10^{-10}	10
10^{-3}	3	10^{-11}	11
10^{-4}	4	10^{-12}	12
10^{-5}	5	10^{-13}	13
10^{-6}	6	10^{-14}	14
10^{-7}	7		

The following symbols indicating hydrogen ion concentration in normality are encountered in the literature $[H^+]$; $[H^\cdot]$; $C_H{}^+$; C_H; h.

THE EFFECT OF DILUTION

A litre of normal acid becomes a fifth normal solution if diluted to 5 litres; the hydrogen ion concentration may in many instances be affected too little for the change to be detected by any but refined methods. This apparent anomaly is frequently encountered and sometimes advantage of it is taken in the dilution of solutions otherwise too dense optically for the application of the indicator method. The effect of dilution upon the hydrogen ion concentration of a solution may be briefly generalized by some approximations.

Consider an acid of the type HA for the dissociation of which we have the equilibrium equation:

$$\frac{[H^+] \times [A^-]}{[HA]} = K_a$$

If K_a is small there must obviously be a large reserve of undissociated acid so long as the concentration of total acid is high. As the solution is diluted this reserve dissociates to keep K_a constant; but there is a readjustment of all components which can be conveniently followed only by means of the simple algebraic equation expressing the equilibrium condition.

If the acid alone is present in the solution we may assume that $[A^-] = [H^+]$. Also if S_a = the total acid, $[HA] = S_a - [H^+]$.

Substituting these in the above equation and solving for $[H^+]$ we have:

$$[H^+] = \sqrt{K_a S_a + \frac{K_a^2}{4}} - \frac{1}{2} K_a \qquad (18)$$

When K_a is small in relation to S_a

$$[H^+] \cong \sqrt{K_a S_a} \qquad (19)$$

Compare the equation on page 19. On these assumptions the hydrogen ion concentration should vary with dilution of the solution (diminution of S_a) only as the square root of $K_a S_a$.

If there is present a salt of the acid we can apply the equation derived on page 24 which shows that the hydrogen ion concentration of a mixture of a weak acid and its highly dissociated salt is determined approximately by the ratio of acid to salt. Since dilution does not change the ratio, such a mixture should not suffer a change of hydrogen ion concentration beyond the narrow limits set by the approximate treatment with which this relation was derived.

Therefore, except for solutions of high hydrogen ion concentration induced by the presence of unneutralized strong acids, the hydrogen ion concentration should vary with dilution somewhere between the zero change indicated by the last approximation and the square root relation first indicated.

Such a conclusion takes no account of changes of equilibrium which sometimes occur in colloidal solutions.

For bases and amphoteric electrolytes similar relations may be deduced. One or two actual cases may be of interest.

Sørensen has given the following table of the pH values of different dilutions of asparagine and glycocoll.

MOLECULAR CONCENTRATION OF GLYCOCOLL	pH	MOLECULAR CONCENTRATION OF ASPARAGINE	pH
1.0	6.089	1.0	2.954
0.1	6.096	0.1	2.973
0.01	6.155	0.01	3.110
0.001	6.413	0.001	3.521
0.0001	6.782	0.0001	4.166

The dilution here is ten-fold at each step, yet the increase in pH is very small while the solutions are beween 1.0-0.01 M.

Walpole (1914) besides giving data on the hydrogen electrode potentials of various dilutions of acetic acid and "standard acetate," has determined the effect of a twenty-fold dilution of various acetic acid-sodium acetate mixtures. The change of pH on twenty-fold dilution of standard acetate is about 0.08 pH; and of mixtures of acetic acid and sodium acetate which lie on the flat part of the curve the change of pH is of the same order of magnitude. When the ratio $\frac{\text{acetic acid}}{\text{sodium acetate}}$ reaches 19/1 the change is about 0.3 pH.

BUFFER ACTION

If we were to add to 1 liter of perfectly pure water of pH 7.0, 1 cc. of 0.01N HCl, the resulting solution would be about pH 5.0 and very toxic to many bacteria. If, on the other hand, we were to add this same amount of acid to a liter of a standard beef infusion medium of pH 7.0, the resulting change in pH would be hardly appreciable. This power of certain solutions to resist change in reaction was commented upon by Fernbach and Hubert (1900) who likened the resistance of phosphate solutions to a "tampon." The word was adopted by Sørensen (1909) and in the German rendition of his paper it became "Puffer" and thence the English "buffer." There has been some objection to this

word so applied but it now possesses a clear technical meaning and is generally used. By buffer action we mean the resistance exhibited by a solution to change in pH through the addition or loss of acid or alkali. This may be illustrated by titration curves such as those shown in figures 4, 5 and 6. The construction of such curves may be illustrated by the following example.

A 1 per cent solution of Witte peptone was found to have a pH value of 6.87. To equal portions of the solution were added successively increasing amounts of 0.1N lactic acid and the resulting pH was measured in each case. There were also added to equal portions of the solution successively increasing amounts of 0.1N NaOH and the resulting pH was measured in each case. The pH values were then plotted on cross section paper as ordinates against the amount of acid or alkali added in each case as abscissas. This gave curve 1 shown in figure 4. The other curve shown in this figure was constructed with data obtained with a 5 per cent solution of Witte peptone. The curves of figures 5 and 6 were obtained in a similar way.

These curves illustrate the following points.

Figure 4 shows that the buffer action of a solution is dependent upon the concentration of the constituents. The 5 per cent solution is much more resistant to change in pH than the 1 per cent solution. It will also be noticed that in either case the buffer action is not the same at all points in the curve. In other words the buffer action can not be expressed by a constant but must be determined for each region of pH. This is illustrated even more clearly by the titration curve for phosphoric acid (fig. 5). At the point where the solution contains only the primary phosphate and again where it contains only the secondary phosphate there is very little buffer effect indeed.

Furthermore the buffer action of a solution may not be due entirely to the nature of the constituents titrated but also to the nature of the substance with which it is titrated. This point may be illustrated by titrating a beef infusion medium in the one case with hydrochloric acid and in the other case with lactic acid, both of the same normality (see fig. 6). It will be seen that at first the two curves are identical. As the region is approached where the dissociation of the "weak" lactic acid is itself suppressed because of the accumulation of lactate ions and the high

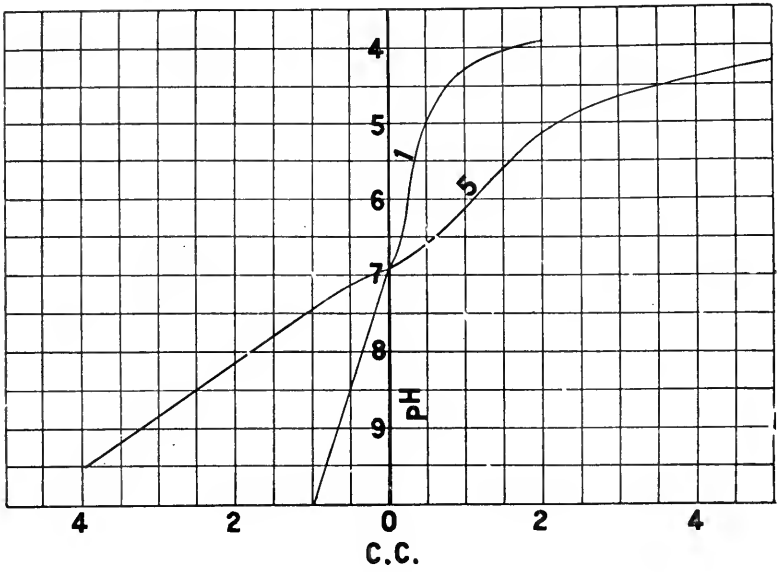


FIG. 4. TITRATION CURVES OF 1 PER CENT AND 5 PER CENT PEPTONE

Ten cubic centimeters of peptone solution titrated with N/10 lactic acid (to right) and with N/10 NaOH (to left).

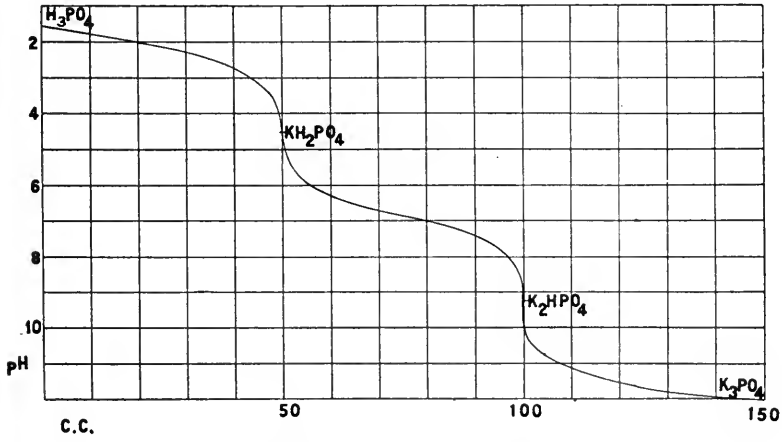


FIG. 5. TITRATION CURVE OF PHOSPHORIC ACID

Fifty cubic centimeters M/10 H_3PO_4 titrated with N/10 KOH.

concentration of the hydrogen ions, further addition of this acid has comparatively little effect. The strong hydrochloric acid on the other hand continues to be effective until its dissociation, too, at very high hydrogen ion concentrations is suppressed.

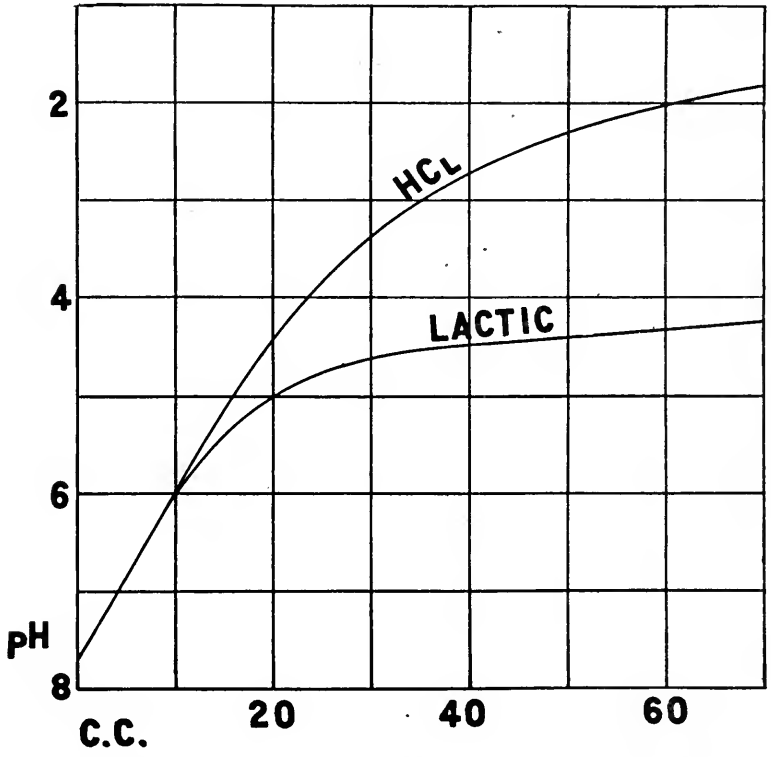


FIG. 6. TITRATION CURVES OF A BEEF INFUSION MEDIUM

One hundred cubic centimeters medium titrated with N/5 HCl and with N/5 lactic acid.

These examples will suffice to make it evident that the buffer action of a solution is dependent upon the nature and the concentration of the constituents, upon the pH region where the buffer action is measured and upon the nature of the acid or alkali added. To connect all these variables is a difficult problem. Koppel and Spiro (1914) have attempted to do so but they have necessarily had to leave out of consideration another factor. If

there are present any bodies which tend to adsorb any of the constituents of a solution which can affect the hydrogen ion concentration of a solution, these bodies will tend to act as buffers or will affect the buffer action of the solution. Henderson (1909) has called attention to this and Bovie (1915) has shown in a very interesting way the buffer action of charcoal. Since some culture media or cultures and many of the solutions whose buffer action must be studied for physiological purposes, contain undissolved or colloidal material which may act in this way, it seems best to consider buffer action in its broadest sense, and to express it by the relative slopes of titration curves determined experimentally. Further illustrations of titration curves of culture media will be found in the papers of Clark (1915) and of Bovie (1915). Titration curves of some inorganic solutions will be found in a paper by Hildebrand (1913).

The reader will have perceived the elementary theory underlying buffer action. The titration curve of phosphoric acid (fig. 5) illustrates the principles discussed on previous pages. The titration curve of a "peptone" solution integrates as it were the effects of acids, bases and ampholytes, in complex mixture.

Returning to figure 1 we see that along the flat portion of the curve considerable alkali has to be added to produce much change in pH. Conversely, the addition of a strong acid would not have anywhere near the effect at this flat portion of the curve that it would have near either end. Thus it is evident that a mixture of a single acid and its salt will tend to stablize the pH of the solution only within a certain narrow zone having vague boundaries. Mixtures buffering the solution within such a pH zone are often referred to as "regulator mixtures." They are of very great value to the analyst and the physiological chemist in that they furnish a means of stablizing the hydrogen ion concentration within a predetermined zone. The middle point of this zone, where the strongest buffer action is exerted, is determined approximately as shown on page 25 by the dissociation constant of the acid or base concerned. Other things being equal the choice of mixtures is thus revealed in a table of dissociation constants.

More theoretical treatments of the subject are given in the papers of Henderson (1909), Sørensen (1909), Sørensen (1912), Michaelis (1914) and Koppel and Spiro (1914).

Unless a solution is buffered to some extent in some way, it is almost impossible to make an accurate electrometric determination of the pH; and because of the influence of traces of carbon dioxid and other acidic or basic contaminations such solutions may be very unsuitable when used for physiological purposes. Thus the failure to buffer against the effect of so-called neutral salts which are not truly neutral may lead to gross error. In like manner the failure to buffer has rendered physiologically unstable certain so-called synthetic and supposedly stable culture media.

In the preparation of standard buffer mixtures it is of course, preferable to use a high grade of water if accuracy is required but there is little need of carrying this to an extreme. "Conductivity water" is sometimes specified for the preparation of special standards because the ordinary distilled water of certain regions of the country is such that "distilled water" means nothing. The exercise of judgment is advantageous.

The maintenance of "neutrality" by such solid reagents as calcium carbonate may be considered as a buffer action. It is very important to note however that the use of calcium carbonate may become a grossly inefficient procedure. To show its inefficiency the author has placed at the bottom of a test tube a deep layer of very finely divided, freshly precipitated and well washed calcium carbonate and overlaid this with cultures of bacteria and molds in sugar media. Indicators show that unless the calcium carbonate is frequently and thoroughly shaken with the medium only the solution in direct contact with the calcium carbonate is neutralized. Molds may develop an acidity as high as pH 2 within a few millimeters of the carbonate.

THE CONDUCT OF STRONG ELECTROLYTES

The relations set forth in the preceding pages, even in the approximate form adopted to keep the distinctive lines of the picture clear, afford in their experimental verification the best of evidence that the theory of electrolytic dissociation is essentially correct. That it is incomplete is shown when we turn to the examination of the quantitative data for strong electrolytes—acids such as hydrochloric and nitric and salts such as the simple chlorides. For instance, if the conductance of a solution is ascribed to the concentration and the mobilities of the ions, and if the mobilities be considered constant at all dilutions, the conductance data should satisfy the Ostwald dilution law and furnish a dissociation *constant*. The Ostwald dilution law is $\frac{a^2}{(1-a)v} = k$

where a is the degree of dissociation, v the dilution and k the equilibrium constant which should be independent of the dilution; a should be equal to the ratio of equivalent conductance at dilution v to equivalent conductance calculated for infinite dilution. For potassium chloride, k varies from 0.049 at 1000 dilution to 0.541 at 10 dilution. The discrepancies with hydrochloric acid are comparable.

The reader will recall that in the derivation of the equilibrium constant (page 19) there was introduced an assumption full of danger. The assumption was that the physical environment, within which occur the reactions of dissociation and recombination, remain constant. It has already been mentioned that a change in temperature changes the equilibrium constant and that a change in solvent produces a more profound effect. Now it is not at all improbable that the presence of relatively large concentrations of ions and especially of the hydrogen or hydroxyl ions constitutes an environment appreciably different from that of a dilute solution. If so, we should hardly expect to find an equilibrium *constant* holding over a great range of concentration. Yet it is by changing concentration that we expect to so alter the distribution of "species" that we may demonstrate the "mass" law experimentally.

But there are other possible difficulties. For instance, data upon what may be called the structure of solutions, the mutual influence of solvent and solute upon association of solvent molecules, association of solute molecules and association of solvent with solute are still hazy. Furthermore it is difficult to say what degree of separation constitutes ionization as measured by different methods. Therefore it is impossible to give rigidly accurate values to the concentrations of active molecules. When, therefore, it is stated that the anomalies of strong electrolytes "disprove the mass law," it may be only a clumsy way of saying that we do not know how to give the case an adequate test.

To give any adequate review of the present status of the problem would require undue space A most valuable review appeared in the discussions which took place in the Faraday Society and which are published in the December, 1919, number of the *Transactions*. It is there made very evident that the "anomalies" of strong electrolytes have been the bugbear of students of ioni-

zation, have stimulated most brilliant researches and promise to be the starting point for new developments which will harmonize the entire body of data.

There have been attempts to formulate the facts by means of purely empirical equations; and then again the pendulum has swung back to a faith that the original simple assumptions could be satisfied if interfering factors were discovered and their numerical magnitudes introduced as *corrections*. More recently there has come to the fore the "activity" concept of Lewis. This will be mentioned again in Chapter XIX. This concept has attained considerable success in systematizing the data; but whether it will have an appeal universal enough to satisfy minds of the type of Lord Kelvin, which reason not only in abstract terms but also demand concrete models, remains to be seen.

When there occur in the development of a science such baffling difficulties as have arisen in the case of "strong electrolytes," it is highly desirable to abandon both complex reasoning and endless corrections, if an entirely new basis can be found. This statement will appear gratuitous or even foolish to those who are so possessed with the idea of the complexity of aqueous solutions that they admit no theory as sufficient that is not itself complex; but the history of other developments has shown that in the face of similar complexities a simple basis of reference has been found and has won acceptance through its convenience.

Whatever may be the opinion of the reader he will doubtless agree that we are in the midst of or at the beginning of a period of transition, and that it is incumbent upon the experimenter to keep his data as free as is convenient from confusions introduced by tacit assumptions. In the following treatment of our subject assumptions common to the age will remain, but at least they will be more clearly recognized than if we straddled the issue that has arisen. We shall therefore proceed with the concept of "concentration" as commonly used, since it is the more convenient for elementary descriptive text. Finally in Chapter XIX we shall redefine certain standards in such a way as to embody current procedures and at the same time relieve the biochemist from embarrassments due to the present state of flux.

Although free acidities of a magnitude that fall within the grosser uncertainties of our knowledge of strong electrolytes are

seldom met in physiological solutions, the whole system of pH measurements is scaled from certain assumptions regarding the now uncertain conduct of HCl as will be shown in Chapter XIX. Furthermore we have continually to deal with solutions containing salts whose conduct is so little understood that precise treatment is impossible. This will appear in the so-called salt error of indicators and the strange fact that the *apparent* hydrogen ion concentration as determined with the hydrogen electrode may be raised above the quantity of available acid present by the addition of sufficient salt. To deal with such questions without tracing back through the subtleties of certain tacit assumptions is a most pernicious practice. It seems wiser to admit at once that certain of the more fundamental assumptions are too insecurely based to provide any adequate systematic treatment at the present time, and for this reason such questions as the salt error of indicators will be given in the subsequent chapters what may at first appear to be too brief a treatment. Experimentally the safest procedure to follow whenever the conduct of strong electrolytes enters into the determination of or the use of pH values is standardization of data.

Standardization of experimental data on the one hand and the maintenance of the more simple concepts of the theory of electrolytic dissociation will then be the policy of the following treatment.

SUPPLEMENTARY REFERENCES

A few references on the conduct of "strong electrolytes" and the "activity" concept. Arrhenius (1887, 1914), Beattie (1920), Bjerrum (1919), Bronsted (1919–1922), Ebert (1921,)* Ferguson (1916), Ferguson-France (1921), Getman (1920), Ghosh (1921), Harkins (1920), Harned (1916, 1920, 1922), Hill (1921), Jahn (1900), Kendall (1921, 1922), Kraus (1920, 1921), Lapworth (1915), Lewis (1907–1922), Linhart (1917, 1919), Noyes (1907), Noyes-MacInnes (1920), MacInnes (1919), Rabinowitsch (1921), Stern (1922). Symposium on theory of electrolytic dissociation Trans. Faraday Society *15*, 1–178, Dec. 1919.

pH calculator. Klopsteg (1921).

pH tables and graphs. Appendix table B. Matula (1916), Roaf (1920), Schmidt-Hoagland (1919), Symes (1916).

* Contains extensive review.

CHAPTER III

Outline of a Colorimetric Method

Acidimetric-alkalimetric indicators are substances, the colors of which correlate with the hydrogen ion concentrations of the aqueous solutions in which they are dissolved.

For each indicator there is a characteristic pH zone. On the acid side of this zone the indicator is completely transformed into its "acid color" and on the alkaline side of this zone the indicator is completely transformed into its "alkaline color." Within the characteristic pH zone there may be observed different proportions of the acid and alkaline colors.

In ordinary titrations conditions are so chosen that when the "end-point" of the titration is reached the pH of the solution passes suddenly through the whole range of the indicator's color-change. The intermediate stages, if observed, are not emphasized. The intermediate colors, however, are the important ones for the present purpose. They can be maintained with buffer solutions; and, being characteristic at definite pH values, they can be used to estimate the pH of tested solutions by a system of comparison with standards. To distinguish the stabilized degree of color transformation from the changing color observed during a titration, we shall adopt Sørensen's term and speak of the *virage* of an indicator when referring to a particular, stabilized degree of color transformation.

For reasons which will be given in Chapter IV the characteristic pH zone, within which differences of virage may be observed, is comparatively narrow. It is therefore necessary to have a series of indicators, the zones of which overlap (see table on page 80). Then if an indicator is found to be transformed completely to its acid color by a solution under test, the indicator next in the series is tried and so on until there is found the indicator which is transformed by the solution to an intermediate virage. It is then known that the solution has a pH value within the limits characteristic of the indicator used.

For some purposes it is sufficient to know the approximate pH and this may be estimated from the degree of color transformation

induced in the indicator. It is a simple matter, however, to take the first step toward accuracy. This is done as follows.

There have been determined by hydrogen-electrode methods the pH values of definite buffer solutions such as mixtures of KH_2PO_4 and Na_2HPO_4. Series of such solutions and the details of their preparation are described in Chapter VI. By adding definite quantities of an indicator to definite volumes of these standard solutions a series of color standards is easily prepared. With these standards the color of the tested solution can be compared. For instance, suppose that the preliminary test of a given solution has shown that it transforms the indicator phenol red neither to a full red nor to a bright yellow but that the proportion of red is low. Previous experience has impressed the fact that such a virage with phenol red indicates the solution to be near pH 7.0. See the color chart. Therefore, one employs those standard buffer solutions giving pH values near 7.0. To a series of uniform test tubes is added seriatim 10 cc. of each of the standard phosphate solutions described in Chapter VI. To each tube is added five drops of phenol red solution. On mixing there will be observed a graded series of virages and perhaps three of them will be recognized at once to have nearly the same color as 10 cc. of the tested solution mixed with 5 drops of the same indicator solution. When closer inspection shows where the color-match occurs, the standard with its known pH value and the tested solution are supposed to have the same pH value. As in this example, it is always necessary to make comparisons between *like concentrations* of indicator viewed through *equal depths* of solution.

An error may be made if the standard and tested solutions differ much in total salt concentration, or if the tested solution contains much protein, or if an unreliable indicator is used. But we shall have to deal with these and other difficulties in subsequent chapters.

When one is familiar with the virages of the indicators at known pH values very fair estimations may be made without the aid of the standards; but there is no way as satisfactory as the setting up of the standards for the establishment of a correct impression of the relations of the various indicators on the pH

scale. On the other hand, the author has discovered in his conversations that there are a great many investigators who would like to use indicators for the occasional rough measurement of pH but who are discouraged by a pressure of work which prevents them from taking the time to carefully prepare the standard solutions. To furnish such investigators with a demonstration of the general relations of the various indicators and to furnish *rough* standards the attempt has been made in figure 8, to reproduce the colors. The colors of standard buffer solutions containing definite quantities of the several indicators were reproduced very faithfully by Mr. Max Broedel of the Johns Hopkins Medical School. It must be remembered, however, that in undertaking a second reproduction by means of the printer's art the publishers are to be commended for their courage and are not to be held responsible for the inadequacy of the result. Aside from the inherent difficulty in freeing a printed color from the effect of the vehicle, there remains the utter impossibility of reproducing upon paper the *exact virage* observed in a liquid solution. The fundamental phenomena are *quantitatively* very different in the two cases. Therefore the user of the chart of colors will have to use discretion and some imagination. If he does not attempt to make the reproductions take the place of the standards he should find them useful for class room demonstrations, for refreshing the memory and for rough standards.[1]

In each case the colors were reproduced from tubes 16 mm. internal diameter containing 10 cc. standard buffer solution. The quantities of indicator solution added in each case were as follows: Thymol blue, acid range (T. B. acid range) 1 cc. 0.04 per cent solution. Brom phenol blue (B. P. B.) 0.5 cc. 0.04 per cent solution. Methyl red (M. R.) 0.3 cc. 0.02 per cent solution. Brom cresol purple (B. C. P.) 0.5 cc. 0.04 per cent solution. Brom thymol blue (B. T. B.) 0.5 cc. 0.04 per cent solution. Phenol red (P. R.) 0.5 cc. 0.02 per cent solution. Cresol red (C. R.) 0.5 cc. 0.02 per cent solution. Thymol blue (T. B.) 0.5 cc. 0.04 per cent solution.

[1] Separates of the color chart may be obtained from the publisher.

Dr. Barnett Cohen of the Hygienic Laboratory has recently (Public Health Reports, U. S. P. H. S., *38*, 199, 1923) synthesized the following new sulfonphthalein. Brom cresol green covers the range of methyl red. Salt and protein errors have not yet been determined.

CHEMICAL NAME	SUGGESTED COMMON NAME	APPARENT DISSOCIATION CONSTANT	pH RANGE
m-Cresol sulfonphthalein	Meta cresol purple	2.8×10^{-2} 5.0×10^{-9}	0.5–2.5 7.6–9.2
Dibromo-dichloro-phenol sulfonphthalein	Brom-chlor phenol blue	7.9×10^{-5}	3.2–4.8
Tetra bromo-m-cresol sulfonphthalein	Brom cresol green	1.0×10^{-5}	4.0–5.6
Dichloro-phenol sulfonphthalein	Chlor phenol red	8.9×10^{-7}	5.0–6.6
Dibromo-phenol sulfonphthalein	Brom phenol red	4.5×10^{-7}	5.4–7.0

The ranges of pH covered by the several indicators in the color chart are:

T. B. (acid range), Thymol blue	1.2–2.8
B. P. B., Brom phenol blue	3.0–4.6
M. R., Methyl red	4.4–6.0
B. C. P., Brom cresol purple	5.4–7.0
B. T. B., Brom thymol blue	6.0–7.6
P. R., Phenol red	6.6–8.2
C. R., Cresol red	7.2–8.8
T. B., Thymol blue	8.2–9.8

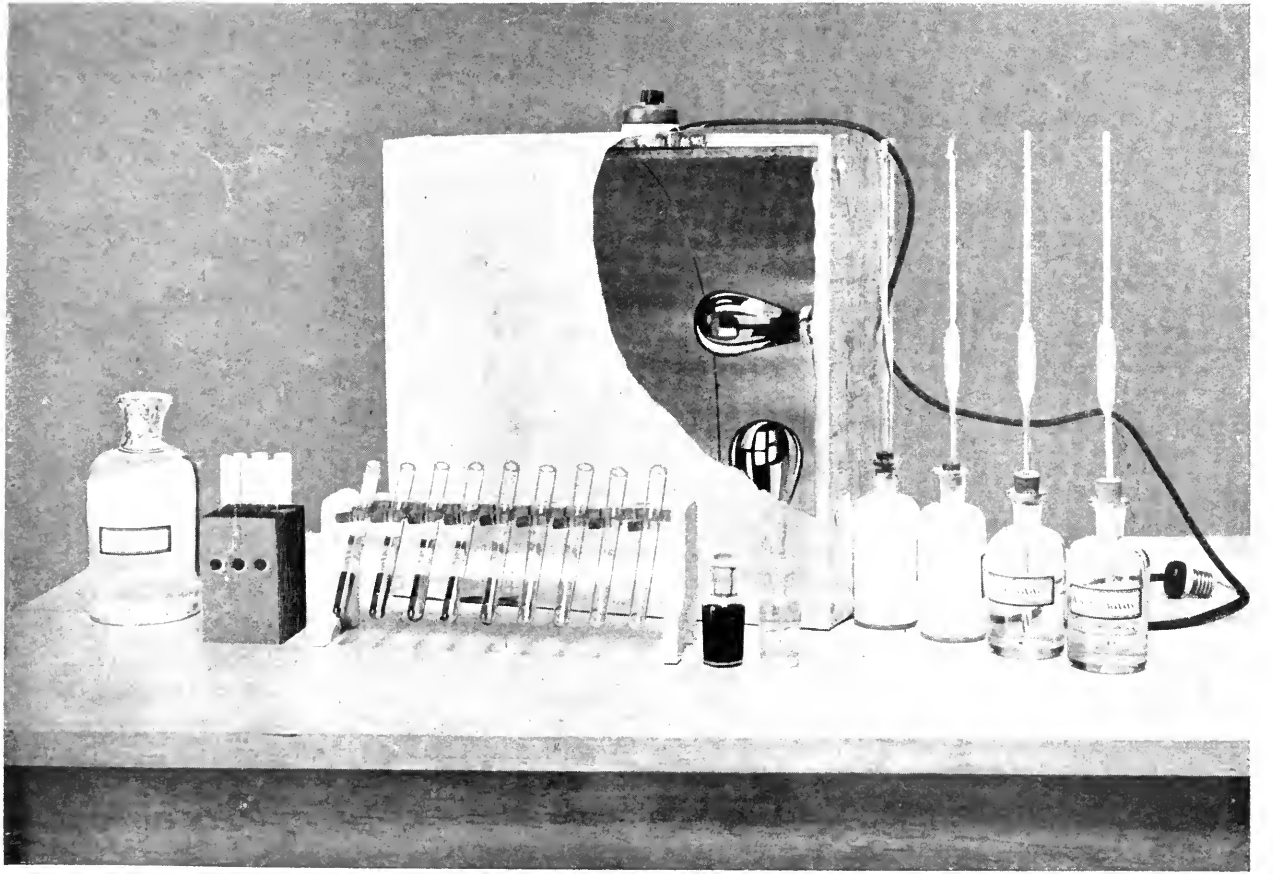
For class-room work it is advantageous to show the position of the several indicators on the pH scale by relining each series so that corresponding pH values overlap.

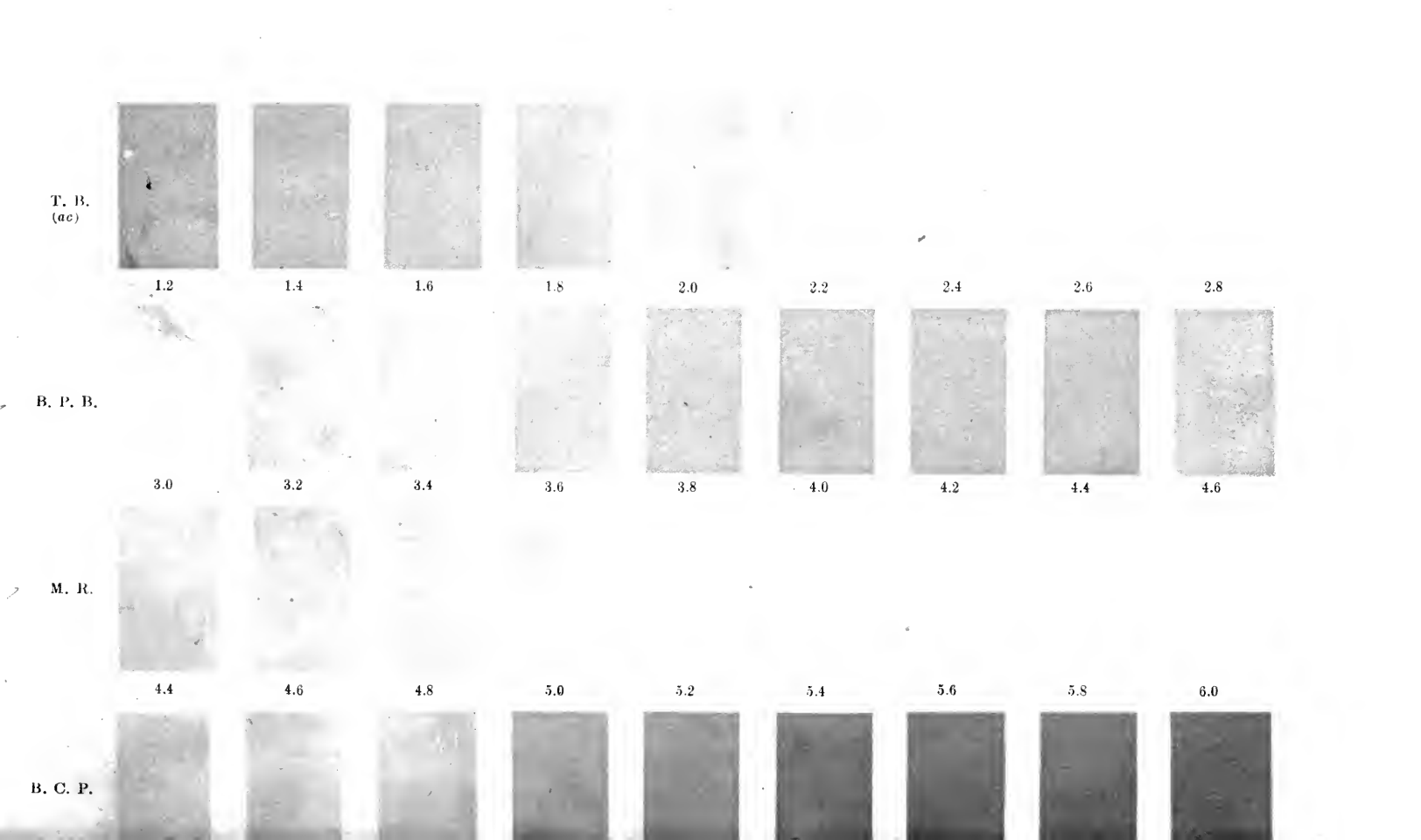
One requirement for the colorimetric method is a set of indicators selected for their relative freedom from the so-called protein and salt errors and for their brilliancy. Beside the brilliant and reliable selection of Clark and Lubs there is the carefully studied selection of Sørensen given on page 78 with Sørensen's summary of properties on page 79.

There are also required standard buffer solutions whose pH values are established from hydrogen electrode measurements. It is in the preparation of these standards that the greater part of the labor of the colorimetric method is involved; but, once the stock solutions are carefully made, the preparation of the mixtures is a simple matter. If only the pH range 5.2 to 8.0 is necessary, the Sørensen mixtures of primary and secondary phosphates are the more convenient. If a wider range is desired the system tabulated on pages 106 to 107 is recommended.

For precise measurements there are required control by hydrogen electrode measurements and constant watchfulness for the several sources of error noted in following chapters. Approximate methods are described in Chapter VIII.

In figure 7 are shown several pieces of equipment useful in colorimetric work. Beginning at the left is, first, a sample of a litre bottle used for holding the standard stock solutions, such as M/5 KH Phthalate, which are not seriously affected by exposure to the carbon dioxide of the laboratory air. In Clark and Lubs' series of standards (see page 99) there are required four such bottles. In this same series there is required a container for




T. B.
(ac)
1.2
1.4
1.6
1.8
2.0
2.2
2.4
2.6
2.8
B. P. B.
3.0
3.2
3.4
3.6
3.8
4.0
4.2
4.4
4.6
M. R.
4.4
4.6
4.8
5.0
5.2
5.4
5.6
5.8
6.0
B. C. P.

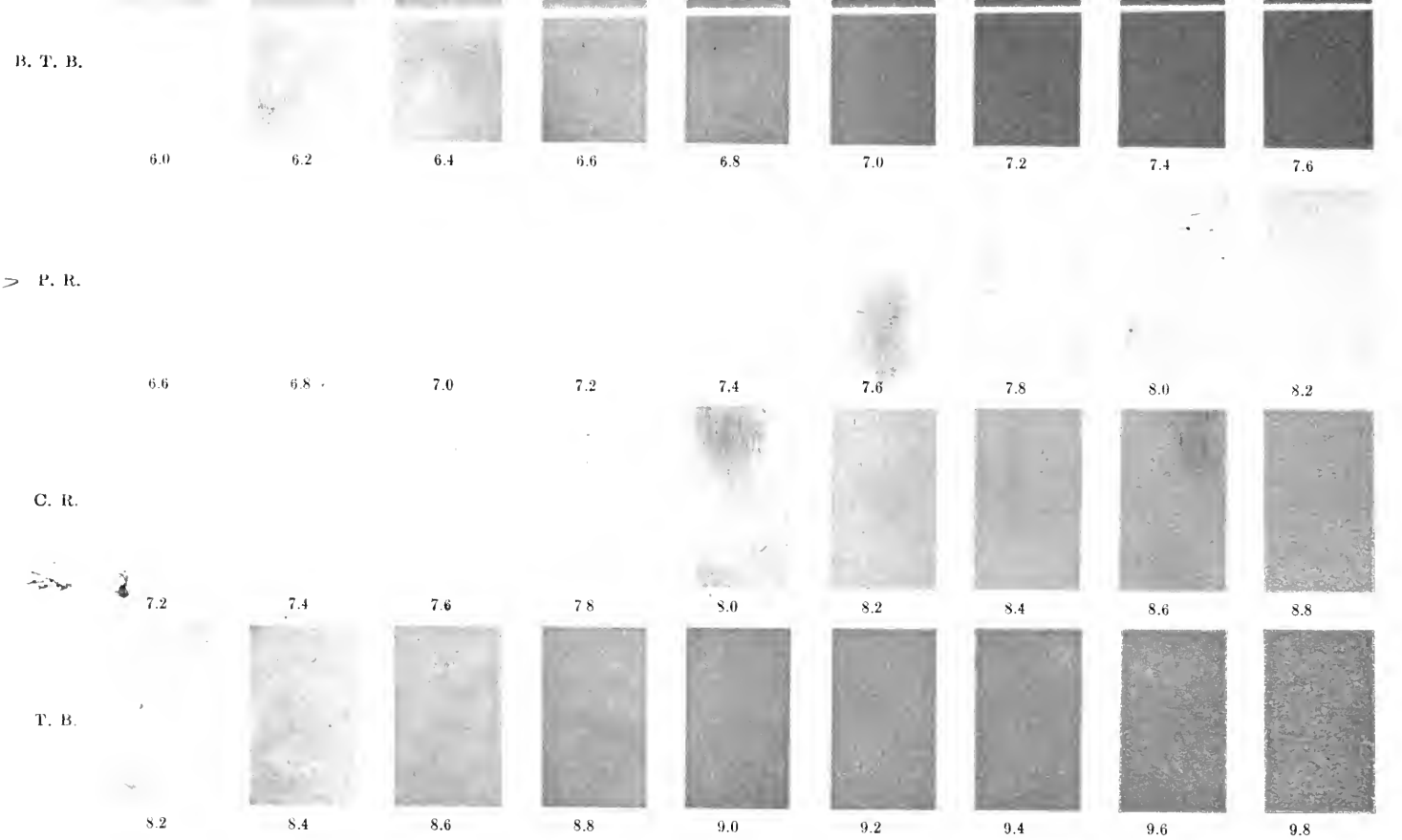


FIG. 8. CHART SHOWING THE COLORS OF CLARK AND LUBS' INDICATORS IN SOLUTIONS OF KNOWN pH

standard M/5 NaOH. This should be a paraffined bottle with calibrated burette and soda-lime guard-tubes attached.

In figure 7 there is next shown a comparator whose construction is given on page 70. This is used in comparing turbid or colored solutions with the standards. When the turbidity of a tested solution brings into evidence the dichromatism of an indicator as described on page 65, the comparator is used with the light screen shown at the back of figure 7 and described on page 67.

For ordinary colorimetric comparisons the test tube rack shown in the figure is very useful. The holders are the clips sold at stationers for holding rubber stamps. Two forms of dropping bottle are next shown and, finally, at the right, two paraffined bottles for alkaline standards and two acid resistant bottles for acid solution. Of such bottles there are required for the series of standards given on pages 106–107 fifty-one bottles and the same number of 10 cc. pipettes. The range of pH thus covered is wider than that called for in special investigations. The pipettes may have their tips broken to allow quicker delivery of solution without serious violation of volume requirements. Sørensen's standards, pages 111-114, are designed so that individual 10 cc. samples are made up as required. Larger quantities such as are specified in table 21 provide for the occasional test.

CHAPTER IV

Theory of Indicators

Les propriétés des corps sont les propriétés des nombres.—De Chancourtois.

Indicator theory is a cross-roads where the cultivators of distinct fields of science meet. Here comes the organic chemist with analyses of plant and animal products, structural formulas of synthetic dyes, tautomers and chromophores. Here comes the physico-chemist with formulations of electrolytic and tautomeric equilibria. Here comes the physicist with the theory of color and the instruments of light analysis. And perhaps there will meet here the psychologist bringing a clearer description of the subjective aspect. As a confluence of trade routes may determine the growth of a city so the confluence of many specialties may sometime lead to a great community of interest where the cross-roads of indicator theory once lay. Indicators themselves are not particularly unique except that they compel the attention of the eye. Through this we are made aware of phenomena of wide occurrence.

According to the inclination of a reviewer one or another of the manifold aspects of indicator theory might be emphasized. We must choose that which is useful to the purpose at hand and for the sake of a necessary brevity we must try to include only so much as will contribute toward an intelligent use of indicators as tools for the determination of hydrogen ion concentration.

In the first place it may be said that the customary manner of using indicators is merely a method of comparison involving little if any theory. The conduct of an indicator may be, and generally is, "calibrated" by means of hydrogen electrode measurements. It is well to emphasize this uninspiring, matter-of-fact aspect because it will remind us that with so much of the fundamental theory at hand the employment of theory may lead to a wider usefulness of the instruments thus far treated empirically. But before this can be done important relationships must be expressed definitely in *numerical data*. How this can be done is the immediate problem before us.

The first consistent attempt to bring the conduct of indicators into relation with electrolytic dissociation was that of Ostwald (1891). He assumed that indicators are acids or bases the undissociated molecules of which have a color different from that of their dissociation products. If this be so, it is evident that the color

TABLE 4

Approximate apparent dissociation constants of indicators

INDICATOR	K_a	pK_a
Phenol sulfon phthalein	1.2×10^{-8}	7.9
o-Cresol sulfon phthalein	5.0×10^{-9}	8.3
Thymol sulfon phthalein	1.2×10^{-9}	8.9
Carvacrol sulfon phthalein	1.0×10^{-9}	9.0
α-Naphthol sulfon phthalein	5.3×10^{-9}	8.2
Tetra bromo phenol sulfon phthalein	7.9×10^{-5}	4.1
Di bromo o-cresol sulfon phthalein	5.0×10^{-7}	6.3
Di bromo thymol sulfon phthalein	1.0×10^{-7}	7.0
Phenol phthalein	2.0×10^{-10}	9.7*
o-Cresol phthalein	4.0×10^{-10}	9.4
α-Naphthol phthalein	4.0×10^{-9}	8.4
Methyl red	7.9×10^{-6}	5.1†
Ethyl red	4.0×10^{-6}	5.4
Propyl red	4.0×10^{-6}	5.4†
Thymol sulfon phthalein (acid range)	2.0×10^{-2}	1.7

* This value is identical with Rosenstein's (1912).

† In the table published in the Journal of the Washington Academy, vol. vi, p. 485, these values for methyl red and propyl red were erroneously interchanged.

Tizard (1910) gives $K_a = 1.05 \times 10^{-5}$ or $pK = 4.98$ for methyl red considered as an acid.

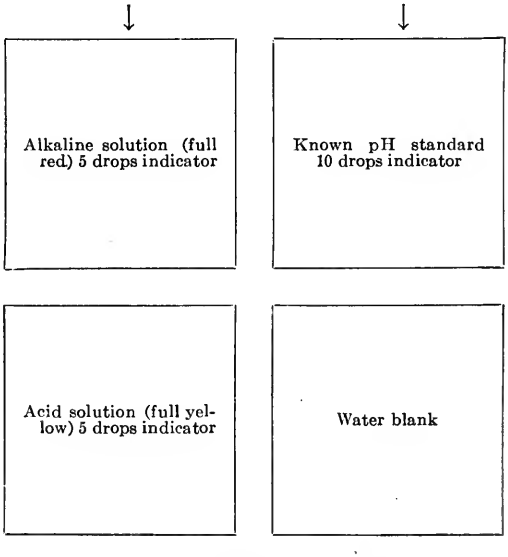
of an indicator should change with the pH of a solution exactly as the dissociation curves described in Chapter I. If, for instance, the indicator is an acid, colorless in the undissociated form, but colored when dissociated as an anion, then the change of color with the hydrogen ion concentration should conform to the equation:

$$\alpha = \frac{K_a}{K_a + [H^+]}$$

where K_a is the dissociation constant of the acid indicator and α is the degree of dissociation. Assuming then that such a rela-

tion does hold, let us determine K_a for a series of indicators in the following way.

From the above equation when $\alpha = \frac{1}{2}$, $K_a = [H^+]$. That is, at a hydrogen ion concentration corresponding numerically to the dissociation constant, the acid is half dissociated. At such a hydrogen ion concentration a colorless-to-red indicator, such as phenolphthalein, should show half the available color; and a yellow-to-red indicator, such as phenol red, should show the half-yellow, half-red state. We can match the half way state of this first solution by superimposing two solutions each of a depth equal to the first, if we have in one of the superimposed solutions only the yellow form and in the other only the red form, each concentration equaling half the concentration in the first solution. Such an arrangement is shown diagraphically in the following figure:



We may not know at the beginning at what pH the half transformation may occur, so we vary the pH of the standard solution until a match with our superimposed solutions does occur. Then we have found, presumably, the hydrogen ion concentration the numerical value of which is the dissociation constant of the indicator. Values so obtained by Clark and Lubs (1917) are given in table 4.

As indicated in Chapter I the determination of the dissociation curve, or of the half transformation point, does not tell us whether we are dealing with the dissociation curve of an acid or the dissociation-residue curve of a base or vice versa. Thus methyl red is treated in table 4 as an acid and plotted in figure 9 as if the color were associated with the undissociated form. Methyl red however could be treated as a base.

Just as it is convenient to deal with a logarithmic function of $[H^+]$ so the dissociation constants can be used in the form $\log \frac{1}{K_a}$. This can be designated pK_a.

Gillespie (1920) gives somewhat different values but, since the method used in each case was approximate, the table given above, as it is found in the paper by Clark and Lubs (1917) will do for purposes of illustration. With the aid of the approximately determined apparent dissociation constants we are enabled to plot the curves shown in figure 9, which reveal graphically the relationships of the various indicators in the series we shall discuss. This figure shows at a glance that an indicator of the simple type we have assumed has no appreciable dissociation and consequently exists in only one colored form at pH values beginning about 2 points below the half transformation point, while at the same distance above this point the indicator is completely dissociated and exists only in its second form. Between these limits the color changes may be observed. The useful range of such an indicator is far less than 4 pH units for optical reasons which will be discussed later.

The illustration (fig. 9) will show how in choosing a set of indicators it is advantageous to include a sufficient number, if reliable indicators can be found, so that their ranges overlap. It shows that each of the indicators, when considered to be of the simple type we have assumed, has an equal range. It also shows that the half transformation point of each indicator occurs nearer one end of the useful range, the useful range being indicated by the shaded part of the curve. This aspect will be discussed later.

It is evident that if the actual color change of an indicator varied with pH in accordance with a curve such as those in figure 9, and if the true dissociation constant were accurately known, then the hydrogen ion concentration of a solution could be determined

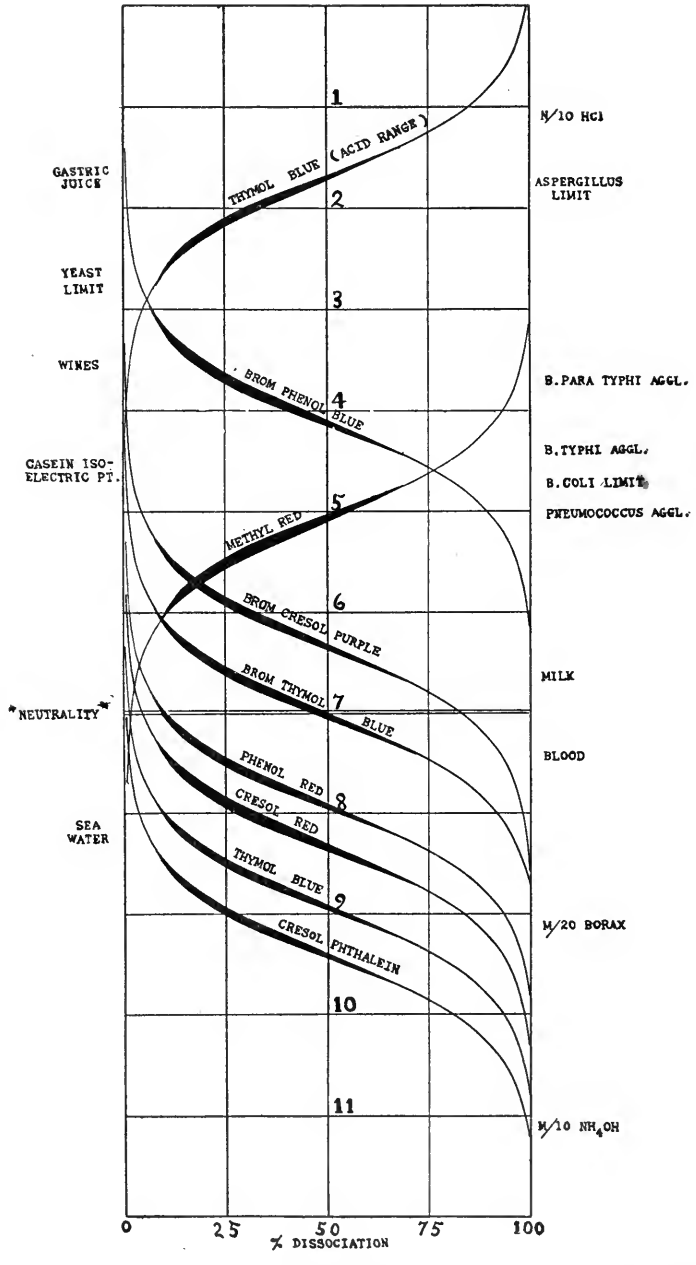


FIG. 9. INDICATOR CURVES AND SIGNIFICANT pH VALUES. SHADING INDICATES USEFUL RANGE

by finding the percentage transformation induced in the indicator. Indeed the dissociation constants of some few indicators have been determined with sufficient accuracy to permit the use of this method when the proper means of determining the color intensities are used. This will be discussed in Chapter VIII.

We have been assuming that the theory of indicators may be treated in the simple manner originally outlined by Ostwald (1891). In his theory it was assumed that the anion of an indicator acid, for instance, has a color different from that of the undissociated molecule. This assumption if unmodified does not harmonize with what is known. Researches in the phenomena of tautomerism have shown that when a change in color is observed in an indicator solution the change is associated with the formation of a new substance which is generally a molecular rearrangement or so-called "tautomer" of the old. If this color change is associated with the transformation of one substance into another, how is it that it seems to be controlled by the hydrogen ion concentration of the solution? As Steiglitz (1903) and others have pointed out, it is the state of these compounds, their existence in a dissociated or undissociated condition, which determines the stability of any one form.

The method of dealing with the tautomeric relations of indicators is shown by the following quotation from Noyes (1910):

We may derive a general expression (as has previously been done by Acree, 1907) for the equilibrium-relations of any pair of tautomeric acids and their ions. The three fundamental equilibrium equations are as follows:

$$\frac{(H^+)\ (In'^-)}{(HIn')} = K_I'; \quad (20) \qquad \frac{(H^+)\ (In''^-)}{(HIn'')} = K''_I; \quad (21)$$

$$\frac{(HIn'')}{(HIn')} = K_T; \quad (22)$$

Multiplying (21) by (22), adding (20) to the product, and substituting in the denominator for (HIn') its value $\frac{(HIn') + (HIn'')}{1 + K_T}$ given by (22), we get

$$\frac{(H^+)\ [(In'^-) + (In''^-)]}{(HIn') + (HIn'')} = \frac{K'_I + K''_I\ K_T}{1 + K_T} = K_{IA} \qquad (23)$$

If the indicator is a base existing as the two tautomeric substances In'OH and In''OH, having ionization constants K'_I and K''_I and a tautomer constant K_T defined by equations analogous to (20), (21) and (22), the

general expression for the equilibrium between the ionized bases and their ions is:

$$\frac{(OH^-)\,[(In'^+) + (In''^+)]}{(In'OH) + (In''OH)} = \frac{K'_I + K''_I\,K_T}{1 + K_T} = K_{IB} \qquad (24)$$

In these expressions a single constant K_{IA} or K_{IB} has been introduced in place of the function of the three constants K'_I, K''_I, and K_T . . . . The constant so calculated for a pair of tautomeric acids or bases can evidently be substituted for the ionization constant of an ordinary (non tautomeric) acid in any derived expression, provided the *sum* of the two ion concentrations and the sum of the two acid or base concentrations are quantities that are to be known or are to be calculated.

If then in equation (23) we substitute (In^-) for $[(In'^-) + (In''^-)]$ and (HIn) for [(HIn′) + (HIn″)] we have:

$$\frac{(H^+)\,(In^-)}{(HIn)} = K_{IA} \qquad (25)$$

Applying to Noyes' equation (25) the derivation given on page 25 we have

$$\alpha = \frac{K_{IA}}{K_{IA} + (H^+)}.$$

From this we may plot the curves of figure 9. Such curves will then represent the color transformations when and only when (In^-) is substantially equal to (In'^-) or to (In''^-), whichever tautomer is associated with the color. The most probable explanation of the fact that such curves do represent very closely the color transformations in certain instances is that K_T (see equation (23)) is so small that the dissociation brought about by salt formation leaves (In^-) dominant.

In other words it is, after all, the degree of dissociation, as determined by the hydrogen ion concentration, that determines which tautomer predominates. Therefore, consideration of the tautomeric equilibria only modifies the original Ostwald treatment to this extent: the true dissociation constant is a function of the several equilibrium and ionization constants involving the different tautomers and must be replaced by what Acree calls the "total affinity constant," or by what Noyes calls the "apparent dissociation constant," when it is desired to show directly how the color depends upon the hydrogen ion concentration.

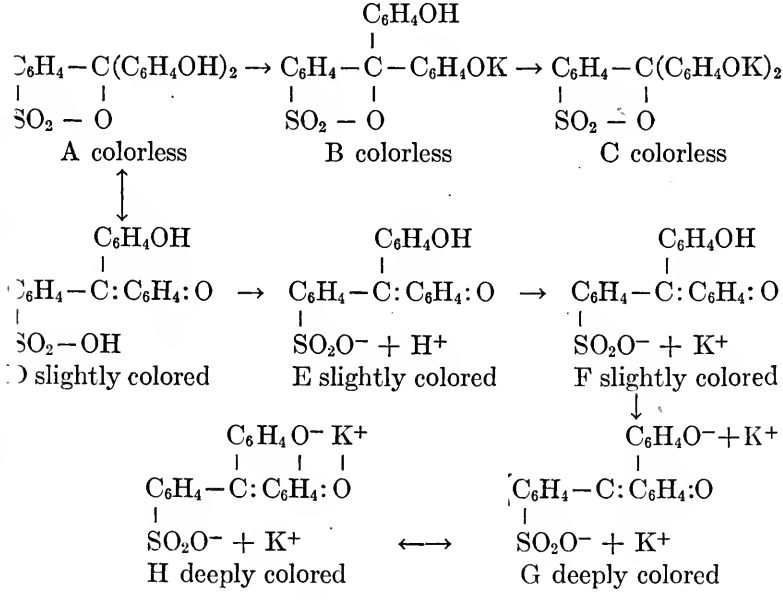
Many indicators are poly-acidic or poly-basic and will not rigidly conform to the treatment for a simple monovalent acid such as we have described. Phenolphthalein, for instance, as was shown by Acree (1908) and by Wegscheider (1908) must be

considered as poly-acidic. The proper equations to apply in this case have been given by Acree (1907, 1908) and also by Wegscheider (1908, 1915). According to Acree and his students (Acree, 1908) (Acree and Slagle, 1909) the chief color change in phenolphthalein is associated with the presence of a quinone group and with the ionization of one of the phenol groups. In the sulfon phthalein series of indicators Acree and his students (White, 1915, and White and Acree, 1918) have found much the same sort of condition.

In the sulfon phthalein series, however, certain unique properties described by Lubs and Acree (1916) make the series eminently suited for experimental demonstration of the seat of color change.

In the sulfon phthalein group of indicators we have to deal with poly-acids; but as Acree has shown, the dissociation constant of the strong sulfonic acid group is so very much greater than that of the weak phenolic group, with which the principal color change is associated, that there is no serious interference. As shown in Chapter I we may, therefore, plot the curves for the chief color-changes as if we were dealing with monobasic acids.

The structures of all the sulfon phthaleins are analogous to that of phenol sulfon phthalein (phenol red) whose various tautomers are given by Lubs and Acree (1916) in the following scheme:



The colorless lactoid A by reason of the strong tendency of the sulfonic acid group to ionize goes over into the quinoid structures illustrated in the second line which are slightly colored yellow. It is the transformation of F to G and H, the ionization of the phenolic group forming a quinone-phenolate structure which correlates with the intense red color of phenol sulfon phthalein (phenol red).

Just as the discovery of tautomerism seemed at first to discredit the original form of the Ostwald theory of color change, so it is now realized that a mere change in structure is of itself quite inadequate to account for the change in the light absorption upon which the color of a solution depends. Light is an electromagnetic phenomenon and the absorption of the energy in any particular train of light is undoubtedly due to the resonance of electrons. Thus the direct connection between light absorption and molecular structure will be found in the relation of molecular structure to the distribution and freedom of the component electrons. It is in this direction that Baly (1915) believes a satisfying theory of the colors of dyes will be found. Although Baly has called attention to difficulties in the correlation of colors with tautomeric changes there seems to be no inherent reason why tautomerism, alteration of the fields of force within the molecule, electrolytic ionization and color should not be correlated. The original Ostwald theory may yet prove to be essentially correct in that the charging of a molecule by ionization should cause a redistribution of the fields of force. Whether or not a molecular rearrangement or absorption of a particular train of visible light follows may well depend upon particular circumstances. But of course all this is left to the future and to quantitative data.

OPTICAL ASPECTS

While the color changes of indicators are correlated with molecular rearrangements controlled by hydrogen ion concentrations, it should not be forgotten that the phenomena observed are optical and that no theory of indicators can be considered complete enough for practical purposes which fails to recognize this. As ordinarily observed in laboratory vessels the color observed

is due to a somewhat complex set of phenomena. It is unfortunate that we have no adequate treatment of the subject which at the same time embraces electrolytic dissociation, tautomerism and the optical phenomena in a manner directly available in the practical application of indicators. The simultaneous treatment of these various aspects is necessary before we can feel quite sure of our ground when dealing with discrepancies often observed in the comparison of colorimetric and electrometric measurements of biological fluids.

Let us first consider the range of an indicator as it is determined by the differentiating power of the eye. An approximate treatment of this is all that will be attempted.

Using equation (10), cf. page 26:

$$\mathrm{pH} = \log \frac{1}{K} + \log \frac{\alpha}{(1 - \alpha)}$$

we find on differentiation that the rate of increase in α with increase of pH is:

$$\frac{d\alpha}{d(\mathrm{pH})} = \alpha (1 - \alpha).$$

When

$$\frac{d^2\alpha}{d(\mathrm{pH})^2} = 0, \ \alpha = \frac{1}{2}.$$

In other words the maximum rate of increase in dissociation is at the half transformation point. This fixes a reference point when indicators are to be employed in distinguishing differences in pH. The question now arises whether or not this is the central point of the optimal conditions for differentiation of pH values. It may be said at once that it is not, because the eye has not only to detect differences but also to resolve these differences from the color already present. Experience shows that the point of maximum rate of increase in α is near one limit of the useful range and that this range lies on the side of lower color. Thus, in the case of the one-color indicator phenolphthalein, the useful zone lies between about 8.4 and 9.8 instead of being centered at 9.7 which corresponds with the point of half-transformation. In the case of a two-color indicator such as phenol red the

same reasoning holds, because the eye instinctively fixes upon the very dominant red. With other two-color indicators the principle holds except when there is no very great difference in the command upon the attention by one or the other color.

It should be mentioned however that these more or less empirical relations are observed in comparing virages at equal increments of pH when the indicator concentration is adjusted to emphasize the differences among the less intensely colored tubes. By suitable dilution of the indicator the differences among the tubes having the higher percentage color may be emphasized and the useful range of the indicator slightly extended. In practice this is a procedure which requires care for it is easy to become confused when dealing with different concentrations of the same indicator.

The fixing of the lower limit of usefulness of a given indicator involves another factor. There is the question of the total indicator which may be brought into action. A dilute solution of phenolphthalein may appear quite colorless at pH 8.4 while a much stronger solution will show a distinct color which would permit distinguishing 8.2 from 8.4. But the concentration is limited by the solubility of the indicator and therefore must be taken into consideration. In short there is no basis upon which to fix definite limits to the pH range of a given indicator, and those limits which are given must be considered to be arbitrary. On the other hand the apparent dissociation curve is quite definitive; and were it not for the greater convenience of the "range of usefulness" it would be preferable to define the characteristics of an indicator in terms of its apparent dissociation constant.

We ordinarily speak of color as it if were an entity. As a matter of fact the color exhibited by an indicator in solution is due to the selective absorption of certain frequencies of the incident light. This results in the partial or complete blocking off of the light in one or more regions of the spectrum, as may be seen by the dark band or bands which appear when the solution is viewed through a spectroscope. The transmitted light instead of being of the continuous spectrum which blends to subjective white is made up of the unaffected wave lengths and of those wave trains the intensities of which have been reduced to a greater or lesser

extent. *The resultant subjective color* must be distinguished from the color associated with a definite region of the spectrum.

We come now to the consideration of a phenomenon which is undoubtedly exhibited with all indicators but which is generally not noticed except in special instances. In some of these instances it becomes of great importance and may lead to serious error unless recognized. The phenomenon we speak of is the dichromatism exhibited, for instance, by solutions of brom phenol blue. Solutions of this indicator appear blue when viewed in thin layers but red in deep layers. The explanation is as follows: The dominant absorption band of the alkaline solution is in the yellow and the green, so that the transmitted light is composed almost entirely of the red and blue. The incident light has an intensity which we may call I. After transmission through unit thickness of solution some of the light has been absorbed and the intensity becomes Ia, where a is a fraction—the transmission coefficient—which depends upon the nature of the absorbing medium and the wave length of the light. After traversing thickness ϵ the intensity becomes Ia^{ϵ}. Now the transmitted blue is $I_b a_b^{\epsilon}$ and the transmitted red $I_r a_r^{\epsilon}$. We do not happen to know what the actual values are, but, merely to illustrate the *principle*, let us assume first that the intensity of the incident blue is 100 and of the red 30 and that $a_b = 0.5$ and $a_r = 0.8$.

For $\epsilon = 1$, $I_b a_b^{\epsilon} = 50$ and $I_r a_r^{\epsilon} = 24$. Hence blue greater than red.

For $\epsilon = 10$, $I_b a_b^{\epsilon} = 0.01$ and $I_r a_r^{\epsilon} = 0.30$. Hence blue less than red.

This example indicates that the solution may appear blue when viewed through thin layers while it may appear red when viewed through thick layers.

If we change the relative intensities of the incident red and blue we can change the color of a given thickness of solution. If in the above example we reversed the intensities of the incident red and blue, then,

For $\epsilon = 1$, $I_b a_b^{\epsilon} = 15$ and $I_r a_r^{\epsilon} = 80$, or red greater than blue.

This is essentially what happens when we carry the solution from daylight, rich in blue, to the light of an electric carbon fila-

ment lamp, poor in blue. The solution which appears blue in daylight appears red in the electric light.

The practical importance of recognizing the nature of this phenomenon may be illustrated in the following way. Suppose we have a solution rich in suspended material such as bacterial cells, and that we wish to determine its pH value by using brom phenol blue. If we view such a solution in deep layers very little of the light incident at the bottom reaches the eye. A large proportion of the light which does reach the eye is that which has entered from the side, has been reflected by the suspended particles, and has traversed only a relatively thin section of the solution. In such a solution then, if it is of the proper pH, brom phenol blue will appear blue, while in a clear comparison solution of the same pH the indicator appears red or purple if the tube is viewed lengthwise. A comparison is therefore impossible under these conditions. If, however, we view the two solutions in relatively thin layers, as from the side of a test tube, they will appear more nearly comparable. There will still remain, however, a clearly recognizable difference in the quality of the color which serves as a warning that the two solutions are not being compared under proper conditions.

Now a change in the quality of the light in which the turbid and the clear solutions are compared will, of course, not avert one fundamental difficulty—a difference in effective path; but a proper change in the quality of the light can eliminate the dichromatism and free the eye from one source of confusion. In the case at hand we might eliminate either the red or the blue. Which had best be eliminated is a question which can not be answered properly until we have before us the necessary spectrometric measurements. Nevertheless the following observations made with a small hand spectroscope, and the deductions therefrom may prove to be illuminating.

The chief absorption bands of brom phenol blue solutions occur in the yellow-green range and in the blue. In alkaline solutions the band in the blue disappears while that in the yellow widens into the green. As the solution is made more acid the band in the blue appears, shutting off the transmitted blue, while that in the yellow-green contracts, permitting the passage of the green. Our light source then should be such that at least one of these

changes may become apparent, and at the same time either the blue or red must be eliminated. The light of the mercury arc fulfills these conditions. It is relatively poor in red and it emits yellow, green and blue lines where the shifts in the absorption bands of brom phenol blue occur. Since the mercury arc is not generally available we have devised a light source to fulfill the alternative condition, namely, one which will permit observation of the contrasts due to the shift in the yellow-green band[1] and which at the same time is free from blue. Such a source is found in electric light from which the blue is screened by a translucent paper painted with a yellow, acid solution of phenol red. One disadvantage of such a screen is that the red transmitted through it is so dominant that it obscures the contrasts which are due to the shifting of the yellow-green absorption band. Nevertheless, such a screen has proved useful in pH determinations with brom phenol blue and particularly useful with brom cresol purple. In either case it is most useful in the more acid ranges covered by these indicators.

The device consists of an ordinary box of convenient size in which are mounted three or four large electric lights (e.g., 30 cp. carbon filaments). A piece of "tin" serves as reflector. The box may be lined with asbestos board. A piece of glass, cut to fit the box, is held in place on one side by the asbestos lining and on the other by a few tacks. This glass serves only to protect the screen and is not essential. The screen is made from translucent paper known to draughtsmen as "Economy" tracing paper. It is stretched across the open side of the box and painted with a solution consisting of 5 cc. of 0.6 per cent phenol red and 5 cc. of M/5 KH_2PO_4 (stock, standard phosphate solution). While the paper is wet it is stretched and pinned to the box with thumb tacks. This arrangement may be constructed in a very short time and will be found very helpful in many cases. It should be used in a dark room or, if such a room is not available, exterior light may be shut off with a photographer's black cloth.

While considering light sources we may call attention to the fact that all the sulfon phthalein indicators may be used in elec-

[1] This should not be confused with the changes in "subjective color." In the screened light no participation of transmitted green will be detected by the unaided eye.

tric light, although brom thymol blue and thymol blue are not well adapted for use in light poor in blue. Doubtless a more thorough investigation of the absorption spectra of the sulfon phthalein indicators will make it possible to devise light sources which will materially increase their efficiency.

So far as we have been able to detect with instruments at hand, the absorption spectra of all the indicators of the sulfon phthalein series are such that the appearance of dichromatism must be expected under certain conditions. It will be observed with phenol red in light relatively poor in red and rich in blue, for example, the light of a mercury arc; and with thymol blue in light relatively poor in blue and rich in red for example, ordinary electric light.

When the colorimeter is employed in the study of colored solutions the applicability of Beer's law is assumed. This may be expressed in the form, $\frac{L_1}{L_2} = \frac{C_2}{C_1}$ where C_1 and C_2 represent the concentrations of color in two solutions and L_1 and L_2 represent the depths of solution traveled by the light when a color match occurs. Applying this relation one is able to obtain the ratio of concentrations and therefrom the concentration in one solution if the concentration in the other be known. But as was shown above we have, in the case of two-color indicators, different transmission coefficients for various regions of the spectrum. Consequently the depth of a solution cannot be altered as it is in the ordinary colorimeter without seriously affecting the quality of the emergent light.

When such shifts in quality occur it is impossible without the aid of elaborate photometric devices to make an accurate comparison of intensities. This at once limits the usefulness of the ordinary colorimeter, a cardinal principle of which is an accurate device for varying and measuring the depth of view. That feature of certain instruments whereby two optical fields are brought into juxtaposition remains most useful.

This last and other mechanical features should at once be developed for the colorimeter devised by Gillespie (1921) which promises to be of very great value in exact indicator work. The principle of Gillespie's colorimeter is shown in figure 10. The vessels A, B, C, D and E are of colorless glass the bottoms of which should be optically polished plane-parallel. A and C are

fixed while B may be moved up or down. The position of B is indicated on a scale the zero mark of which corresponds to the position of B when B and C are in contact and the 100 mark of which corresponds to the position of B when B is in contact with A. If now there is placed in B a solution of the acid form of an indicator and in C a solution of the same concentration of the indicator transformed completely to the alkaline form, it is obvious that the position of the vessel B will determine the ratio of the two forms of the indicator which will be within the view.

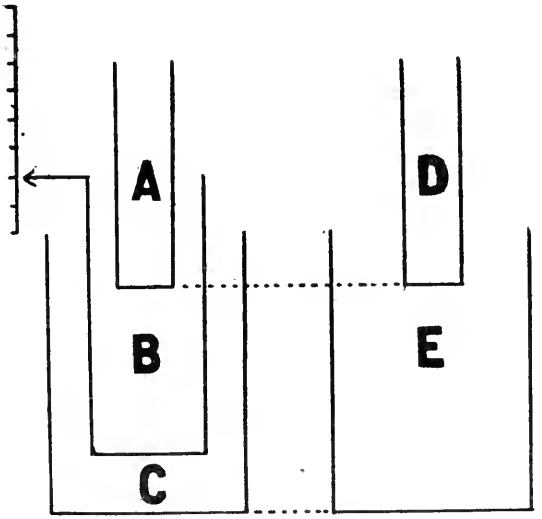


FIG. 10. DIAGRAMMATIC SECTION OF GILLESPIE'S COLORIMETER

For comparison studies a solution to be tested is placed in E together with that concentration of indicator that occurs in the optical system B-C. For colored solutions tubes A and D are used as in the Walpole system, which will presently be described. As Gillespie has indicated this colorimeter should be useful for certain general work where the exact principles of colorimetry have often been neglected.

There have been two chief methods of dealing with the interfering effect of the natural color of solutions. The first method, used by Sørensen, consists in coloring the standard comparison solutions until their color matches that of the solution to be tested, and subsequently adding to each the indicator.

Sørensen's coloring solutions are the following:

a. Bismarck brown (0.2 gram in 1 litre of water).
b. Helianthin II (0.1 gram in 800 cc. alcohol, 200 cc. water).
c. Tropeolin O (0.2 gram in 1 litre of water).
d. Tropeolin OO (0.2 gram in 1 litre of water).
e. Curcumein (0.2 gram in 600 cc. alcohol, 400 cc. water).
f. Methyl violet (0.02 gram in 1 litre of water).
g. Cotton blue (0.1 gram in 1 litre of water).

The second method was introduced by Walpole (1910). It consists in superimposing a tube of the colored solution over the standard comparison solution to which the indicator is added, and comparing this combination with the tested solution plus indicator superimposed upon a tube of clear water.

A somewhat crude but nevertheless helpful application of Walpole's principle may be made from a block of wood. Six deep holes just large enough to hold ordinary test tubes are bored parallel to one another in pairs. Adjacent pairs are placed as close to one another as can be done without breaking through the intervening walls. Perpendicular to these holes and running through each pair are bored *smaller* holes through which the test tubes may be viewed. The center pair of test tubes holds first the solution to be tested plus the indicator and second a water blank. At either side are placed the standards colored with the indicator and each backed by a sample of the solution under test. This is the so called "comparator" of Hurwitz, Meyer, and Ostenberg (1915). Before use it is well to paint the whole block and especially the holes a non-reflecting black. To produce a "dead" black use a soft wood and an alcohol wood-stain. This simple comparator is illustrated in figure 7.

One or another of the means described serves fairly well in overcoming the confusing influence of moderate color in solutions to be tested. In bacteriological work, however, a most serious difficulty is presented by the suspension of cells and precipitates.

If one views lengthwise a tube containing suspended particles, or even particles of grosser colloid dimensions, much of the light incident at the bottom is absorbed or reflected before it reaches the eye, and, if the tube is not screened, some of the light which reaches the eye is that which has entered from the side and has been scattered. Consequently, a comparison with a clear standard is inadequate.

Sørensen (1909) has attempted to correct for this effect by the use of a finely divided precipitate suspended in the comparison solution. This he accomplishes by forming a precipitate of $BaSO_4$ through the addition of chemically equivalent quantities of $BaCl_2$ and Na_2SO_4. Strictly speaking, this gives an imperfect imitation, but like the attempt to match color it does very well in many instances. The Walpole superposition method may be used with turbid solutions as well as with colored, as experience with the device of Hurwitz, Meyer and Ostenberg has shown. In passing, attention should be called to the fact that the view of a turbid solution should be made through a relatively thin layer. When the comparison is made in test tubes, for instance, the view should be from the side.

There are some solutions, however, which are so dark or turbid that they cannot be handled with much precision by any of these methods. On the other hand a combination of these methods with moderate and judicious dilution [as was indicated in Chapter II this may not seriously alter the pH of a solution], permits very good estimates with solutions which at first may appear impossible. Some of the deepest colored solutions permit reasonably good determinations and when sufficiently transparent permit the application of spectrometric devices. Turbidity on the other hand is sometimes unmanageable. Even in the case of milk where comparison with a standard is out of the question a two colored indicator presents a basis for judgment.

This brings us to a phase of the question the detailed analysis of which will not be attempted. It may simply be stated as a fact of experience that the color change of a two-color indicator, presenting as it does change in intensities of what we may summarily describe as two colors, is a change in *quality* which is unmistakable within narrow limits. When there is added to this that brilliancy which is characteristic of the sulfon phthalein indicators the subjective aspect of indicator work is taken care of in a way that may surprise one.

The spectrophotometer and allied instruments which have served in many of the investigations of indicators have not yet been brought within the range of ordinary colorimetric procedure for the determination of pH. Where there occurs a great change in the absorption bands, as at the endpoint of a titration, the hand

spectroscope may be applied but it is doubtful if such an instrument is of much value for slight differences of virage. For the possibilities which remain for development in this field the reader is referred to the special literature.

This brief sketch of some of the principal aspects of indicator theory would be incomplete were attention not called to the value of indicators for demonstrating to students important relations among acids and bases. Indicators also call our attention to molecular transformations which we seldom think of as occurring among substances the light absorptions of which are in regions of the spectrum beyond the reach of the eye.

And finally, indicator colors bring to the thoughtful observer their own intrinsic beauty and also reminders of how far we have come along the road of understanding and of how very, very far we still have to go.

CHAPTER V

Choice of Indicators

From the enormous number of colored compounds found in nature and among the products of the laboratory many have been called into use as acidimetric-alkalimetric indicators. Among those of plant origin litmus and alizarine are the more familiar. One indicator of animal origin, cochineal, an extract of an insect, was formerly used to some extent. Walpole's (1913) treatment of litmus, Walbum's (1913) study of the coloring matter of the red cabbage and some of the more recent work, has given us a little data on properties of plant and animal pigments which are applicable to hydrogen ion determinations. But for the most part indicators of natural origin have been neglected for the study of synthetic compounds.

Litmus has played so important a rôle in acidimetry that it is worthy of brief, special mention.

Litmus is obtained by the oxidation in the presence of ammonia of the orcin contained in lichens, generally of the species *Roccella* and *Lecanora*. The material which comes upon the market is frequently in the form of cubes composed of gypsum or similar material and comparatively little of the coloring matter. The coloring matter is a complex from which there have been isolated many compounds, chief among which are azolitmin, erythrolitmin, erythroleïn and spaniolitmin. Of these the azolitmin is the most important; but the azolitmin of commerce is of uncertain composition, Scheitz (1910). The composition of the different preparations varies with the source and also with the extent of the action of alkali and air upon the crude material.

The following method of preparing a sensitive litmus solution is taken from Morse (1905).

The crushed commercial litmus is repeatedly extracted with fresh quantities of 85 per cent alcohol for the purpose of removing a violet coloring matter which is colored by acids but not made blue by alkalies. The residue, consisting mainly of calcium carbonate, carbonates of the alkalies and the material to be isolated, is washed with more hot alcohol upon a filter

and then digested for several hours with cold distilled water. The filtered aqueous extract has a pure blue color and contains an excess of alkali, a part of which is in the form of carbonate and a part in combination with litmus. To remove the alkaline reaction the solution is heated to the boiling point and cautiously treated with very dilute sulfuric acid until it becomes very distinctly and permanently red. Boil till all CO_2 is dispelled. Treat with a dilute solution of barium hydroxide until the color changes to a violet. Filter, evaporate to a small volume and precipitate the litmus with strong alcohol. Wash with alcohol and dry.

Dr. P. Rupp (private communication) prefers to make a final washing with water which removes much of the salt at the expense of some dye.

Synthetic indicators have for the most part displaced those of natural origin until litmus and alizarine, turmeric and cochineal are becoming more and more unfamiliar in the chemical laboratory. Indeed Bjerrum (1914) states that the two synthetic indicators, methyl red and phenolphthalein, particularly because of the zones of hydrogen ion concentration within which they change color, are sufficient for most titrimetric purposes.

But the two indicators mentioned above cover but a very limited range of hydrogen ion concentration so that they are insufficient for the purpose we now have under consideration. A survey of indicators suitable for hydrogen ion determinations was opened in Nernst's laboratory in 1904 by Salessky. This survey was extended in the same year by Friedenthal, by Fels and by Salm and the results were summarized in Salm's famous table (cf. *Z. physik. Chem.*, **57**).

Then came the classic work of Sørensen of the Carlsberg laboratory in Copenhagen. The array of available indicators had become so large as to be burdensome. Sørensen in an extensive investigation of the correspondence between colorimetric and electrometric determinations of hydrogen ion concentrations revealed discrepancies which were attributed mainly to the influence of protein and salts. He chose those indicators which were relatively free from the so-called protein and salt errors, constructed solutions of known and reproducible hydrogen ion concentrations and thus furnished the biochemist with selected tools of beautiful simplicity. It is well to emphasize the labor of elimination which Sørensen performed because without it we might still be consulting such tables as that published by Thiel (1911), or the

ponderous tables 8–19, pages 84–94, and be bewildered by the very extensive array.

Sørensen's work, coupled as it was with a most important contribution to enzyme chemistry gave great impetus to the use of indicators in biochemistry. His selection of indicators was therefore soon enlarged by additions of new indicators which fulfilled the criteria of reliability which he had laid down. Alpha naphthol phthalein, a compound first synthesized by Grabowski (1871), was shown by Sørensen and Palitzsch (1910) to have a range of pH 7–9 and was found useful in biological fluids. Methyl red (Rupp and Loose, 1908) was given its very useful place by the investigations of Palitzsch (1911). Henderson and Forbes (1910) introduced 2–5 di nitro hydroquinone as an indicator possessing several steps of color change and therefore useful over a wide range of pH. Walpole (1914) called attention to several indicators of potential value. Hottinger (1914) recommended "lacmosol," a constituent of lacmoid, and Scatchard and Bogert (1916) advocated the use of dinitro benzoylene urea. There remain a host of indicators which have been tried out in the empirical practices of titration but which have never had their pH ranges determined; and there remain an unlimited number of possibilities embodied in existing compounds such as Dox's (1915) phenol quinolinein, Rupp's (1915) syntheses in the methyl red series and untouched homologues of phenol phthalein and of phenol sulfon phthalein. Furthermore, there undoubtedly are still unsynthesized compounds of various types, old and new, which will some day displace those now in use.

In 1915 Levy, Rowntree and Marriott, without applying the tests of reliability which Sørensen had employed, used phenol sulphon phthalein in determining the pH of the dialyzate of blood. This compound, first synthesized in Remsen's laboratory by Sohon (1898), has received considerable attention from Acree and his co-workers because it furnishes excellent material for the quinone-phenolate theory of indicators. To further such studies Acree and White had synthesized new derivatives of phenol sulphon phthalein at the time when the work of Levy, Rowntree and Marriott attracted the attention of Clark and Lubs. These authors were looking for more brilliant indicators for use in bacterial culture media and were attracted by the well known brilliance of

phenol sulphon phthalein. Through the courtesy of Professor Acree some of the derivatives which White had prepared were obtained. New homologues were synthesized by Lubs. The applicability of these and numerous other indicators in the determination of the pH values of biological fluids was then studied.

In the sulfon phthalein series the following were studied:

Phenol sulfon phthalein, Sohon (1898).

Tetra nitro phenol sulfon phthalein, White and Acree (1915).

Phenol nitro sulfon phthalein, Lubs and Clark (1915).

Tetra bromo phenol sulfon phthalein, White and Acree (1915).

Tetra chloro phenol sulfon phthalein, Lubs and Clark.

Ortho cresol sulfon phthalein, Sohon (1898).

Di bromo ortho cresol sulfon phthalein, Sohon (1898).

Thymol sulfon phthalein, Lubs and Clark (1915).

Thymol nitro sulfon phthalein, Lubs and Clark.

Di bromo thymol sulfon phthalein, Lubs and Clark (1915).

α-napthol sulfon phthalein, Lubs and Clark (1915).

Carvacrol sulfon phthalein, Lubs and Clark.

Orcinol sulfon phthalein, Gilpin (1894).

The attractiveness of methyl red led to the study of the following compounds:

o-carboxy benzene azo mono methyl aniline, Sive and Jones (1915).

o-carboxy benzene azo di methyl aniline, Rupp and Loose (1908).

o-carboxy benzene azo mono ethyl aniline, Lubs and Clark (1915).

o-carboxy benzene azo di ethyl aniline, Lubs and Clark (1915).

o-carboxy benzene azo mono propyl aniline, Lubs and Clark (1915).

o-carboxy benzene azo di propyl aniline, Lubs and Clark (1915).

o-carboxy benzene azo (?) amyl aniline, Lubs and Clark (1915).

o-carboxy benzene azo di methyl α naphthyl amine, Howard and Pope (1911).

o-carboxy benzene azo α naphthyl amine, Howard and Pope (1911).

o-carboxy benzene azo di phenyl amine, Howard and Pope (1911).

Meta carboxy benzene azo di methyl aniline, Lubs and Clark.

The mono alkyl homologues of methyl red were found to be much less brilliant than the di alkyl compounds and were therefore rejected. For the same reason or because of large protein errors we rejected the other compounds with the exception of di ethyl and di propyl red. Of these we retained di propyl red because it is very useful in solutions of a little lower hydrogen ion concentration than those which may be studied with methyl red.

Propyl red is, however, not included in table 6 because it precipitates too easily from buffer solutions to be of general usefulness.

As the result of an extensive series of comparisons between colorimetric and electrometric measurements, made for the most part upon solutions of interest to bacteriologists, Clark and Lubs (1917) suggested the series of indicators given in table 6. This series is made up for the most part of the brilliant and more reliable sulfon phthaleins but contains the still indispensable but not very stable methyl red.

In the course of their investigations these authors resurrected ortho cresol phthalein (Baeyer and Freude, 1880), found it quite as reliable as phenolphthalein and more brilliant with a color better adapted to titrations in artificial light.

In spite of the fact that Sørensen rejected the greater number of the indicators which he studied and that Clark and Lubs, after a resurvey of the subject and the preparation of many new compounds, listed but few indicators as reliable, there has recently appeared a tendency to resurrect the rejects. Now many of these are useful in special cases and undoubtedly there is an occasional individual to be found in the lists which has been insufficiently studied and unjustly rejected. Nevertheless, the indiscriminate use of miscellaneous indicators may lead to gross errors or at least to such a diversity of data that their correlation will become complex during the coming period when the specific salt-errors and general conduct of the individual indicators are still being worked up.

It is therefore advisable to use the more thoroughly studied lists. Three such lists are given (tables 5, 6 and 7). The indicators therein listed should cover all ordinary needs. Sørensen's list is given in table 5 and to this is appended Sørensen's comments. For general purposes the selection of indicators given

in table 6 will be found the most satisfactory especially because of their brilliancy. Each of these however has its own special limitations as every indicator has. For the study of colorless solutions where salt errors are to be reduced the nitro phenols listed in table 7 should be valuable.

TABLE 5

Sørensen's selected indicators

Figures in parentheses refer to Schultz (1914). Figures 1–20 are Sørensen's

INDICATOR	pH RANGE
1. Methyl violet 6B extra, (517)	0.1– 3.2
2. Mauvein, Rosolane, (688)	0.1– 2.9
3. Diphenylamino-azo-benzene	1.2– 2.1
4. Diphenylamino-azo-p-benzene sulfonic acid, Tropaeolin OO, (139)	1.4– 2.6
5. Diphenylamino-azo-m-benzene sulfonic acid, Metanil yellow, (134)	1.2– 2.3
6. Benzyl anilino-azo-benzene	2.3– 3.3
7. Benzylanilino-azo-p-benzene sulfonic acid	1.9– 3.3
8. Metachloro diethyl-anilino-azo-p-benzene sulfonic acid	2.6– 4.0
9. Dimethyl anilino-azo-benzene, (32)	2.9– 4.0
10. Methyl orange, Helianthine, (138)	3.1– 4.4
11. α naphthylamino-azo-benzene	3.7– 5.0
12. α-naphthylamino-azo-p-benzene sulfonic acid	3.5– 5.7
13. Para nitro phenol	5.0– 7.0
14. Neutral red, (670)	6.8– 8.0
15. Rosolic acid, Aurin, (555)	6.8– 8.0
16. Orange I, Tropaeolin OOO No. 1, (144)	7.6– 8.9
17. Phenolphthalein	8.3–10.0
18. Thymolphthalein	9.3–10.5
19. Paranitrobenzene-azo-salicylic acid, Alizarine yellow R, (58)	10.1–12.1
20. Resorcin-azo-p-benzene sulfonic acid, Tropaeolin O, (143)	11.1–12.7

In tables 8–20 are a few indicators which are undoubtedly reliable but little used, a few which are definitely unreliable though often used, and very many of uncertain character but for the most part bearing the stamp of disapproval by competent judges. Since the indicators in tables 5, 6 and 7 cover all ordinary requirements it seems hardly worth while to venture upon an analysis of the remaining tables.

In table 5 is Sørensen's list of indicators; concerning these indicators Sørensen remarks:

Not all these indicators furnish equally well defined virages and above all they are not of equal applicability under all circumstances. In the choice of an indicator from among those which we have been led to recommend it is necessary to use judicious care and especially to take into consideration the following facts:

a. The indicators of the methyl violet group (nos. 1 and 2) (see table 5) are especially sensitive to the action of neutral salts; furthermore the intensity of color changes on standing and the change is the more rapid the more acid the medium.

b. The basic indicators (nos. 3, 6, 9, 11, 14) are soluble in toluene and in chloroform. The first four separate partially on prolonged standing of the experimental solution.

c. In the presence of high percentages of natural proteins most of the indicators are useless although certain of them are still serviceable; nos. 1, 2, 13, 16, 17, 18.

d. In the presence of protein decomposition products even in considerable proportions the entire series of indicators may render real service. Yet even in these conditions some of the acid azo indicators may give notable errors (nos. 4, 5, 7, 8, 10) in which case one should resort to the corresponding basic indicators.

e. When only small percentages of protein or their decomposition products are concerned the acid azo indicators are more often preferable to the basic for they are not influenced by toluene or chloroform and do not separate from solution on standing.

f. In all doubtful cases—for example in the colorimetric measurement of solutions whose manner of reaction with the indicator is not known, the electrometric measurement as a standard method should be used. Then the worth of the indicator will be determined by electrometric measurement with colorimetric comparison.

In table 6 will be found the final selection of Clark and Lubs with the common names which they suggested for laboratory parlance, the concentration of indicator convenient for use, a rough indication of the nature of the color, and the useful pH range.

With the improved method for the preparation of the sulfon phthalein indicators described by Lubs and Clark (1915) they may easily be made from materials readily obtained. The indicators can also now be purchased in this country and abroad from chemical supply houses.

The indicators recommended by Clark and Lubs are marketed both in the form of a dry powder and in stock solutions. In cases where the acidity of the free acid dye in the indicator solution

does not interfere with accuracy and when alcohol is not objectionable the alcoholic solutions of the dyes may be used. Clark and Lubs prefer to use aqueous solutions of the alkali salts in concentrations which may be conveniently kept as stock solutions. These are diluted for the test solutions used in the dropping bottles.

TABLE 6

Clark and Lubs' list of indicators

CHEMICAL NAME	COMMON NAME	CONCENTRATION	COLOR CHANGE	RANGE pH
		per cent		
Thymol sulfon phthalein (acid range)	Thymol blue (see below)	0.04	Red-yellow	1.2–2.8
Tetra bromo phenol sulfon phthalein	Brom phenol blue	0.04	Yellow-blue	3.0–4.6
Ortho carboxy benzene azo di methyl aniline	Methyl red	0.02	Red-yellow	4.4–6.0
Di bromo ortho cresol sulfon phthalein	Brom cresol purple	0.04	Yellow-purple	5.2–6.8
Di bromo thymol sulfon phthalein	Brom thymol blue	0.04	Yellow-blue	6.0–7.6
Phenol sulfon phthalein	Phenol red	0.02	Yellow-red	6.8–8.4
Ortho cresol sulfon phthalein	Cresol red	0.02	Yellow-red	7.2–8.8
Thymol sulfon phthalein	Thymol blue	0.04	Yellow-blue	8.0–9.6
Ortho cresol phthalein	Cresol phthalein	0.02	Colorless-red	8.2–9.8

For the preparation of these stock solutions one decigram (0.1 gram) of the dry powder is ground in an agate mortar with the following quantities of N/20 NaOH. When solution is complete dilute to 25 cc. with water.

MOLECULAR WEIGHT	INDICATOR	N/20 NaOH PER DECIGRAM
		cc.
354	Phenol red	5.7
669	Brom phenol blue	3.0
382	Cresol red	5.3
540	Brom cresol purple	3.7
466	Thymol blue	4.3
624	Brom thymol blue	3.2
269	Methyl red	7.4

If there be no particular reason to maintain exact equivalents it may be found easier to dissolve the dyes in 1.1 equivalents of alkali instead of one equivalent as indicated above.

When made up to 25 cc. as noted above there is obtained in each case a 0.4 per cent solution of the original dye itself. For tests they should be diluted further. To place the dyes upon a comparable basis the final dilution should be nearly the same when calculated upon a molar basis and, by reason of the great change in molecular weight caused by the introduction of bromine and other group substituents, equal molecular concentrations will be very far apart in percentage concentration. For all ordinary purposes, however, this may be neglected and the solutions mentioned above if diluted in each case to a concentration of 0.04 per cent will be found satisfactory for use in testing 10 cc. of a solution with about five drops of indicator.

From various sources have come complaints that the method outlined above for the preparation of the aqueous alkali salt solution of brom cresol purple leads to a solution of much lower tinctorial power than when the same material is taken up directly in alcohol. No such difficulty was experienced with the material described by Lubs and Clark but it has appeared not infrequently since. The source of the difficulty is not yet definitely traced, but is suspected to be due to impurities. If so it should be avoided by purchasing the highly purified material which is now made specially.

While the aqueous alkali salt solution of methyl red is preferred for some purposes a methyl red solution can be more conveniently prepared by dissolving 1 decigram in 100 cc. alcohol and diluting to 200 with distilled water.

Ortho cresol phthalein and phenol phthalein are used in a 0.04 per cent solution of 95 per cent alcohol.

Methyl red and brom cresol purple may be recrystallized from hot toluol, cresol red and brom phenol blue from glacial acetic acid, thymol blue from hot alcohol.

Tables 8–20 have been compiled with the aid of Dr. Barnett Cohen and Dr. Elias Elvove with several purposes in view. In the first place there exist in the older literature a great many observations recorded in terms of the color of a given indicator. These data can often be translated into modern terms if the pH range of the given indicator is known. In the second place there

TABLE 7

Michaelis' indicators and their ranges as used in the method of Michaelis and Gyemant (see Chapter VIII)

Picric acid	colorless	0.0– 1.3 yellow
2, 4-dinitro phenol	colorless	2.0– 4.7 yellow
α dinitro phenol		
2, 6-dinitro phenol	colorless	1.7– 4.4 yellow
β dinitro phenol		
2, 5-dinitro phenol	colorless	4.0– 6.0 yellow
γ-dinitro phenol		
m-nitro phenol	colorless	6.3– 9.0 yellow
p-nitro phenol	colorless	4.7– 7.9 yellow
Phenolphthalein	colorless	8.5–10.5 red
Alizarine yellow GG	colorless	10.0–12.0 yellow
Salicyl yellow		

are circumstances when for one reason or another it becomes necessary to draw upon the miscellaneous list. It should therefore be available. Lastly, and perhaps most important, our review of the literature and of indicator labeling has shown that there is great confusion and an initial step in the clarification of the subject will be taken if there is available a tabulation of existing data to serve as a basis for revision.

In examining a large collection of indicators the labeling was found to be insufficient in a large percentage of cases. On studying the literature we find evidence that others have encountered the same difficulty without stating so, for in many instances the indicator names given were evidently those

of one or another dealer who cared so little for the scientific uses of his commodity that he left from the label the designation essential to its identification. This habit has become more or less prevalent. In some instances our own uncertainty may be due to an arbitrary adherence to the nomenclature found in various editions of Schultz. For instance when we see the indicator croceïne listed and refer to Schultz (1914) we find four croceïnes with various distinguishing marks and seven other compounds for the names of which "croceïne" is used in one or another combination. But Schultz lists no croceïne. We are not helped in going back to the lists of Schultz and Julius (1902). Now we might assume that "croceïne" was used in Salm's table as a term having a definite meaning outside the dye industry. On this principle we should find that "helianthine" has been employed in accordance with scientific usage. However we find that an old sample of helianthine from Salm's dealer is not the helianthine of methyl orange but corresponds in pH-range to Salm's Helianthine I, which, together with Salm's Helianthine II we have not identified.

Again there are other difficulties such as are illustrated by the case of Tropaeolin OOO No. 1 and Tropaeolin OOO No. 2. No. 1 is prepared from p-sulfanilic acid and *α-naphthol.* No 2 is prepared from p-sulfanilic acid and *β-naphthol.* In this there is agreement by Schultz and Julius 1902, Green 1904 and Beilstein (third edition). In accord with this Sørensen describes his α-naphthol preparation as Tropaeolin OOO No. 1. In the second edition of *Indicators and Test Papers,* Cohn (1914) has given synonyms for the α and β compounds which agree with Green, but has reversed the No. 1 and No. 2 at the headings of his descriptions and uses "No. 1" and "No. 2" inconsistently in the text. Prideaux (1917) has called the β compound Tropaeolin OOO and gives the range as 7.6–8.9, which looks suspiciously like Sørensen's 7.6–8.9 for the α compound. Prideaux uses the synonym Orange II for the β compound in harmony with Green but on the next page describes the α compound as Orange II. The identity of Salm's Tropaeolin OOO is not clear. It was evidently different from the Tropaeolin OOO No. 1 used by Sørensen. We find that an old sample with the label "Tropaeolin OOO" agrees with neither Sørensen's nor Salm's data.

Many other instances might be cited to show the confused

state of the subject. Because it is serious the reader will have to use the following tables with caution, and he need not be surprised if a sample of indicator which he tests does not give a pH range corresponding to that recorded.

In the compilation of the lists we have followed competent advice in using the nomenclature of *Farbstofftabellen*, Gustav

TABLE 8

Nitro compounds

SERIAL NUMBER	INDICATOR	pH RANGE
1	Picric acid (5)..... 2, 4, 6-trinitro-phenol	colorless 0.0– 1.3 yellow
2	2, 6-Dinitro-phenol (β).....	colorless 1.7– 4.4 yellow
3	Martius yellow (6)..... 2, 4-dinitro-α-naphthol	light yellow 2.0– 4.0 yellow
4	2, 4-Dinitro-phenol (α).....	colorless 2.0– 4.7 yellow
5	2, 5-Dinitro-hydroquinone....	3.0– 9.0 various colors
6	2, 3-Dinitro-phenol (ε).....	colorless 3.9– 5.9 yellow
7	2, 5-Dinitro-phenol (γ).....	colorless 4.0– 6.0 yellow
8	*iso*-Picramic acid..... 2, 6-dinitro-4-amino-phenol	pink 4.1– 5.6 yellow
9	3, 4-Dinitro-phenol (δ).....	colorless 4.3– 6.3 yellow
10	p-Nitro-phenol.....	colorless 5.0– 7.0 yellow
11	Dinitrobenzoylene-urea.....	colorless 6.0– 8.0 yellow
12	m-Nitro-phenol.....	colorless 6.3– 9.0 yellow
13	Nitramine (?).....	11.0–12.5
14	1, 3, 5-Trinitro-benzene.....	colorless 11.5–14.0 orange
15	2, 4, 6-Trinitro-toluene (TNT)	pink 11.5–14.0 orange

Schultz, fifth revised edition, Berlin, 1914. In a few cases there have been added to the synonyms in table 20 terms which are obsolete in the dye industry but which are still used in the nomenclature of indicators. Schultz numbers are to be found in tables 8 to 19 following the name of each indicator when the given indicator is listed by Schultz. Since it is unimportant for indicator work, no distinction has been made between acids and their salts. The classification by structure follows in the main that of Schultz.

TABLE 9

Monoazo compounds

SERIAL NUMBER	INDICATOR	pH RANGE
16	Curcumein (?).....	orange 0.0– 1.0 yellow, yellow 13–15 green
17	o-Carboxybenzene-azo-(di or mono?) amyl-aniline...	purple 0.0– 1.6 orange (fluorescent), orange 5.6–7.6 yellow
18	o-Carboxybenzene-azo-m-phenylenediamine.....	yellow 0.0– 4.6 orange, orange 4.6–7.6 yellow
19	p-Toluene-azo-phenyl-aniline.	1.0– 2.0
20	p-Carboxybenzene-azo-dimethyl-aniline (Para methyl red).....	red 1.0– 3.0 yellow
21	p-Toluene-azo-phenyl-α-naphthylamine.....	1.1– 1.9
22	Benzene-azo-diphenylamine..	1.2– 2.1
23	Metanil yellow extra (134).... m-sulfobenzene-azo-diphenylamine	red 1.2– 2.3 yellow
24	Benzene-azo-phenyl-α-naphthylamine	1.4– 2.6
25	Orange IV (139)..... p-sulfobenzene-azo-diphenylamine	pink 1.4– 2.6 yellow
26	o-Toluene-azo-o-toluidine....	1.4– 2.9
27	p-Toluene-azo-benzyl-α-naphthylamine.....	1.6– 2.6
28	p-Toluene-azo-benzyl-aniline.	1.6– 2.8
29	Benzene-azo-benzyl-α-naphthylamine.....	1.9– 2.9
30	Amino-azo-benzene (31).....	light yellow 1.9– 3.3 yellow
31	p-Benzenesulfonic acid-azo-aniline.....	1.9– 3.3
32	p-Benzenesulfonic acid-azo-benzyl-aniline.....	1.9– 3.3
33	m-Carboxybenzene-azo-dimethylaniline.....	red 2.0– 4.0 yellow
34	Benzene-azo-benzyl-aniline...	2.3– 3.3

TABLE 9—*Continued*

SERIAL NUMBER	INDICATOR	pH RANGE
35	p-Benzenesulfonic acid-azo-metachloro-dimethyl-aniline	2.6– 4.0
36	Orange III (47) m-nitrobenzene-azo-β-naphthol-3, 6-disulfonic acid	red 2.6– 4.6 yellow
37	Butter yellow O (32) benzene-azo-dimethyl-aniline	red 2.9– 4.0 yellow
38	o-Carboxybenzene-azo-diphenylamine	pink 3.0– 4.6 yellow, purple 0.0–1.6 pink
39	p-Benzenesulfonic acid-azo-methyl-aniline	3.1– 4.2
40	p-Benzenesulfonic acid-azo-ethyl-aniline	3.1– 4.4
41	Methyl orange (138) p-benzenesulfonic acid-azo-dimethyl-aniline	orange red 3.1– 4.4 yellow
42	p-Benzenesulfonic acid-azo-diethyl-aniline (Ethyl orange)	pink 3.5– 4.5 yellow
43	p-Benzenesulfonic acid-azo-α-naphthylamine	3.5– 5.7
44	Benzene-azo-α-naphthyl-amine	3.7– 5.0
45	p-Toluene-azo-α-naphthylamine	3.7– 5.0
46	o-Carboxybenzene-azo-monomethylaniline	red 4.0– 6.0 yellow
47	Chrysoïdin (33) benzene-azo-m-phenylenediamine	orange 4.0– 7.0 yellow
48	o-Carboxybenzene-azo-monoethylaniline	red 4.2– 6.2 yellow
49	o-Carboxybenzene-azo-mono-n-propylaniline	red 4.2– 6.2 yellow
50	o-Carboxybenzene-azo-dimethylaniline (Methyl red)	red 4.2– 6.3 yellow
51	o-Carboxybenzene-azo-diethylaniline	red 4.4– 6.2 yellow

TABLE 9—*Concluded*

SERIAL NUMBER	INDICATOR	pH RANGE
52	o-Carboxybenzene-azo-di-n-propylaniline (Propyl red)	red 4.6– 6.6 yellow
53	Benzene-azo-dimethyl-α-naphthylamine	4.8– 5.5
54	p-Benzenesulfonic acid-azo-dimethyl-α-naphthylamine	5.0– 5.7
55	o-Carboxybenzene-azo-α-naphthylamine	pink 5.6– 7.0 yellow
56	o-Carboxybenzene-azo-dimethyl-α-naphthylamine	red 5.6– 7.6 orange
57	Naphthylamine brown (160).. 4-sulfonaphthalene-azo-α-naphthol	orange 6.0– 8.4 pink
58	6-Sulfo-α-naphthol-1-azo-m-hydroxybenzoic acid	orange 7.0– 8.0 blue, violet 12–13 red
59	Orange I (144) p-sulfobenzene-azo-α-naphthol	7.6– 8.9
60	Orange II (145) p-sulfobenzene-azo-β-naphthol	7.6– 8.9 (?)
61	Alizarine yellow GG (48) m-nitrobenzene-azo-salicylic acid	colorless 10.0–12.0 yellow
62	Alizarine yellow R (58) p-nitrobenzene-azo-salicylic acid	pale yellow 10.1–12.1 orange
63	Fast red A (161) 5-sulfonaphthalene-azo-β-naphthol	10.5–12.1
64	Fast red B (112) α-naphthalene-azo-β-naphthol-3, 6-disulfonic acid	pink 10.5–12.5 orange
65	Chrysoïn (143) p-sulfobenzene-azo-resorcin	yellow 11.1–12.7 orange
66	Orange G (38) benzene-azo-β-naphthol-γ-disulfonic acid	yellow 11.5–14.0 pink

TABLE 10

Disazo compounds

SERIAL NUMBER	INDICATOR	pH RANGE
67	Benzopurpurin B (365)..... ditolyl-disazo-bi-β-naphthylamine-β-sulfonic acid	blue—0.3– 1.0 violet, violet 1.0–5.0 yellow, yellow 12.0–14.0 rose
68	Congo (307)..... diphenyl-disazo-binaphthionic acid	blue 3.0– 5.0 red
69	Azo blue (377)..... ditolyl-disazo-bi-α-naphthol-4-sulfonic acid	violet 10.5–11.5 pink

TABLE 11

Triphenylmethane compounds

SERIAL NUMBER	INDICATOR	pH RANGE
70	Crystal violet (516)..... hexamethyl pararosaniline	green 0.0– 2.0 blue
71	Malachite green (495)..... tetramethyl-di-p-amino-triphenyl-carbinol	yellow 0.0– 2.0 green, blue 11.5–14.0 fades
72	Red violet 5R extra (514).... mixture of mono-, di- and tri-methyl or ethyl rosanilines and pararosanilines	green 0.0– 2.0 blue
73	Brilliant green (499)..... tetraethyl-di-p-amino-triphenyl-carbinol	yellow 0.0– 2.6 green
74	Iodine green..... heptamethyl rosaniline	yellow 0.0– 2.6 blue
75	Ethyl violet (518)..... hexaethyl pararosaniline	yellow 0.0– 3.6 blue
76	Ethyl green (methyl green)... ethyl-hexamethyl-pararosaniline bromid	0.1– 2.3

TABLE 11—*Continued*

SERIAL NUMBER	INDICATOR	pH RANGE
77	Methyl violet 6B extra (517).. mixture of benzyl-tetra- and pentamethyl-p-rosaniline and hexamethyl-p-rosaniline	0.1– 3.2
78	Fuchsin (512) (base)..... mixture of rosaniline and pararosaniline	purple 1.2– 3.0 fades
79	Red violet 5RS (525)..... trisulfonate of ethyl rosaniline	pink 3.6– 6.0 colorless
80	Water blue (539)..... di- and tri-sulfonic acids of triphenyl-p-rosaniline and di-phenyl-rosaniline	blue 4.7– 7.0 colorless,* purple 10.5–14.0 rose
81	Aurin (p-rosolic acid) (555)... complex mixture	yellow 6.9– 8.0 red
82	Alkali blue (536)..... mixture of diphenyl-rosaniline-mono-sulfonic acid and triphenyl-pararosaniline-mono-sulfonic acid	lilac 9.4–14.0 pink
83	Methyl blue (538)..... triphenylpararosaniline-di- and trisulfonic acids	blue 10.0–13.0 pink
84	Fuchsin S (524)..... di- and trisulfonic acids of rosaniline and p-rosaniline	red 12.0–14.0 fades

* Samples of Water blue (China blue) which we have tested vary considerably. The color change in the neutral range is instantaneous with some samples but requires a long period (several hours at room temperature) for others.

TABLE 12

Quinoline compounds

SERIAL NUMBER	INDICATOR	pH RANGE
85	Quinoline blue (Cyanin) (611). $C_{28}H_{35}N_2I$	colorless 7.0–8.0 violet

TABLE 18

Oxazine compounds

SERIAL NUMBER	INDICATOR	pH RANGE
86	Alizarin green B (657)..... dihydroxy-naphth-azoxonium sulfonate	lilac—0.3– 1.0 flesh, brownish yellow 12.0–14.0 brown, then green
87	Nile blue 2B (654)..... diethyl-benzyl-diamino-naphtho-phenazoxonium chlorid	blue 7.2– 8.6 rose
88	Nile blue A (653)..... diethyl-diamino-naphtho-phenazoxonium sulfate	blue 10.2–13.0 rose

TABLE 14

Azines

SERIAL NUMBER	INDICATOR	pH RANGE
89	Methylene violet BN powder (680)..... dimethyl-diamino-phenyl-phenazonium chloride	purple 0.0– 1.2 violet
90	Rosolane (688)..... phenyl and tolyl safranines	0.1– 2.9
91	Rose magdala (694)..... mixtures of amino naphthyl-naphthazonium chlorid and diamino-naphthyl-naphthazonium chloride	rose 3.0– 4.0 red, lilac 12.0–14.0 violet
92	Indulin, spirit soluble (697).. mixtures of dianilido-amido-tri-anilido- and tetranilido-phenyl-phenazonium chlorides	blue 5.6– 7.0 violet
93	Neutral red (670)..... dimethyl-diamino-tolu-phenazine	red 6.8– 8.0 yellow
94	Neutral blue (676)..... dimethyl-amino-phenyl-phenonaphthazonium chloride	9.3–10.2

TABLE 15

Anthraquinone compounds

SERIAL NUMBER	INDICATOR	pH RANGE		
95	Alizarin Blue X (803)..... dihydroxy-anthraquinone-β-quinoline	pink	0.0–1.6	yellow, yellow 6.0–7.6 green
96	Purpurin (783)..... 1, 2, 4-trihydroxy-anthraquinone	yellow	0.0–4.0	orange, orange 4.0–8.0 rose, lilac 12.0–14.0 violet
97	Alizarin red S (780)..... mono sulfonic acid of alizarin V_1	yellow	5.0–6.8	pink
98	Alizarin V_1 (Alizarine) (778)... 1, 2-dihydroxy-anthraquinone	yellow	5.5–6.8	red, violet 10.1–12.1 purple
99	Alizarin Blue S (804)..... Na bisulfite compound of alizarin blue X	yellow	6.0–8.0	green, green 11.0–13.0 blue

TABLE 16

Indigos

SERIAL NUMBER	INDICATOR	pH RANGE
100	Indigotine Ia in powder (Indigo carmine) (877).... Indigo disulfonate	blue 11.6–14.0 yellow

TABLE 17

Phthalein and xanthone compounds

SERIAL NUMBER	INDICATOR	pH RANGE
101	Rhodamine B (573)..... diethyl m-amino-phenol-phthalein	orange—0.1– 1.2 pink
102	Galleïn (599)..... pyrogallol phthalein	yellow 0.0– 2.6 brown, brown 3.6–7.0 pink, pink 9.4–14.0 purple
103	Eosin G (587)..... tetrabromo fluorescein	orange 0.0– 3.0 pink
104	Erythrosin*.....	orange 0.0– 3.6 pink
105	Phloxin Red BH (Grübler)...	pink 1.4– 3.6 red
106	Uranin (Fluorescein) (585)... resorcin phthalein	light yellow 3.6– 5.6 yellow (fluorescent)
107	Dichloro fluorescein.....	yellow 4.0– 6.6 yellow (fluorescent)
108	o-α-Naphthol phthalein.....	yellowish 7.0– 9.0 green
109	p-α-Naphthol phthalein.....	yellowish 7.0– 9.0 blue
110	Tetrabromophenol phthalein.	colorless 8.0– 9.0 violet
111	o-Cresol phthalein.....	colorless 8.2– 9.8 red
112	Phenol phthalein.....	colorless 8.3–10.0 red
113	1, 2, 3-Xylenol phthalein.....	colorless 8.9–10.2 blue
114	Thymol phthalein.....	colorless 9.3–10.5 blue
115	Eosin BN (590)..... dibromo dinitro fluorescein	pink 10.5–14.0 yellow

* The identity of this erythrosin is in doubt. Erythrosin R, G, yellowish, and Iodeosin G are synonyms of di-iodo-fluorescein. Erythrosin extra bluish, D, B, J extra, JNV, W extra, and Iodeosin B are synonyms for the tetra-iodo-fluorescein.

TABLE 18

Sulfonphthaleins

SERIAL NUMBER	INDICATOR	pH RANGE
116	Di-iodophenol sulfonphthalein*	orange 0.0– 1.2 yellow, yellow 3.2–7.0 purple
117	Catechol sulfonphthalein	pink 0.2– 0.8 orange, yellow 4.0–7.0 green, violet 8.5–10.2 blue, blue 10.2–12.5 green
118	Thymol sulfonphthalein *Thymol blue* (acid range) (alkaline range)	red 1.2– 2.8 yellow†, yellow 8.0–9.6 blue
119	Tetranitrophenol sulfonphthalein	yellow 2.8– 3.8 red
120	Tetrabromophenol sulfonphthalein *Brom phenol blue*	yellow 3.0– 4.6 blue
121	Tetrachlorophenol sulfonphthalein	yellow 3.0– 4.6 blue
122	Dibromo-o-cresol sulfonphthalein *Brom cresol purple*	yellow 5.2– 6.8 purple
123	Dibromothymol sulfonphthalein *Brom thymol blue*	yellow 6.0– 7.6 blue
24	Phenol nitro sulfonphthalein	yellow 6.6– 8.4 purple
25	Phenol sulfonphthalein *Phenol Red*	yellow 6.8– 8.4 red
26	o-Cresol sulfonphthalein *Cresol Red*	yellow 7.2– 8.8 red
27	Salicyl sulfonphthalein	yellow 7.2– 9.2 pink
28	Thymol nitro sulfonphthalein	yellow 7.2– 9.4 blue
29	α-Naphthol sulfonphthalein	yellow 7.5– 9.0 blue
30	Carvacrol sulfonphthalein	yellow 7.8– 9.6 blue
31	Orcin sulfonphthalein	yellow 8.6–10.0 pink (fluorescent)
32	Nitrothymol sulfonphthalein	violet 9.2–11.5 yellow

* Purity not established.

† All sulfonphthaleins show color changes at high acidities but those o thymol sulfonphthalein are the most useful.

TABLE 19

Miscellaneous indicators

SERIAL NUMBER	INDICATOR	pH RANGE
133	Croceine (?)	blue—0.3– 0.0 rose, rose 12.0–14.0 violet
134	Eosin methylene blue	green—0.3– 1.0 blue, violet 14.0–15.0 lilac
135	Safranin (679?)	blue—0.3– 1.0 red, red 14.0–15.0 violet
136	Hematein (Logwood) (938)	variable from 0.0–15.0
137	Gentian violet	0.4– 2.7
138	Red cabbage extract	red 2.4– 4.5 green
139	1-Oxy-naphtho-chino-methane	colorless 2.7– 3.7 purple
140	Troger and Hille's indicator. $C_{14}H_{15}N_4SO_3H$	orange 2.8– 4.0 yellow
141	Phenacetolin	yellow 3.0– 6.0 red, red 10.0–13.0 colorless
142	Lacmosol	red 4.4– 5.5 blue
143	Lacmoid	red 4.4– 6.2 blue
144	Azolitmin (Litmus)	red 4.5– 8.3 blue
145	Carminic acid (from cochineal) (932)	orange 4.6– 7.8 rose, violet 11.0–14.0 pink
146	Cochineal (932)	yellow 4.8– 6.2 lilac
147	Archil (Orchil) (934)	pink 5.6– 7.6 lilac
148	Brazil wood, Red wood, Brasileïn (935)	colorless 6.0– 8.0 pink
149	Guaiac tincture	colorless 7.0– 8.0 greenish
150	Lygosine di-o-hydroxy-styryl ketone	yellow 7.3– 8.7 green
151	Mimosa flower extract	7.7– 9.6
152	Turmeric (Curcuma) (927) $C_{21}H_{20}O_6$	yellow 8.0–10.2 orange
153	Alkanin	red 8.3–10.0 blue
154	α-Naphthol benzein	yellow 8.5– 9.8 green
155	Benzoazurin (?)	purple 10.5–12.0 pink
156	Helianthin I (?)	orange 11.0–12.0 orange red
157	Poirrier's blue	blue 11.0–13 0 red
158	Helianthin II (?)	brownish-yellow 13.0–14.0 lilac

TABLE 20

The more common synonyms of indicators

This table contains the names and synonyms of the various indicators in alphabetical order. Following each name, or group of synonyms, is a number in **bold face** type. This number is the serial number of the compound as found in the preceding tables.

Some names apply to two or more entirely different dyes. If such dyes are in our tables, their serial numbers are given; and if the particular dyes are not in the preceding tables there is given *in italics* in parentheses the 1914 Schultz number and name. Thus: "Helianthin, **36, 41** (*141, Azogelb 3G conc.*)," means that the name Helianthin is applied to Orange III, to Methyl orange and to Schultz No. 141, Azogelb 3G conc.

Acetin blue R..... **92**
Acid fuchsin, B, G, O, S..... **84**
Acid magenta, O..... **84**
Acid orange..... **60**
Acid yellow, cryst, D extra, DMP..... **25**
Acid yellow RS..... **65**
Acme yellow..... **65**
Alizarin, 1e..... **98**
Alizarin-Blaustich I and Ia..... **98**
Alizarin blue A, ABI, BM in Teig, C, DNW in Teig, F, G, GG, GW, R, RR, RR in Teig, WA in Teig, WC, WN in Teig, WR, WRR, WX, X, XA in Teig..... **95**
Alizarin blue S, SR, SRW, SW..... **99**
Alizarin blue soluble ABS..... **99**
Alizarin carmine..... **97**
Alizarin dark blue S, SW..... **99**
Alizarin green B..... **86**
Alizarin mono sulfonate..... **97**
Alizarin No. 1..... **98**
Alizarin No. 6..... **96**
Alizarin orange R, 2R-paste and powder.. **62**
Alizarin P..... **98**
Alizarin powder SA, W, W extra..... **97**
Alizarin purpurin..... **96**
Alizarin red IWS, S..... **97**
Alizarin sulfacid..... **97**
Alizarin yellow G, GG, GGW, 3G paste and powder..... **61**
Alizarin yellow R, RW paste and powder **62**
Alizarin V1..... **98**
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Alkali blue, B-5B, No. 2, No. 4, No. 6, R-5R, RR..... **82**
Alkanin..... **153**
Amidoazobenzol..... **30**
Anilin brown..... **78**
Anilin purple..... **90**
Anilin red..... **78**
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Benzoïn blue R..... **69**
Benzopurpurin B..... **67**
Benzyl violet, 7B..... **77**
Betanaphthol orange..... **60**
Bitter almond oil green..... **71**
Blau CB, spirit sol..... **92**
Bleu alcalin, 4B..... **82**
Bleu 3BS, C4B, de Lyon O..... **80**
Bleu méthyl..... **83**
Bleu neutre..... **94**
Bleu Nicholson 4B..... **82**
Bleu soluble pur..... **80**
Blue extra, water soluble for wool and silk..... **80**
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* The term curcumeïn has been applied to several different compounds.

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CHAPTER VI

Standard Buffer Solutions for Colorimetric Comparison

The standard solutions used in the colorimetric method of determining hydrogen ion concentrations are buffer solutions with such well defined compositions that they can be accurately reproduced, and with pH values accurately defined by hydrogen electrode measurements. They generally consist of mixtures of some acid and its alkali salt. Several such mixtures have been carefully studied. An excellent set has been described by Sørensen (1912). This set may be supplemented by the acetic acid—sodium acetate mixtures, most careful measurements of which have been made by Walpole (1914), and by Palitzsch's (1915) excellent boric acid-borax mixtures.

Clark and Lubs (1916) have designed a set of standards which they believe are somewhat more conveniently prepared than are the Sørensen standards. Their set is composed of the following mixtures:

Potassium chlorid + HCl
Acid potassium phthalate + HCl
Acid potassium phthalate + NaOH
Acid potassium phosphate + NaOH
Boric acid, KCl + NaOH

For a discussion of these mixtures, the methods used in determining their pH values, and the potential measurements we refer the reader to the original paper (*Journal of Biological Chemistry*, 1916, **25**, no. 3, p. 479). We may proceed at once to describe the details of preparation.

The various mixtures are made up from the following stock solutions: M/5 potassium chlorid (KCl), M/5 acid potassium phosphate (KH_2PO_4), M/5 acid potassium phthalate ($KHC_8H_4O_4$), M/5 boric acid with M/5 potassium chlorid (H_3BO_3, KCl), M/5 sodium hydroxid (NaOH), and M/5 hydrochloric acid (HCl). Although the subsequent mixtures are diluted to M/20 the above concentrations of the stock solutions are convenient for several reasons.

The water used in the crystallization of the salts and in the preparation of the stock solutions and mixtures should be redistilled. So-called "conductivity water," which is distilled first from acid chromate solution and again from barium hydroxid, is recommended, but it is not necessary.

M/5 potassium chlorid solution. (This solution will not be necessary except in the preparation of the most acid series of mixtures.) The salt should be recrystallized three or four times and dried in an oven at about 120°C. for two days. The fifth molecular solution contains 14.912 grams in 1 liter.

M/5 acid potassium phthalate solution. Acid potassium phthalate may be prepared by the method of Dodge (1915) modified as follows. Make up a concentrated potassium hydroxid solution by dissolving about 60 grams of a high-grade sample in about 400 cc. of water. To this add 50 grams of the commercial *resublimed* anhydrid of ortho phthalic acid. Test a cool portion of the solution with phenol phthalein. If the solution is still alkaline, add more phthalic anhydrid; if acid, add more KOH. When roughly adjusted to a slight pink with phenol phthalein[1] add as much more phthalic anhydrid as the solution contains and heat till all is dissolved. Filter while hot, and allow the crystallization to take place slowly. The crystals should be drained with suction and recrystallized at least twice from distilled water.[2]

Crystallization should not be allowed to take place below 20°C., for Dodge (1920) states:

> A saturated solution of the acid phthalate on chilling will deposit crystals of a more acid salt, having the formula $2KHC_8H_4O_4 \cdot C_8H_6O_4$. These crystals are in the form of prismatic needles, easily distinguished under the microscope from the 6-sided orthorhombic plates of the salt $KHC_8H_4O_4$.

Dry the salt at 110°–115°C. to constant weight.

A fifth molecular solution contains 40.836 grams of the salt in 1 liter of the solution.

M/5 acid potassium phosphate solution. A high-grade commercial sample of the salt is recrystallized at least three times

[1] Use a diluted portion for the final test.

[2] Samples of phthalic anhydrid which are now found on the market are frequently grossly impure. With some samples ten recrystallizations are necessary. Hence it is economy to purchase only the highest grades.

from distilled water and dried to constant weight at 110°–115°C. A fifth molecular solution should contain in 1 liter 27.232 grams. The solution should be distinctly red with methyl red and distinctly blue with brom phenol blue.

M/5 boric acid, M/5 potassium chlorid. Boric acid should be recrystallized several times from distilled water. It should be air dried[3] in thin layers between filter paper and the constancy of weight established by drying small samples in thin layers in a desiccator over $CaCl_2$. Purification of KCl has already been noted. It is added to the boric acid solution to bring the salt concentration in the borate mixtures to a point comparable with that of the phosphate mixtures so that colorimetric checks may be obtained with the two series where they overlap. One liter of the solution should contain 12.4048[4] grams of boric acid and 14.912 grams of potassium chlorid.

M/5 sodium hydroxid solution. This solution is the most difficult to prepare, since it should be as free as possible from carbonate. A solution of sufficient purity for the present purposes may be prepared from a high grade sample of the hydroxid in the following manner. Dissolve 100 grams NaOH in 100 cc. distilled water in a Jena or Pyrex glass Erlenmeyer flask. Cover the mouth of the flask with tin foil and allow the solution to stand over night till the carbonate has settled. Then prepare a filter as follows. Cut a "hardened" filter paper to fit a Buchner funnel. Treat it with warm, strong [1:1] NaOH solution. After a few minutes decant the sodium hydroxid and wash the paper first with absolute alcohol, then with dilute alcohol, and finally with large quantities of distilled water. Place the paper on the Buchner funnel and apply gentle suction until the greater part of the water has evaporated; but do not dry so that the paper curls. Now pour the concentrated alkali upon the middle of the paper, spread it with a glass rod making sure that the paper, under gentle suction, adheres well to the funnel, and draw the solution

[3] Boric acid begins to lose "water of constitution" above 50°C.

[4] This weight was used on the assumption that the atomic weight of boron is 11.0. The atomic weight has since been revised and appears as 10.9 in the 1920 table.

Because the solutions were standardized with the above weight of boric acid this weight should be used.

through with suction. The clear filtrate is now diluted quickly, after rough calculation, to a solution somewhat more concentrated than N/1. Withdraw 10 cc. of this dilution and standardize roughly with an acid solution of known strength, or with a sample of acid potassium phthalate. From this approximate standardization calculate the dilution required to furnish an M/5 solution. Make the required dilution with the least possible exposure, and pour the solution into a *paraffined*[5] bottle to which a calibrated 50 cc. burette and soda-lime guard tubes have been attached. The solution should now be most carefully standardized. One of the simplest methods of doing this, and one which should always be used in this instance, is the method of Dodge (1915) in which use is made of the acid potassium phthalate purified as already described. Weigh out accurately on a chemical balance with standardized weights several portions of the salt of about 1.6 grams each. Dissolve in about 20 cc. distilled water and add 4 drops phenol phthalein. Pass a stream of CO_2-free air through the solution and titrate with the alkali till a faint but distinct and permanent pink is developed. It is preferable to use a factor with the solution rather than attempt adjustment to an exact M/5 solution.

If one should be fortunate enough to find that the concentrated sodium hydroxid solution had clarified itself without leaving suspended carbonate, the clear solution might be carefully pipetted from the sediment. Cornog (1921) describes another method as follows:

> Distilled water contained in an Erlenmeyer flask is boiled to remove any carbon dioxide present, after which, when the water is cooled enough, ethyl ether is added to form a layer 3 or 4 cm. in depth. Pieces of metallic sodium, not exceeding about 1 cm. in diameter are then dropped into the flask. They will fall no further than the ether layer where they remain suspended. The water contained in the ether layer causes the slow formation of sodium hydroxid, which readily passes below to the water layer.

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[5] The author finds that thick coats of paraffine are more satisfactory than the thin coats sometimes recommended. Thoroughly clean and *dry* the bottle, warm it and then pour in the melted paraffine. Roll gently to make an even coat and just before solidification occurs stand the bottle upright to allow excess paraffine to drain to the bottom and there form a very substantial layer.

Cornog depends upon the evaporation of the ether as a barrier to CO_2. There are various ways in which the protection can be made more sure, and there are also various ways in which the aqueous solution may be separated from the ether.

From time to time there appear in the literature suggestions regarding the use of barium salts to remove the carbonate in alkali solutions.

In the author's opinion the next step to take, if the separation of carbonate from very concentrated NaOH solutions is not considered refined enough for the purpose at hand, is to proceed directly to the electrolytic preparation of an amalgam. Given a battery and two platinum electrodes this is a simple process. A *deep* layer of *redistilled* mercury is placed in a conical separatory funnel. The negative pole of the battery is led to this mercury by a glass-protected platinum wire. Over the mercury is placed a concentrated solution of recrystallized sodium chlorid and in this solution is dipped a platinum electrode connected with the positive pole of the battery. The battery may be 4 to 6 volts. Electrolysis is continued with occasional *gentle* shaking to break up amalgam crystals forming on the mercury surface.

Boil the CO_2 out of a litre or so of redistilled water, and, while steam is still escaping, stopper the flask with a cork carrying a siphon, a soda-lime guard tube and a corked opening for the separatory funnel.

When the water is cool introduce the delivery tube of the separatory funnel and deliver the amalgam. Allow reaction to take place till a portion of the solution, when siphoned off to a burette and standardized, shows that enough hydroxid has been formed. Then siphon approximately the required amount into a boiled-out and protected portion of water. Mix thoroughly and standardize.

M/5 hydrochloric acid solution. Dilute a high grade hydrochloric acid solution to about 20 per cent and distill. Dilute the distillate to approximately M/5 and standardize with the sodium hydroxid solution previously described. If convenient, it is well to standardize this solution carefully by the silver chlorid method and check with the standardized alkali.

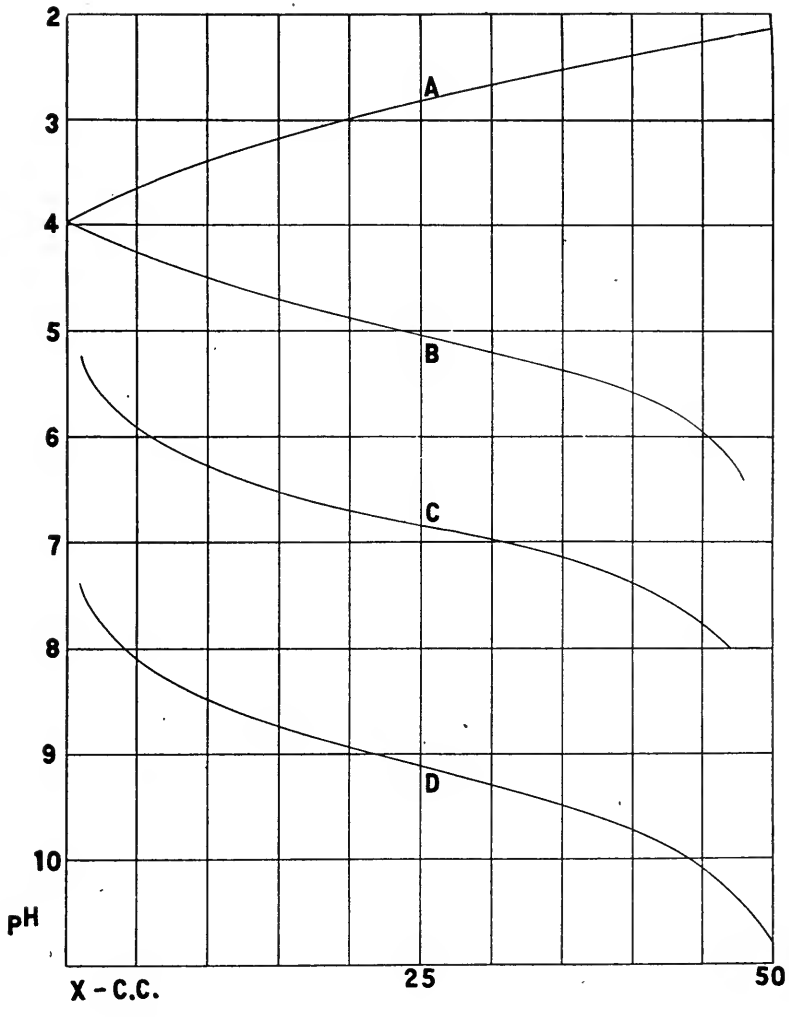


FIG. 11. CLARK AND LUBS' STANDARD MIXTURES

A. 50 cc. 0.2M KHPhthalate + X cc. 0.2M HCl. Diluted to 200 cc.
B. 50 cc. 0.2M KHPhthalate + X cc. 0.2M NaOH. Diluted to 200 cc.
C. 50 cc. 0.2M KH_2PO_4 + X cc. 0.2M NaOH. Diluted to 200 cc.
D. 50 cc. 0.2M H_3BO_3, 0.2M KCl + X cc. 0.2M NaOH. Diluted to 200 cc.

The only solution which it is absolutely necessary to protect from the CO_2 of the atmosphere is the sodium hydroxid solution. Therefore all but this solution may be stored in ordinary bottles of resistant glass. The salt solutions, if adjusted to exactly M/5, may be measured from clean calibrated pipettes.

These constitute the stock solutions from which the mixtures are prepared. The general relationships of these mixtures to their pH values are shown in figure 11. In this figure pH values are plotted as ordinates against X cc. of acid or alkali as abscissas. It will be found advantageous to plot this figure from table 21 with greatly enlarged scale so that it may be used as is Sørensen's chart (1909). The compositions of the mixtures at even intervals of 0.2 pH are given in table 21.

In any measurement the apportionment of scale divisions should accord with the precision. Scale divisions should not be so coarse that interpolations tax the judgment nor so fine as to be ridiculous. What scale divisions are best in the method under discussion it is difficult to decide, since the precision which may be attained depends somewhat upon the ability of the individual eye, and upon the material examined, as well as upon the means and the judgment used in overcoming certain difficulties which we shall mention later. Sørensen (1909) has arranged the standard solutions to differ by even parts of the components, a system which furnishes uneven increments in pH. Michaelis, (1910) on the other hand, makes his standards vary by about 0.3 pH so that the corresponding hydrogen ion concentrations are approximately doubled at each step. Certain general considerations lead to the conclusion that for most work estimation of pH values to the nearest 0.1 division is sufficiently precise, and that this precision can be obtained when the nature of the medium permits if the comparison standards differ by increments of 0.2 pH.

It is convenient to prepare 200 cc. of each of the mixtures and to preserve them in bottles each of which has its own 10 cc. pipette thrust through the stopper. It takes but little more time to prepare 200 cc. than it does to prepare a 10 cc. portion, and if the larger volume is prepared there will not only be a sufficient quantity for a day's work but there will be some on hand for the occasional test.

Unless electrometric measurements can be used as control, we

urge the most scrupulous care in the preparation and preservation of the standards. We have specified several recrystallizations of the salts used because no commercial samples which we have yet encountered are reliable.

TABLE 21

Composition of mixtures giving pH values at 20°C. at intervals of 0.2

KCl-HCl mixtures*

pH			
1.2	50 cc. M/5 KCl	64.5 cc. M/5 HCl	Dilute to 200 cc.
1.4	50 cc. M/5 KCl	41.5 cc. M/5 HCl	Dilute to 200 cc.
1.6	50 cc. M/5 KCl	26.3 cc. M/5 HCl	Dilute to 200 cc.
1.8	50 cc. M/5 KCl	16.6 cc. M/5 HCl	Dilute to 200 cc.
2.0	50 cc. M/5 KCl	10.6 cc. M/5 HCl	Dilute to 200 cc.
2.2	50 cc. M/5 KCl	6.7 cc. M/5 HCl	Dilute to 200 cc.

* The pH values of these mixtures are given by Clark and Lubs (1916) as *preliminary* measurements.

Phthalate-HCl mixtures

2.2	50 cc. M/5 KHPhthalate	46.70 cc. M/5 HCl	Dilute to 200 cc.
2.4	50 cc. M/5 KHPhthalate	39.60 cc. M/5 HCl	Dilute to 200 cc.
2.6	50 cc. M/5 KHPhthalate	32.95 cc. M/5 HCl	Dilute to 200 cc.
2.8	50 cc. M/5 KHPhthalate	26.42 cc. M/5 HCl	Dilute to 200 cc.
3.0	50 cc. M/5 KHPhthalate	20.32 cc. M/5 HCl	Dilute to 200 cc.
3.2	50 cc. M/5 KHPhthalate	14.70 cc. M/5 HCl	Dilute to 200 cc.
3.4	50 cc. M/5 KHPhthalate	9.90 cc. M/5 HCl	Dilute to 200 cc.
3.6	50 cc. M/5 KHPhthalate	5.97 cc. M/5 HCl	Dilute to 200 cc.
3.8	50 cc. M/5 KHPhthalate	2.63 cc. M/5 HCl	Dilute to 200 cc.

Phthalate-NaOH mixtures

4.0	50 cc. M/5 KHPhthalate	0.40 cc. M/5 NaOH	Dilute to 200 cc.
4.2	50 cc. M/5 KHPhthalate	3.70 cc. M/5 NaOH	Dilute to 200 cc.
4.4	50 cc. M/5 KHPhthalate	7.50 cc. M/5 NaOH	Dilute to 200 cc.
4.6	50 cc. M/5 KHPhthalate	12.15 cc. M/5 NaOH	Dilute to 200 cc.
4.8	50 cc. M/5 KHPhthalate	17.70 cc. M/5 NaOH	Dilute to 200 cc.
5.0	50 cc. M/5 KHPhthalate	23.85 cc. M/5 NaOH	Dilute to 200 cc.
5.2	50 cc. M/5 KHPhthalate	29.95 cc. M/5 NaOH	Dilute to 200 cc.
5.4	50 cc. M/5 KHPhthalate	35.45 cc. M/5 NaOH	Dilute to 200 cc.
5.6	50 cc. M/5 KHPhthalate	39.85 cc. M/5 NaOH	Dilute to 200 cc.
5.8	50 cc. M/5 KHPhthalate	43.00 cc. M/5 NaOH	Dilute to 200 cc.
6.0	50 cc. M/5 KHPhthalate	45.45 cc. M/5 NaOH	Dilute to 200 cc.
6.2	50 cc. M/5 KHPhthalate	47.00 cc. M/5 NaOH	Dilute to 200 cc.

KH_2PO_4-NaOH mixtures

5.8	50 cc. M/5 KH_2PO_4	3.72 cc. M/5 NaOH	Dilute to 200 cc.
6.0	50 cc. M/5 KH_2PO_4	5.70 cc. M/5 NaOH	Dilute to 200 cc.
6.2	50 cc. M/5 KH_2PO_4	8.60 cc. M/5 NaOH	Dilute to 200 cc.
6.4	50 cc. M/5 KH_2PO_4	12.60 cc. M/5 NaOH	Dilute to 200 cc.
6.6	50 cc. M/5 KH_2PO_4	17.80 cc. M/5 NaOH	Dilute to 200 cc.
6.8	50 cc. M/5 KH_2PO_4	23.65 cc. M/5 NaOH	Dilute to 200 cc.
7.0	50 cc. M/5 KH_2PO_4	29.63 cc. M/5 NaOH	Dilute to 200 cc.
7.2	50 cc. M/5 KH_2PO_4	35.00 cc. M/5 NaOH	Dilute to 200 cc.
7.4	50 cc. M/5 KH_2PO_4	39.50 cc. M/5 NaOH	Dilute to 200 cc.
7.6	50 cc. M/5 KH_2PO_4	42.80 cc. M/5 NaOH	Dilute to 200 cc.
7.8	50 cc. M/5 KH_2PO_4	45.20 cc. M/5 NaOH	Dilute to 200 cc.
8.0	50 cc. M/5 KH_2PO_4	46.80 cc. M/5 NaOH	Dilute to 200 cc.

Boric acid, KCl-NaOH mixtures

7.8	50 cc. M/5 H_3BO_3, M/5 KCl	2.61 cc. M/5 NaOH	Dilute to 200 cc.
8.0	50 cc. M/5 H_3BO_3, M/5 KCl	3.97 cc. M/5 NaOH	Dilute to 200 cc.
8.2	50 cc. M/5 H_3BO_3, M/5 KCl	5.90 cc. M/5 NaOH	Dilute to 200 cc.
8.4	50 cc. M/5 H_3BO_3, M/5 KCl	8.50 cc. M/5 NaOH	Dilute to 200 cc.
8.6	50 cc. M/5 H_3BO_3, M/5 KCl	12.00 cc. M/5 NaOH	Dilute to 200 cc.
8.8	50 cc. M/5 H_3BO_3, M/5 KCl	16.30 cc. M/5 NaOH	Dilute to 200 cc.
9.0	50 cc. M/5 H_3BO_3, M/5 KCl	21.30 cc. M/5 NaOH	Dilute to 200 cc.
9.2	50 cc. M/5 H_3BO_3, M/5 KCl	26.70 cc. M/5 NaOH	Dilute to 200 cc.
9.4	50 cc. M/5 H_3BO_3, M/5 KCl	32.00 cc. M/5 NaOH	Dilute to 200 cc.
9.6	50 cc. M/5 H_3BO_3, M/5 KCl	36.85 cc. M/5 NaOH	Dilute to 200 cc.
9.8	50 cc. M/5 H_3BO_3, M/5 KCl	40.80 cc. M/5 NaOH	Dilute to 200 cc.
10.0	50 cc. M/5 H_3BO_3, M/5 KCl	43.90 cc. M/5 NaOH	Dilute to 200 cc.

It is important to check the consistency of any particular set of these mixtures by comparing "5.8" and "6.2 phthalate" with "5.8" and "6.2 phosphate" using brom cresol purple. Also "7.8" and "8.0 phosphate" should be compared with the corresponding borates using cresol red.

Sørensen's standards are made as follows. The stock solutions are:

1. A carefully prepared exact tenth normal solution of HCl.
2. A carbonate-free exact tenth normal solution of NaOH.
3. A tenth molecular glycocoll solution containing sodium chlorid, 7.505 grams glycocoll and 5.85 grams NaCl in 1 litre of solution.
4. An M/15 solution of primary potassium phosphate which contains 9.078 grams KH_2PO_4 in 1 litre of solution.

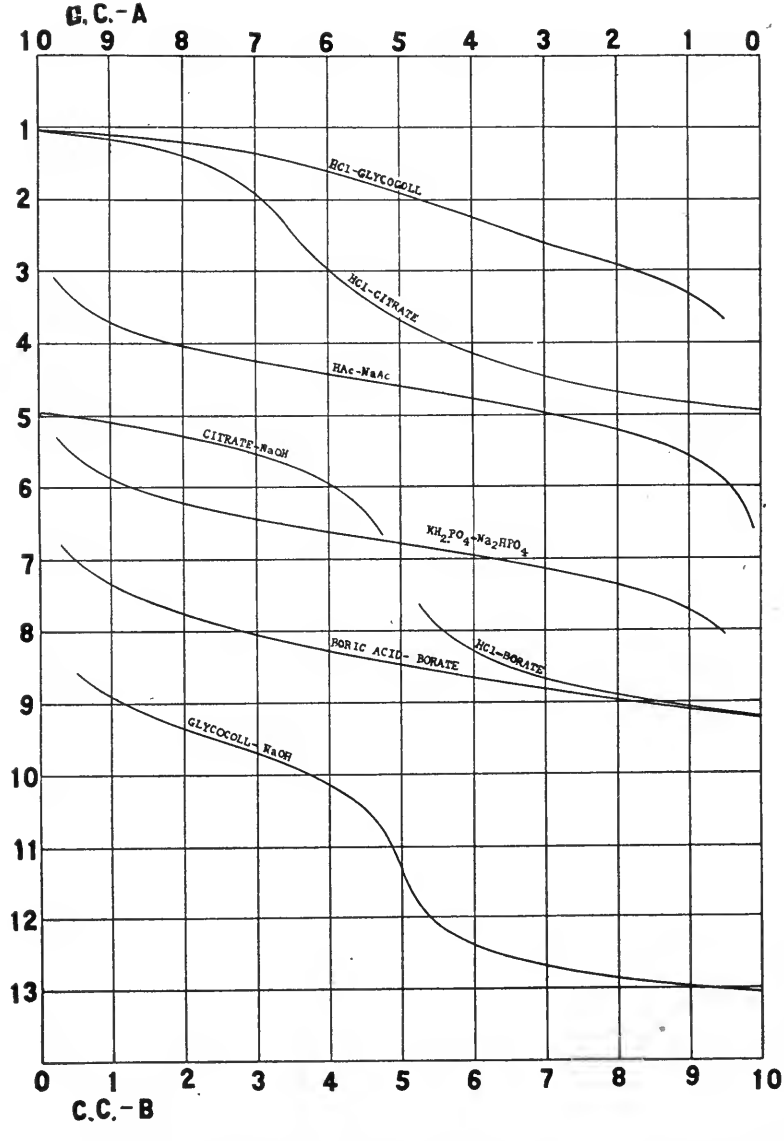


FIG. 12. SØRENSEN'S STANDARD MIXTURES, WALPOLE'S ACETATE SOLUTIONS AND PALITZSCH'S BORATE SOLUTIONS

Mixtures of A parts of acid constituent and B parts of basic constituent.

5. An M/15 solution of secondary sodium phosphate which contains 11.876 grams $Na_2HPO_4,2H_2O$ in 1 litre of solution.

6. A tenth molecular solution of secondary sodium citrate made from a solution containing 21.008 grams crystallized citric acid and 200 cc. carbonate-free N/1 NaOH diluted to 1 litre.

7. An alkaline borate solution made from 12.404 grams boric acid dissolved in 100 cc. carbonate-free N/1 NaOH and diluted to 1 litre.

The materials for these solutions are described by Sørensen as follows.

The water shall be boiled, carbon dioxid-free, distilled water, and the solutions shall be protected against contamination by CO_2.

Glycocoll (Glycine)

Two grams glycocoll should give a clear solution in 20 cc. water and should test practically free of chlorid or sulfate. Five grams should yield less than 2 mgm. of ash. Five grams should yield, on distillation with 300 cc. of 5 per cent sodium hydroxid, less than 1 mgm. of nitrogen as ammonia. The nitrogen content as determined by the Kjeldahl method should be 18.67 ± 0.1 per cent.

Primary phosphate, KH_2PO_4

The salt must dissolve clear in water and yield no test for chlorid or for sulfate. When dried under 20 or 30 mm. pressure for a day at 100°C. the loss in weight should be less than 0.1 per cent, and on ignition the loss should be 13.23 ± 0.1 per cent. When compared colorimetrically with citrate mixtures the stock phosphate solution should lie between "7" and "8 citrate-HCl." On addition of a drop of tenth-normal alkali or acid to 100 cc. the color of this phosphate solution with an indicator should be widely displaced.

Secondary phosphate, Na_2HPO_4, $2H_2O$

The salt with this content of water of crystallization is prepared by exposing to the ordinary atmosphere the crystals con-

taining twelve mols of water.[6] About two weeks exposure is generally sufficient. The salt should dissolve clear in water and yield no test for chlorid or sulfate. A day of drying under 20 to 30 mm. pressure at 100°C. and then careful ignition to constant weight, should result in a 25.28 ± 0.1 per cent loss. The stock solution should correspond on colorimetric test with "10 borate-HCl" and should be displaced beyond "8 borate-HCl" on addition of a drop of N/10 acid, and beyond "8 borate-NaOH" with a drop of alkali to 100 cc.

Citric acid, $C_6H_8O_7,H_2O$

The acid should dissolve clear in water, should yield no test for chlorid or sulfate and should give practically no ash. The water of crystallization may be determined by drying under 20 to 30 mm. pressure at 70°C. On drying in this manner the acid should remain colorless and lose 8.58 ± 0.1 per cent. The acidity of the citric acid solution is determined by titration with 0.2 N barium hydroxid with phenolphthalein as indicator. Titration is carried to a distinct red color of the indicator.

Boric acid, H_3BO_3

Twenty grams of boric acid should go completely into solution in 100 cc. of water when warmed on a strongly boiling water bath. This solution is cooled in ice water and the filtrate from the crystallized boric acid is tested as follows. It should give no tests for chlorides or sulfates. It should be orange to methyl orange. A drop of N/10 HCl added to 5 cc. should make the filtrate red to methyl orange. Twenty cubic centimeters of the filtrate evaporated in platinum, treated with about 10 grams of hydrofluoric acid and 5 cc. of concentrated sulfuric acid and reëvaporated, ignited and weighed, should yield less than 2 mgm. when corrected for non-volatile matter in the HF.

The following tables give the Sørensen mixtures with the corresponding pH values. Mixtures whose pH values are consid-

[6] Certain samples of secondary sodium phosphate sold for the preparation of buffer standards and called "Sørensen's Phosphate" are wrongly labeled Na_2HPO_4.

Temperature	10°	12°	14°	16°	18°	20°	22°	24°	26°	28°	30°	32°	34°	37°	40°
9.5 Glycine + 0.5 NaOH	8.75	8.70	8.66	8.62	8.58	8.53	8.49	8.45	8.40	8.37	8.32	8.28	8.24	8.18	8.12
9.0 Glycine + 1.0 NaOH	9.10	9.06	9.02	8.97	8.93	8.88	8.84	8.79	8.75	8.71	8.67	8.62	8.58	8.52	8.45
8.0 Glycine + 2.0 NaOH	9.54	9.50	9.45	9.40	9.36	9.31	9.26	9.22	9.17	9.13	9.08	9.04	9.00	8.92	8.85
7.0 Glycine + 3.0 NaOH	9.90	9.85	9.80	9.75	9.71	9.66	9.61	9.56	9.51	9.46	9.42	9.37	9.32	9.25	9.18
6.0 Glycine + 4.0 NaOH	10.34	10.29	10.24	10.18	10.14	10.09	10.03	9.98	9.93	9.88	9.83	9.78	9.73	9.66	9.58
5.5 Glycine + 4.5 NaOH	10.68	10.63	10.58	10.53	10.48	10.42	10.37	10.32	10.27	10.22	10.17	10.12	10.07	9.99	9.91
5.1 Glycine + 4.9 NaOH	11.29	11.24	11.18	11.12	11.07	11.01	10.96	10.90	10.85	10.79	10.74	10.68	10.62	10.54	10.46
5.0 Glycine + 5.0 NaOH	11.53	11.48	11.42	11.36	11.31	11.25	11.20	11.14	11.09	11.03	10.97	10.92	10.86	10.78	10.70
4.9 Glycine + 5.1 NaOH	11.80	11.74	11.68	11.62	11.57	11.51	11.45	11.39	11.33	11.27	11.22	11.16	11.10	11.02	10.93
4.5 Glycine + 5.5 NaOH	12.34	12.28	12.22	12.16	12.10	12.04	11.98	11.92	11.86	11.80	11.74	11.68	11.62	11.53	11.44
4.0 Glycine + 6.0 NaOH	12.65	12.59	12.52	12.46	12.40	12.33	12.27	12.21	12.15	12.09	12.03	11.96	11.90	11.81	11.72
3.0 Glycine + 7.0 NaOH	12.92	12.86	12.80	12.73	12.67	12.60	12.54	12.48	12.42	12.35	12.29	12.23	12.17	12.07	11.98
2.0 Glycine + 8.0 NaOH	13.12	13.06	12.99	12.92	12.86	12.79	12.73	12.66	12.60	12.53	12.47	12.41	12.34	12.25	12.15
1.0 Glycine + 9.0 NaOH	13.23	13.16	13.09	13.03	12.97	12.90	12.83	12.77	12.70	12.64	12.57	12.51	12.45	12.35	12.25

Temperature	42°	44°	46°	48°	50°	52°	54°	56°	58°	60°	62°	64°	66°	68°	70°
9.5 Glycine + 0.5 NaOH	8.07	8.03	7.99	7.95	7.91	7.86	7.82	7.78	7.74	7.69	7.65	7.61	7.56	7.52	7.48
9.0 Glycine + 1.0 NaOH	8.41	8.37	8.32	8.28	8.24	8.19	8.14	8.10	8.06	8.02	7.97	7.93	7.88	7.84	7.79
8.0 Glycine + 2.0 NaOH	8.81	8.76	8.72	8.67	8.63	8.58	8.53	8.49	8.44	8.40	8.35	8.30	8.26	8.21	8.16
7.0 Glycine + 3.0 NaOH	9.13	9.08	9.03	8.99	8.94	8.89	8.84	8.79	8.74	8.70	8.65	8.60	8.55	8.50	8.45
6.0 Glycine + 4.0 NaOH	9.53	9.48	9.43	9.38	9.33	9.28	9.23	9.18	9.13	9.08	9.03	8.98	8.93	8.88	8.82
5.5 Glycine + 4.5 NaOH	9.86	9.81	9.76	9.71	9.66	9.61	9.56	9.51	9.46	9.41	9.35	9.30	9.25	9.20	9.15
5.1 Glycine + 4.9 NaOH	10.40	10.35	10.29	10.24	10.18	10.13	10.07	10.02	9.96	9.90	9.85	9.79	9.74	9.68	9.62
5.0 Glycine + 5.0 NaOH	10.64	10.59	10.54	10.48	10.43	10.37	10.32	10.26	10.20	10.14	10.09	10.04	9.98	9.93	9.87
4.9 Glycine + 5.1 NaOH	10.87	10.81	10.75	10.69	10.64	10.58	10.52	10.46	10.40	10.35	10.29	10.23	10.17	10.11	10.05
4.5 Glycine + 5.5 NaOH	11.38	11.32	11.26	11.20	11.14	11.08	11.02	10.96	10.90	10.84	10.78	10.72	10.66	10.60	10.54
4.0 Glycine + 6.0 NaOH	11.65	11.59	11.53	11.47	11.41	11.34	11.28	11.22	11.16	11.10	11.03	10.97	10.91	10.84	10.78
3.0 Glycine + 7.0 NaOH	11.91	11.85	11.79	11.73	11.66	11.60	11.54	11.47	11.41	11.35	11.28	11.22	11.16	11.09	11.03
2.0 Glycine + 8.0 NaOH	12.08	12.02	11.96	11.89	11.83	11.77	11.70	11.64	11.57	11.51	11.44	11.38	11.31	11.25	11.18
1.0 Glycine + 9.0 NaOH	12.19	12.13	12.06	12.00	11.94	11.87	11.80	11.74	11.67	11.61	11.54	11.48	11.41	11.35	11.28

TABLE 23

Sørensen's borate—NaOH mixtures after Walbum

Temperature.....	10°	12°	14°	16°	18°	20°	22°	24°	26°	28°	30°	32°	34°	37°	40°
10 Borate.....	9.30		9.27		9.24		9.21		9.18		9.15		9.13	9.11	9.08
9 Borate + 1 NaOH	9.42		9.39		9.36		9.33		9.29		9.26		9.23	9.20	9.18
8 Borate + 2 NaOH	9.57		9.54		9.50		9.46		9.43		9.39		9.35	9.32	9.30
7 Borate + 3 NaOH	9.76		9.72		9.68		9.63		9.59		9.55		9.50	9.47	9.44
6 Borate + 4 NaOH	10.06	10.04	10.02	9.99	9.97	9.94	9.91	9.88	9.86	9.83	9.80	9.78	9.75	9.71	9.67
5 Borate + 5 NaOH	11.24	11.20	11.16	11.12	11.08	11.04	10.99	10.95	10.91	10.86	10.82	10.78	10.74	10.68	10.61
4 Borate + 6 NaOH	12.64	12.58	12.51	12.45	12.38	12.32	12.25	12.19	12.13	12.06	12.00	11.93	11.87	11.77	11.68

Temperature.....	42°	44°	46°	48°	50°	52°	54°	56°	58°	60°	62°	64°	66°	68°	70°
10 Borate.....		9.05		9.02		9.00		8.97		8.93		8.90			8.86
9 Borate + 1 NaOH		9.15		9.11		9.08		9.05		9.01		8.98			8.94
8 Borate + 2 NaOH		9.26		9.22		9.18		9.15		9.11		9.08			9.02
7 Borate + 3 NaOH		9.40		9.35		9.31		9.27		9.22		9.18			9.12
6 Borate + 4 NaOH	9.64	9.62	9.59	9.56	9.54	9.51	9.48	9.46	9.43	9.40	9.38	9.35	9.33	9.30	9.28
5 Borate + 5 NaOH	10.57	10.53	10.49	10.44	10.40	10.36	10.32	10.27	10.23	10.19	10.13	10.10	10.06	10.02	9.98
4 Borate + 6 NaOH	11.61	11.55	11.48	11.42	11.36	11.29	11.23	11.17	11.10	11.04	10.98	10.91	10.85	10.78	10.72

TABLE 24

Sørensen's borate—HCl mixtures after Walbum

Temperature	10°	20°	30°	40°	50°	60°	70°
10.0 Borate	9.30	9.23	9.15	9.08	9.00	8.93	8.86
9.5 Borate + 0.5 HCl	9.22	9.15	9.08	9.01	8.94	8.87	8.80
9.0 Borate + 1.0 HCl	9.14	9.07	9.01	8.94	8.87	8.80	8.74
8.5 Borate + 1.5 HCl	9.06	8.99	8.92	8.86	8.80	8.73	8.67
8.0 Borate + 2.0 HCl	8.96	8.89	8.83	8.77	8.71	8.65	8.59
7.5 Borate + 2.5 HCl	8.84	8.79	8.72	8.67	8.61	8.55	8.50
7.0 Borate + 3.0 HCl	8.72	8.67	8.61	8.56	8.50	8.45	8.40
6.5 Borate + 3.5 HCl	8.54	8.49	8.44	8.40	8.35	8.30	8.26
6.0 Borate + 4.0 HCl	8.32	8.27	8.23	8.19	8.15	8.11	8.08
5.75 Borate + 4.25 HCl	8.17	8.13	8.09	8.06	8.02	7.98	7.95
5.5 Borate + 4.5 HCl	7.96	7.93	7.89	7.86	7.82	7.79	7.76
5.25 Borate + 4.75 HCl	7.64	7.61	7.58	7.55	7.52	7.49	7.47

TABLE 25

Sørensen's citrate—NaOH mixtures after Walbum

Temperature	10°	20°	30°	40°	50°	60°	70°
10.0 Citrate	4.93	4.96	5.00	5.04	5.07	5.10	5.14
9.5 Citrate + 0.5 NaOH	4.99	5.02	5.06	5.10	5.13	5.16	5.20
9.0 Citrate + 1.0 NaOH	5.08	5.11	5.15	5.19	5.22	5.25	5.29
8.0 Citrate + 2.0 NaOH	5.27	5.31	5.35	5.39	5.42	5.45	5.49
7.0 Citrate + 3.0 NaOH	5.53	5.57	5.60	5.64	5.67	5.71	5.75
6.0 Citrate + 4.0 NaOH	5.94	5.98	6.01	6.04	6.08	6.12	6.15
5.5 Citrate + 4.5 NaOH	6.30	6.34	6.37	6.41	6.44	6.47	6.51
5.25 Citrate + 4.75 NaOH	6.65	6.69	6.72	6.76	6.79	6.83	6.86

TABLE 26

Sørensen's glycocoll—HCl mixtures

GLYCOCOLL	HCl	pH
cc.	*cc.*	
0.0	10.0	1.038
1.0	9.0	1.146
2.0	8.0	1.251
3.0	7.0	1.419
4.0	6.0	1.645
5.0	5.0	1.932
6.0	4.0	2.279
7.0	3.0	2.607
8.0	2.0	2.922
9.0	1.0	3.341
9.5	0.5	3.679

TABLE 27

Sørensen's phosphate mixtures

SECONDARY	PRIMARY	pH
cc.	*cc.*	
0.25	9.75	5.288
0.5	9.5	5.589
1.0	9.0	5.906
2.0	8.0	6.239
3.0	7.0	6.468
4.0	6.0	6.643
5.0	5.0	6.813
6.0	4.0	6.979
7.0	3.0	7.168
8.0	2.0	7.381
9.0	1.0	7.731
9.5	0.5	8.043

TABLE 28

Sørensen's citrate—HCl mixtures

CITRATE	HCl	pH
cc.	*cc.*	
0.0	10.0	1.038
1.0	9.0	1.173
2.0	8.0	1.418
3.0	7.0	1.925
3.33	6.67	2.274
4.0	6.0	2.972
4.5	5.5	3.364
4.75	5.25	3.529
5.0	5.0	3.692
5.5	4.5	3.948
6.0	4.0	4.158
7.0	3.0	4.447
8.0	2.0	4.652
9.0	1.0	4.830
9.5	0.5	4.887
10.0	0.0	4.958

TABLE 29

Walpole's acetate buffer mixtures, recalculated for intervals of 0.2 pH. Total acetate 0.2 molecular

pH	CONCENTRATION (MOLALITY)	
	Acetic Acid	Sodium acetate
3.6	0.185	0.015
3.8	0.176	0.024
4.0	0.164	0.036
4.2	0.147	0.053
4.4	0.126	0.074
4.6	0.102	0.098
4.8	0.080	0.120
5.0	0.059	0.141
5.2	0.042	0.158
5.4	0.029	0.171
5.6	0.019	0.181

TABLE 30

Palitzsch's borax-boric acid mixtures

M/20 BORAX	M/5 BORIC ACID, M/20 NaCl	pH
cc.	*cc.*	
10.0	0.0	9.24
9.0	1.0	9.11
8.0	2.0	8.98
7.0	3.0	8.84
6.0	4.0	8.69
5.5	4.5	8.60
5.0	5.0	8.51
4.5	5.5	8.41
4.0	6.0	8.31
3.5	6.5	8.20
3.0	7.0	8.08
2.5	7.5	7.94
2.3	7.7	7.88
2.0	8.0	7.78
1.5	8.5	7.60
1.0	9.0	7.36
0.6	9.4	7.09
0.3	9.7	6.77

ered by Sørensen to be too uncertain and which he has indicated by brackets are omitted in these tables. The third decimal of Sørensen's tables are given by Sørensen in small type.

TABLE 31

McIlvaine's standards

pH required	0.2 M Na_2HPO_4	0.1 M citric acid
	cc.	cc.
2.2	0.40	19.60
2.4	1.24	18.76
2.6	2.18	17.82
2.8	3.17	16.83
3.0	4.11	15.89
3.2	4.94	15.06
3.4	5.70	14.30
3.6	6.44	13.56
3.8	7.10	12.90
4.0	7.71	12.29
4.2	8.28	11.72
4.4	8.82	11.18
4.6	9.35	10.65
4.8	9.86	10.14
5.0	10.30	9.70
5.2	10.72	9.28
5.4	11.15	8.85
5.6	11.60	8.40
5.8	12.09	7.91
6.0	12.63	7.37
6.2	13.22	6.78
6.4	13.85	6.15
6.6	14.55	5.45
6.8	15.45	4.55
7.0	16.47	3.53
7.2	17.39	2.61
7.4	18.17	1.83
7.6	18.73	1.27
7.8	19.15	0.85
8.0	19.45	0.55

Walbum (1920) has determined the pH values for the Sørensen mixtures at temperatures of 10°, 18°, 28°, 37°, 46°, 62° and 70°C. and has interpolated data for intervening temperatures. He uses a system of reference essentially that which is described

in Chapter XIX as standard. He finds that upon this basis the alteration of pH with temperature is for the most part negligible for the phosphate mixtures, the glycocoll-HCl mixtures and the citrate-HCl mixtures. Data for the other mixtures are tabulated in tables 22, 23, 24 and 25. In these will be found Sørensen's values at 18°. Tables 26, 27 and 28 are taken from Sørensen's paper of 1912.

The stock solutions for the Palitzsch mixtures given in table 30 are an M/20 Borax solution containing 19.108 grams[7] $Na_2B_4O_7$ $10H_2O$ in 1 litre; and an M/5 Boric acid, NaCl solution containing 12.404 grams[7] H_3BO_3 and 2.925 grams NaCl in 1 litre.

McIlvaine (1921) has given the electrometrically determined pH values for a series of mixtures of 0.2 M disodium phosphate and 0.1 M citric acid. Since the citrate exerts a buffer action at the steep part of the phosphate curve near the position where the mono alkali phosphate alone is present McIlvaine's mixtures give a continuous buffer action from pH 2.2 to pH 8.0. His data are shown in table 31.

Acree and his coworkers have published curves for other mixtures giving more or less smooth slopes over wide ranges of pH.

Kolthoff in his 1921 text has recalculated the following data from Ringer (1909):

TABLE 32

Ringer's mixtures of 0.15M Na_2HPO_4 and 0.1M NaOH

MIXTURE	pH
50 cc. Na_2HPO_4 + 15 cc. NaOH	10.97
50 cc. Na_2HPO_4 + 25 cc. NaOH	11.29
50 cc. Na_2HPO_4 + 50 cc. NaOH	11.77
50 cc. Na_2HPO_4 + 75 cc. NaOH	12.06

[7] The values given by Palitzsch were calculated upon the basis of 11.0 as the atomic weight of boron. Since this was the value used, the new value of 10.9 given in the atomic weight table in the report of the international committee for 1922 should not be used in calculating the composition of the specific solutions given by Palitzsch.

CHAPTER VII

Sources of Error in Colorimetric Determinations

There are errors of technique such as incorrect apportionment of the indicator concentration in tested and standard solution and the use of unequal depths of solutions through which the colors are viewed that may be passed over with only a word of reminder. Likewise we may recall certain of the optical effects mentioned in Chapter IV and then pass on to the more serious difficulties in the application of the indicator method.

In the comparison of electrometric and colorimetric measurements discrepancies have often been traced so clearly to two definite sources of error that they have been given categorical distinction. They are the so-called "protein" and "salt" errors.

From what has already been said in previous pages, it will be seen that, if there are present in a tested solution bodies which remove the indicator or its ions from the field of action either by adsorption or otherwise, the equilibria which have formed the basis of our treatment will be disturbed. An indicator in such a solution may show a color intensity, or even a quality of color, which is different from that of the same concentration of the indicator in a solution of the same hydrogen ion concentration where no such disturbance occurs. We could easily be led to attribute very different hydrogen ion concentrations to the two solutions. This situation is not uncommon when we are dealing with protein solutions, for in some instances there is distinctly evident the removal of the indicator from the field. In other cases the discrepancy between electrometric and colorimetric measurements is not so clear, nor can it always be attributed solely to the indicator measurement.

If two solutions of inorganic material, each containing the same concentration of hydrogen ions, are tested with an indicator, we should expect the same color to appear. If, however, these two solutions have different concentrations of salt, it may happen that the indicator color is not the same. As Sørensen (1909) and Sørensen and Palitzsch (1913) have demonstrated, this effect of

the salt content of a solution cannot be tested by adding the salt to one of two solutions which have previously been brought to the same hydrogen ion concentration. The added salt, no matter if it be a perfectly neutral salt, will change either the hydrogen ion concentration or the hydrogen ion activity of the solution or so affect the electrode equilibrium that it appears as if the hydrogen ion activity is altered.

So long as hydrogen electrode measurements are made the standard it is *convenient* to throw the burden of the "salt effect" upon the indicator; but neutral salts are known to displace electrode potential differences from the point estimated from the expected hydrogen ion *concentration*. Tentatively we may deal with the salt effect as if the hydrogen electrode measurement were the point of reference, and this will doubtless harmonize with future developments of theory.

Bjerrum (1914) gives an example of a case where the influence of the neutral salt is evidently upon the buffer equilibrium rather than on the indicator. An ammonium-ammonium salt buffer mixture and a borate buffer mixture are both made up to the same color of phenolphthalein. On the addition of sodium chloride the color of phenolphthalein becomes stronger in the ammonium mixture and weaker in the borate mixture.

The following table taken from Prideaux (1917) illustrates the order of magnitude of the "salt error" in some instances.

INDICATOR	BUFFER USED	CHANGE OF pH IN PRESENCE OF 0.5 N NaCl
Para benzene sulphonic acid azo naphthylamine.	Phosphate	−0.10
Para nitro phenol	Phosphate	+0.15
Alizarine sulphonic acid	Phosphate	+0.26
Neutral red	Phosphate	−0.09
Rosolic acid	Phosphate	+0.06
Para benzene sulphonic acid azo α-naphthol	Phosphate	+0.12
Phenolphthalein	Phosphate	+0.12

In cases where the solutions under examination are of the same general nature hydrogen electrode measurements may be taken as the standard and colorimetric measurements calibrated accordingly. Sørensen and Palitzsch (1910) did this in their study of

the salt errors of indicators in sea water. They acidified the sea water and passed hydrogen through to displace carbon dioxid, and then neutralized it to the ranges of various indicators with buffer mixtures and compared colorimetric with electrometric measurements. In this way they found the following deviations.

INDICATOR	BUFFER	PARTS PER 1000 OF SALTS AND CORRESPONDING ERRORS			
		35	20	5	1
Paranitro phenol.....	Phosphate	+0.12	+0.08		
Neutral red.....	Phosphate	−0.10	−0.05	0	0
α-Napththol phthalein..	Borate	+0.22	+0.17	+0.03	−0.07
	Phosphate	+0.16	+0.11	−0.04	−0.14
Phenolphthalein.....	Borate	+0.21	+0.16	+0.05	−0.03

If, for example, sea water of about 3.5 per cent salt is matched against a standard borate solution with phenolphthalein and appears to be pH 8.43 the real value is pH 8.22.

Such calibration is doubtless the very best way to deal with the salt errors since it tends to bring measurements to a common experimental system of reference.

Kolthoff (1922) gives the following table showing the corrections to be applied for the "salt error" of various indicators. It should be noted that Kolthoff includes in this table data obtained when the hydrogen electrode potentials were taken as standard and also data in which the pH values were calculated. The two sets are not strictly comparable (see Chapter XIX) and therefore must be used with caution in theoretical work. We have eliminated from Kolthoff's table Congo red, Azolitmin, and Tropaeolin O (Chrysoïn) which Kolthoff describes as having salt errors so large that these indicators are useless.

Michaelis and his coworkers have determined the salt errors for a number of the nitrophenols, but, since the corrections are often intimately related to the constants used in Michaelis' method of operating, the reader is referred to the original literature for the details. See Chapter VIII.

The reader was warned in Chapter II that the treatment to be given the so-called salt errors of indicators would not deal with the theory. There are various theories that have been advanced,

TABLE 33

Salt error of indicators, after Kolthoff

INDICATOR	SALT	SALT CONCENTRATION	CORRECTION	REMARKS
Tropaeolin OO (Orange IV)	KCl	0.10 N	−0.05	Indicator suitable. NaCl has about same influence
	KCl	0.25 N	−0.01	
	KCl	0.50 N	+0.06	
	KCl	1.00 N	+0.23	
Methyl orange	KCl	0.10 N	−0.08	Indicator suitable. NaCl has about same influence
	KCl	0.25 N	−0.08	
	KCl	0.50 N	+0.02	
	KCl	1.00 N	+0.23	
Butter yellow.....	KCl	0.10 N	−0.08	Same errors as methyl orange but indicator flocculates with salt
Thymol blue (acid range)	KCl	0.10 N	−0.06	NaCl has same influence
	KCl	0.20 N	−0.06	
	KCl	0.50 N	−0.04	
	KCl	1.00 N	+0.05	
Brom phenol blue	KCl	0.10 N	−0.05	Corrections large at weaker concentration of salt
	KCl	0.25 N	−0.15	
	KCl	0.50 N	−0.35	
	KCl	1.00 N	−0.35	
Brom cresol purple..	NaCl	0.50 N	−0.25	
Phenol red.....	NaCl	0.50 N	−0.15	At small concentrations of salt correction of opposite sign
Thymol blue.....	NaCl	0.50 N	−0.17	
Methyl red.....	NaCl	0.50 N	+0.10	
p-Nitro phenol.....	NaCl	0.50 N	−0.05	
Azo yellow 3G.....	NaCl	0.50 N	0.00	
Phenolphthalein.....	NaCl	0.50 N	−0.17	
Nitramine (?).....	KCl	0.10 N	−0.06	NaCl has about same influence
	KCl	0.25 N	−0.12	
	KCl	0.50 N	−0.10	
	KCl	1.00 N	−0.29	

but up to a recent time none has been entirely satisfactory. Whether the newer concepts of the conduct of strong electrolytes will furnish a basis for the correlation of experimental data remains to be seen. This much at least will be demanded, that the habit of indiscriminately jumbling together dissociation constants and other data obtained by widely different methods and bearing different implications shall cease. Until a thoroughly consistent method of approach and of calculation is accomplished and its value established, the only safe procedure to follow is to calibrate salt errors by experimental hydrogen electrode measurements.

In dealing with protein solutions calibration is less certain. When solutions to be tested vary greatly, not only in protein content but also in the composition and concentration of their salt content, systematic calibration becomes very difficult. When there are added the difficulties presented by strong coloration and turbidity, calibration is impossible. Such is the situation to be faced when dealing with the media and the cultures which the bacteriologist must handle.

It is sometimes helpful to construct titration curves of a solution under examination, making measurements after addition of graded quantities of acid and alkali, in one case with the hydrogen electrode and in the other case with indicators, preferably indicators of different types. The indicator readings may then reveal breaks not to be expected from the hydrogen ion relations of the solution. If, however, no comparison is made with hydrogen electrode measurements, the observer must rely to a considerable extent upon his judgment. "Protein errors" are generally the larger the more complex and concentrated the protein and tend to decrease with increase in the extent of protein hydrolysis.

There seems to be no way then to deal with either the protein or the salt error of indicators but to rely upon the use of those indicators which give *relatively* small errors, to keep in mind the order of magnitude of the error to be expected from the general nature of the solution tested, and, in important cases, to standardize to the electrometric basis as an arbitrary provisional standard.

Because of the great variety of solutions tested by the colorimetric method it is impracticable to give a condensed statement of the probable errors. Elaborate tables of colorimetric and electrometric comparisons are given by Sørensen (1909) for the

cases he studied. Clark and Lubs (1917) have tabulated their results with the sulphonphthalein indicators.

In the work of Michaelis or that of Kolthoff salt corrections are for the most part established by means of hydrogen electrode measurements. Wells (1920) has tabulated some data for cresol red in a manner useful for a certain type of water study (cf. McClendon 1917), and Brightman, Meacham and Acree (1920) have recorded the effects of different concentrations of phosphate buffer.

The "protein error" and the "salt error" have been given prominence in the literature partly because both have to be taken into consideration in dealing with biological solutions, and partly because there is to be perceived underlying the salt error a most interesting phenomenon of rather general interest. However, this emphasis should not obscure the fact that there are specific conditions for each indicator which render that indicator useless for the determination of pH. For instance alizarine, in passing from the phosphate to the borate buffer mixtures exhibits a sudden transition which has all the appearances of a specific effect of the borate upon the indicator. And alizarine is not alone in this peculiarity. This same alizarine in the presence of aluminium may form a lake and with proper pH control may be made a useful reagent for aluminium in place of a very poor acid-base indicator. Zoller (1921) has called attention to the incompatibility between certain dyes and the phthalate buffers. Many indicators are easily reduced or like methyl red easily reduced and then so altered that the reduction is irreversible. A number of indicators undergo their color changes slowly or else fade and are lost to the observer. Other indicators precipitate with certain cations, for instance Orange IV and Congo with alkali earths. In short all possibilities must be watched lest the investigator, venturing upon the study of some new solution, be misled by the mark of reliability placed upon an indicator tried under limited circumstances.

Wherever possible it is good practice to test doubtful cases with two indicators of widely different chemical composition.

As to the effect of temperature variation, comparatively little work has been done. Gillespie and others have some notes on the subject and more recently Michaelis and his coworkers have

included temperature data in stating the constants used in the Michaelis and Gyemant method (see Chapter VIII). Kolthoff (1921) has extended the theory of School in which account is taken of the acidic or basic nature of an indicator, but there often remains some question as to how a given indicator is to be classified. Kolthoff using the values of Kohlrausch and Heydweiller for the dissociation constant of water at various temperatures has reduced his observations to the following table. In this

TABLE 34

Displacement of indicator exponent between 18°C. and 70°C. after Kolthoff

INDICATOR	pH DISPLACEMENT	pOH DISPLACEMENT	RATIO OF DISSOCIATION CONSTANT AT 70°C. TO THAT AT ORDINARY TEMPERATURE
Nitramine	−1.45	0.0	1.0
Phenol phthalein	−0.9 to 0.4	−0.55 to 1.05	About 5
Thymol blue, alkaline range	−0.4	−1.05	2.5
α-naphthol phthalein	−0.4	−1.05	2.5
Curcumine	−0.4	−1.05	2.5
Phenol red	−0.3	−1.15	2.0
Neutral red	−0.7	−0.75	
Brom cresol purple	0.0	−1.45	1.0
Azolitmin	0.0	−1.45	1.0
Methyl red	−0.2	−1.25	
Lacmoid	−0.4	−1.05	2.5
p-nitro phenol	−0.5	−0.95	3.2
Methyl orange	−0.3	−1.15	14.0
Butter yellow	−0.18	−1.17	15.0
Bromphenol blue	0.0	−1.45	1.0
Tropaeolin OO	−0.45	−1.0	10.0
Thymol blue, acid range	0.0	−1.45	1.0

table the displacement of −0.4 for thymol blue means that if thymol blue in a solution at 70°C. shows the same color as the same concentration of this indicator in a buffer of pH 9.4 at ordinary temperature then the pH of the solution at 70°C. is 9.0. Corrections for temperatures between room temperature and 70°C. may be interpolated from the data in the table.

REFERENCES

Abegg-Bose (1899), Arrhenius (1899), Bjerrum (1914), Brightman-Meacham Acree (1920), Chow (1920), Clark-Lubs (1917), Dawson-Powis (1913), Gillespie-Wise (1918), Harned (1915), Kolthoff (1916, 1918, 1922), Lewis (1912), McBain-Coleman (1914), McBain-Salmon (1920), Michaelis (1920–21), Michaelis-Gyemant (1920), Michaelis-Krüger (1921), Michaelis-Rand (1909), Palmaer-Melander (1915), Poma (1914), Poma-Patroni (1914), Prideaux (1917), Rosenstein (1912), Sackur (1901), Sørensen (1909), Sørensen-Palitzsch (1910), (1913), Wells (1920), Zoller (1921).

See also Chapter II and page 341.

CHAPTER VIII

Approximate Determinations with Indicators

If you can measure that of which you speak, and can express it by a number, you know something of your subject; but if you cannot measure it, your knowledge is meagre and unsatisfactory.—Lord Kelvin

The distinctive advantages of the indicator method are the ease and the rapidity with which the approximate hydrogen ion concentration of a solution may be measured. The introduction of improved indicators, the charting of their pH ranges, better definition of degree in "acidity" or "alkalinity," and the illumination of the theory of acid-base equilibria have developed among scientific men in general an appreciation of how indefinite were those old, favorite terms—"slightly acid," "distinctly alkaline," and "neutral." There is now a clear recognition of the distinct difference between quantity and intensity of acidity; and for each aspect there may be given *numerical values* admitting no misunderstanding.

Furthermore the clarification of the subject has brought a perspective which may show where accuracy is unnecessary and where fair approximation is desirable. In such a case the investigator turns to the indicator method knowing that even if his results are rough they can still be expressed in numerical values having a definite meaning to others.

Now a very good approximation may be attained by color memory and without the aid of the standard buffer solutions or of the systems presently to be described in which the standard buffer solutions are dispensed with.

To establish a color memory as well as to refresh it a set of "permanent" standards is convenient. These may be prepared with the standard buffer solutions in the ordinary way, protected against mold growth by means of a drop of toluol, and sealed by drawing off the test tubes in a flame or by corking with the cork protected by tinfoil or paraffine. For exhibition purposes long homeopathic vials make a very good and uniform container. They may be filled almost to the brim and a cork inserted, if a

slit is made for the escape of excess air and liquid. The slit may then be sealed with paraffine. A hook of spring-brass snapped about the neck makes a support by which the vial may be fastened to an exhibition board. A neater container is the so-called typhoid-vaccine ampoule which is easily sealed in the flame.

If one of a series of standards so prepared should alter, the change can generally be detected by the solution falling out of the proper slope of color gradation. But if all in a series should change, it may be necessary to compare the old with new standards. Because such changes do occur, "permanent" standards are to be used with caution. The sulfon phthalein indicators make fairly permanent standards but the methyl red which is an important member of the series of indicators recommended by Clark and Lubs (1917) often deteriorates within a short time.

A device which furnishes a color standard to be interpreted by means of a dissociation curve is the color wedge of Bjerrum (1914). This is a long rectangular box with glass sides and a diagonal glass partition which divides the interior into two equal wedges. One compartment contains a solution of the indicator fully transformed into its alkaline form, the other a like concentration of the indicator transformed to the acid form. A view through these wedges should imitate the view of a like depth and concentration of the indicator transformed to that degree which is represented by the ratio of wedge thicknesses at the point under observation. Compare Barnett and Barnett (1920) and Myers (1922).

As an aid to memory the dissociation curves of the indicators are helpful even when used alone. The color chart shown in Chapter III is a still better aid to memory and within the limitations mentioned the colors may be used as rough standards.

Sondén (1921) has used colored glasses and Kolthoff (1922) inorganic salt solutions as color standards.

Colorimetric determination of hydrogen ion concentration without the use of standard buffer solutions

We have already seen that if an indicator is an acid, its degree of dissociation, α, is related to the hydrogen ion concentration of the solution by the equation

$$[H^+] = K_a \frac{1 - \alpha}{\alpha}$$

We have also seen that if Ka, the true dissociation constant is replaced by the so-called apparent dissociation constant, K_{IA}, which is a function of Ka and of the constants of tautomeric equilibria, then α represents the degree of color transformation. We then have

$$[H^+] = K_{IA} \frac{1 - \alpha}{\alpha}$$

or the more convenient form

$$pH = \log \frac{1}{K_{IA}} + \log \frac{\alpha}{1 - \alpha}$$

where α may now be considered as to the degree of color transformation. If, for instance, an indicator conducts itself as a simple acid with apparent dissociation constant 1×10^{-6}, we can construct the dissociation curve with its central point of inflection at pH 6, and then, assuming that this curve represents the relation of the percentage color transformation to pH, we can determine the pH of a solution if we can determine the percentage color transformation which this indicator displays in a tested solution. Proceeding on these simple and often unjustified assumptions we can now devise a very simple way of detecting the percentage color transformation. The following is quoted from Gillespie (1920):

> We may assume that light is absorbed independently by the two forms of the indicator, and hence that the absorption, and in consequence of this the residual color emerging, will be the same whether the two forms are present together in the same solution or whether the forms are separated for convenience in two different vessels and the light passes through one vessel after the other. Therefore, if we know what these percentages are for a given indicator in a given buffer mixture, we can imitate the color shown in the buffer mixture by dividing the indicator in the proper proportion between two vessels, and putting part of it into the acid form with excess of acid, the rest into the alkaline form with excess of alkali.

Gillespie sets up in the comparator (see page 70) two tubes, one of which contains, for example, three drops of a given indicator fully transformed into the acid color, and the other of which contains seven drops of the indicator fully transformed into the alkaline form. The drop ratio 3:7 should correspond to the ratio of the concentrations of acid and alkaline forms of ten drops of the

indicator in a solution of that pH which is shown by the dissociation curve of the indicator to induce a seventy per cent transformation. If then the two comparison tubes and the tested solution are kept at the same volume, and the view is through equal depths of each, a matching of colors should occur between the virage of the two superposed comparison tubes and that of the tested solution.

Barnett and Chapman (1918) applied this method with one indicator phenol red but without using the dissociation curve. Gillespie (1920) extended the procedure to several other indicators and made use of the dissociation curves so that he was able to smooth out to more probable values the experimental data relating drop ratios to pH. This is important because the experimental error in judging color is not inconsiderable and if the purely empirical data be made the sole basic standardization of the method there may be involved a systematic error, which, added to the error of the individual measurement may make the cumulative error unnecessarily large. That this had already occurred was indicated by Gillespie's comparison of the values for the drop ratios of phenol red given by Barnett and Chapman on the one hand and the report of the bacteriologists' committee (Conn, *et al.*, 1919) on the other hand.

Gillespie found the correspondence between the experimental and the theoretical results predicted on the basis of the simplifying assumptions mentioned above to be very good for the sulfon phthaleins, doubtless because of the reasons mentioned in Chapter IV. He also showed good correspondence in the case of methyl red but reiterates the fact that phenol phthalein cannot be treated by means of the simple dissociation curve for a mono acidic acid, as was mentioned in Chapter IV.

In table 35 are given the pH values corresponding to various drop ratios of seven indicators as determined by Gillespie. At the bottom of the table are shown the quantities of acid used to obtain the acid color in each case. The use of acid phosphate instead of hydrochloric acid in two cases is because the stronger acid might transform the indicator into that red form which occurs with all the sulfon phthalein indicators at very high acidities. The 0.05 *M* HCl is prepared with sufficient accuracy by diluting 1 cc. concentrated hydrochloric acid (specific gravity 1.19) to 240 cc.

The alkaline form of the indicator is obtained in each case with a drop of alkali (two drops in the case of thymol blue). The alkali solution used for this purpose may be prepared with sufficient accuracy by making up a 0.2 per cent solution with ordinary stick NaOH. The indicator solutions may be made up as described on page 81. Gillespie prefers the alcoholic solution in the case of methyl red and specifies it for soil work.

TABLE 35

Gillespie's table of pH values corresponding to various drop-ratios

DROP-RATIO	BROM-PHENOL BLUE	METHYL RED	BROM-CRESOL PURPLE	BROM-THYMOL BLUE	PHENOL RED	CRESOL RED	THYMOL BLUE
1:9	3.1	4.05	5.3	6.15	6.75	7.15	7.85
1.5:8.5	3.3	4.25	5.5	6.35	6.95	7.35	8.05
2:8	3.5	4.4	5.7	6.5	7.1	7.5	8.2
3:7	3.7	4.6	5.9	6.7	7.3	7.7	8.4
4:6	3.9	4.8	6.1	6.9	7.5	7.9	8.6
5:5	4.1	5.0	6.3	7.1	7.7	8.1	8.8
6:4	4.3	5.2	6.5	7.3	7.9	8.3	9.0
7:3	4.5	5.4	6.7	7.5	8.1	8.5	9.2
8:2	4.7	5.6	6.9	7.7	8.3	8.7	9.4
8.5:1.5	4.8	5.75	7.0	7.85	8.45	8.85	9.55
9:1	5.0	5.95	7.2	8.05	8.65	9.05	9.75
Produce acid color with	1 cc. of 0.05M HCl	1 drop of 0.05M HCl	1 drop of 0.05M HCl	1 drop of 0.05M HCl	1 drop of 0.05M HCl	1 drop of 2 per cent H_2KPO_4	1 drop of 2 per cent H_2KPO_4

Gillespie proceeds as follows:

Test tubes 1.5 cm. external diameter and 15 cm. long are suitable for the comparator and for the strengths given for the indicator solutions. It is advisable to select from a stock of tubes a sufficient number of uniform tubes by running into each 10 cc. water and retaining those which are filled nearly to the same height. A variation of 3 to 4 mm. in a height of 8 cm. is permissible. Test tubes without flanges are preferable. The tubes may be held together in pairs by means of one rubber band per pair, which is wound about the tubes in the form of two figure 8's.

It is convenient to use metal test tube racks, one for each indicator, each rack holding two rows of tubes, accommodating one tube of each pair in front and one in back. For any desired indicator a set of color standards is prepared by placing from 1 to 9 drops of the indicator solution in the 9 front tubes of the pairs and from 9 to 1 drops in the back row of tubes. A

drop of alkali is then added to each of the tubes in the front row (2 drops in the case of thymol blue), sufficient to develop the full alkaline color and a quantity of acid is added to each of the tubes in the back row to develop the full acid color without causing a secondary change of color (see table 35 for quantities). . . . . The volume is at once made up in all the tubes to a constant height (within about one drop) with distilled water, the height corresponding to 5 cc.

These pairs are used in the comparator and matched with the tested solution. This tested solution is added to ten drops of the proper indicator until a volume of 5 cc. is attained and the tube is then placed in the comparator backed by a water blank.

Gillespie describes the use of the comparator (page 70) and a modification for the accommodation of sets of three tubes used when colored solutions have to be compared. He also discusses a number of minor points and cautions against the indiscriminate comparison of measurements taken at different temperatures. For the details the original papers should be consulted. Were it not that the writer has seen evidence that the method has been applied with neglect of volume or concentration relations called for by the theory involved and carefully specified by Gillespie, it would seem unnecessary to advise that the principles be understood before the method is used. Certain other misconceptions of theory and practice found in a treatment of the method by Medalia (1920) have been corrected by Gillespie (1921).

A very judicious appraisal of the method's value was given by Gillespie in these words:

> The method should be of especial use in orienting experiments, or in occasional experiments involving hydrogen ion exponent measurements, either where it is unnecessary to push to the highest degree of precision obtainable, or where the investigator may be content to carry out his measurements to his limit of precision and to record his results in such a form that they may be more closely interpreted when a more precise study of indicators shall have been completed.

For the elaboration of certain manipulative details see Van Alstine (1920).

If an indicator has only one color, for instance if it is yellow in the alkaline form and colorless in the acid form, it is evident that the method employed by Gillespie may be used with the elimination of one of the sets of tubes. Thus if 10 cc. of a tested

solution containing 1 cc. of para nitro phenol matches 10 cc. of an alkaline solution containing 0.2 cc. of the same solution of the same indicator, it is known that the tested solution has induced a 20 per cent transformation of the indicator. Then α is 0.2. If now K_{IA} has been determined, and if it has been shown that the simple dissociation formula holds for the indicator in use, equation 10 may be solved for pH.

This procedure has been developed by Michaelis and coworkers; Biochem. Z. (1920) **109**, 165; Biochem. Z. (1921) **119**, 307; Deut. med. Wochenschr. (1920) **46**, 1238; **47**, 465, 673; Z. Nahr. Genussm. (1921) **42**, 75; Z. Immunitätsf. (1921) **32**, 194; Wochenschrift Brau. (1921) **38**, 107. Calculations are aided by the use of a table relating α to $\log \frac{\alpha}{1-\alpha}$. Such a table, somewhat more elaborate than that required for this special purpose, will be found on page 460 of the appendix.

It is obviously necessary that K_{IA} shall have been determined or that the actual experimental data relating the degree of color transformation to pH along the "dissociation curve" shall have been obtained. This necessary, fundamental "calibration" has been worked out by Michaelis and Gyemant (1920) and Michaelis and Krüger (1921) (using hydrogen electrode measurements as a basis) for a series of one-color indicators. In the following table are the pH values of the half-transformation points of the indicators used by Michaelis and Gyemant. These points correspond to $\log \frac{1}{K_{IA}}$ (see p. 26).

TABLE 36

pH values of the half-transformation points of indicators. After Michaelis

	TEMPERATURE				
	10°	20°	30°	40°	50°
2, 6 dinitro phenol	3.74	3.68	3.62	3.56	3.51
2, 4 dinitro phenol	4.11	4.05	3.99	3.93	3.85
p-nitro phenol	7.27	7.16	7.04	6.93	6.81
m-nitro phenol	8.43	8.32	8.21	8.09	7.99
Phenolphthalein	9.82	9.70	9.58	9.46	9.34
Alizarine Yellow GG	11.26	11.13	11.00	10.87	10.74

Now phenolphthalein, as we have already mentioned, is polyacidic with dissociation constants so close to one another that the simple equation of a mono acid cannot be used. Alizarine Yellow GG suffers the same disadvantage. Consequently it is necessary in these cases to abandon the simple equation and the dissociation constants given above and to tabulate the experimental data. Michaelis and Gyemant have given the following tabulations.

TABLE 37

Degree of color, α, shown by phenolphthalein at indicated pH values. Temperature 18°C.

α	pH	α	pH	α	pH
0.01	8.45	0.16	9.10	0.55	9.80
0.014	8.50	0.21	9.20	0.60	9.90
0.030	8.60	0.27	9.30	0.65	10.00
0.047	8.70	0.34	9.40	0.70	10.10
0.069	8.80	0.40	9.50	0.75	10.20
0.090	8.90	0.45	9.60	0.80	10.30
0.120	9.00	0.50	9.70	0.845	10.40
				0.873	10.50

TABLE 38

Degree of color, α, shown by alizarine yellow GG at indicated pH values. Temperature 20°C.

α	pH	α	pH
0.13	10.00	0.56	11.20
0.16	10.20	0.66	11.40
0.22	10.40	0.75	11.60
0.29	10.60	0.83	11.80
0.36	10.80	0.88	12.00
0.46	11.00		

For 2, 5-dinitrophenol $\log \frac{1}{K_{IA}}$ is 5.15 for solutions of very low salt concentrations, 5.08 for solutions of 0.15 M salt concentration and 5.02 for solutions of 0.5 M salt concentration.

For 3, 4-dinitro phenol $\log \frac{1}{K_{IA}}$ is about 5.3 and for 2, 3-dinitrophenol about 4.8.

With these data we are now prepared to measure pH values without the use of standard buffer solutions.

The following indicator solutions are used:

1. 2, 4 dinitro phenol (α dinitro phenol) 0.05 per cent aqueous solution
2. 2, 6 dinitro phenol (β dinitro phenol) saturated aqueous solution formed at high temperature and filtered from crystals.
3. 2, 5 dinitro phenol (γ dinitro phenol) 0.025 per cent aqueous solution.
4. 3, 4 dinitro phenol (δ dinitro phenol) concentration not given.
5. 2, 3 dinitro phenol (ϵ dinitro phenol) concentration not given.
6. p-nitro phenol 0.1 per cent aqueous solution.
7. m-nitro phenol 0.3 per cent aqueous solution.
8. phenol phthalein 0.04 per cent solution in 30 per cent alcohol.
9. Alizarine yellow GG (salicyl yellow, m-nitrobenzene azo salicylic acid) saturated alcoholic solution diluted to convenient strength.

Test tubes must be of equal bore. A measured amount of the solution to be tested (e.g. 10 cc.) is mixed with the proper indicator in such amount that a rather weak color is developed. To a second test tube containing 9 cc. N/100 NaOH there is added such a volume of the indicator solution that the color developed approximately matches that of the first tube. The volume of the second tube is now made up to the volume of the first. If the two tubes do not match in color, another trial is made with more or less indicator until a color match is obtained. The amount of fully transformed indicator in the second tube then corresponds to that amount of indicator in the first tube which has been transformed to the colored tautomer. Let us assume that 1.0 cc. was added to the tested solution and that a color match occurs when 0.1 cc. of the same indicator solution was placed in the second alkaline tube and made up to the volume of the first. Then the percentage color transformation induced by the tested solution was 10.

Hence $\alpha = 0.1$ and $\log \frac{\alpha}{1-\alpha} = -0.95$.

If the indicator used was p-nitrophenol and the temperature was 20°C. pH = 7.16 − .95 = 6.21 (6.2)

If the indicator was phenolphthalein table 37 shows that the pH was about 9.0.

For *routine* work in the range pH 2.8 to 8.4 Michaelis (1921) recommends the following system.

To uniform test tubes are added seriatim the volumes of indicator solution given in the following tables, the indicator solution being prepared by diluting the stock solutions (page 134) ten times with 0.1 normal soda solution (sic). Each tube is now filled to a 7 cc. mark with the soda (sic) solution. (In the original paper Michaelis and Gyemant describe dilutions with N/100 NaOH solution.)

TABLE 39

m-nitro phenol

Tube number	1	2	3	4	5	6	7	8	9
Cubic centimeters of indicator	5.2	4.2	3.0	2.3	1.5	1.0	0.66	0.43	0.27
pH	8.4	8.2	8.0	7.8	7.6	7.4	7.2	7.0	6.8

p-nitro phenol

Tube number	10	11	12	13	14	15	16	17	18
Cubic centimeters of indicator	4.05	3.0	2.0	1.4	0.94	0.63	0.4	0.25	0.16
pH	7.0	6.8	6.6	6.4	6.2	6.0	5.8	5.6	5.4

2, 5-dinitro phenol (γ dinitro phenol)

Tube number	19	20	21	22	23	24	25	26
Cubic centimeters of indicator	6.6	5.5	4.5	3.4	2.4	1.65	1.1	0.74
pH	5.4	5.2	5.0	4.8	4.6	4.4	4.2	4.0

2, 4-dinitro phenol (α dinitro phenol)

Tube number	27	28	29	30	31	32	33	34	35
Cubic centimeters of indicator	6.7	5.7	4.6	3.4	2.5	1.74	1.20	0.78	0.51
pH	4.4	4.2	4.0	3.8	3.6	3.4	3.2	3.0	2.8

The test tubes are sealed with paraffined corks and when not in use are protected from the light.

To test a solution for its pH value 6 cc. are taken and 1 cc. indicator solution added. The solution is then compared with the standards.

For testing the pH values of waters Michaelis (1921) operates as follows:

A stock solution containing 0.3 gram pure m-nitro phenol in 300 cc. distilled water is diluted before use by adding to 1 cc. of the stock 9 cc. distilled water. There are used flat bottom tubes of about 25 cm. height and 14 mm. internal diameter having such uniformity that 40 cc. of water will stand at a height of

between 22 and 23 cm. To six such tubes are added seriatim 0.25; 0.29; 0.33; 0.38; 0.45 and 0.50 cc. of the diluted m-nitro phenol solution. To each tube are added 40 cc. of an approximately N/50 NaOH solution freshly prepared by dilution of an approximately normal solution. These are the standards.

To test a water, 40 cc. are added to a tube of correct dimensions and to this is added sufficient indicator to develop a color within the range of the standards, preferably near the brighter of the standards. Comparison is now made as in Nesslerization, after having waited two minutes for completion of the mixing.

The amount of indicator in the alkaline, matching standard corresponds to the amount transformed to the colored form by the tested solution. Therefore, the cubic centimeters of indicator in the standard divided by the cubic centimeters in the tested solution is α, the degree of color transformation, or when multiplied by 100 the percentage color transformation.

Michaelis and his co-workers have tabulated corrections for temperature and for salt concentrations. The operator should determine for himself not only the order of accuracy required in his problem but his own ability to make readings with that precision which will make corrections significant. He may then refer to the original papers for tables giving corrections for salt effects and for temperature. The order of magnitude of these corrections may be seen in the following example.

For m-nitrophenol Michaelis and Krüger give the following values of $\log \frac{1}{K_{AI}}$ at 17°C. in solutions of the indicated salt concentrations.

TABLE 40

MOLECULAR SALT CONTENT	$\log \frac{1}{K_{AI}}$
0–0.01	8.33
0.05	8.28
0.10	8.23
0.15	8.22
0.20	8.18
0.3–0.6	8.17
to 1.0	8.15

The temperature corrections to be added when m-nitrophenol is used at temperatures other than 17.5°C. are as follows.

TABLE 41

t°	CORRECTION	t°	CORRECTION
5	+0.10	25	−0.06
10	+0.06	30	−0.11
15	+0.02	35	−0.15
17.5	±0.00	40	−0.18
20	−0.02	45	−0.22
		50	−0.26

In spite of the fact that the nitro compounds used by Michaelis and Gyemant are of wan color and those tried in the survey made by Clark and Lubs were neglected for this reason, Michaelis and Gyemant describe the application of their method to colored solutions.

Advantage is taken of the fact that many solutions are inappreciably altered in pH by diluting five or even ten times (see page 37). For dilution, Michaelis and Gyemant use freshly boiled NaCl solution of a concentration approximately that of the test solution. If on dilution the natural color still interferes with the use of an indicator, the natural color may be duplicated in the standard by the use of supplementary dyes such as Sørensen uses. Or, *if addition of alkali does not alter the natural color of the solution under test,* the standard may be made up with an alkaline solution of the tested solution itself. In this case it is necessary to be on guard against the buffer action and to add alkali until no increase in the color of the indicator is observed.

Without doubt the preferable procedure to follow when applying the Michaelis and Gyemant method or any other method to colored solutions is to use the "comparator" described on page 70 and illustrated in figure 8.

The method of Michaelis and Gyemant is fundamentally the same as that of Gillespie and should, therefore, be used with the qualifications which Gillespie has stated. There is a distinct advantage in the use of the nitro phenols for they have been found to have relatively small protein and salt errors. It is sometimes necessary to use very high concentrations of the indicator, and

in such circumstances one must be on guard against the effect of the indicator itself or of impurities.

Indicator paper. Although a favorite form of indicator is the deposit on a strip of paper (for example the familiar litmus paper) it is to be avoided unless the use of an indicator solution is precluded. It is to be avoided because the factors involved in the reaction between solution and indicator are made complex by the capillary action of the paper or the material entrained in these capillaries. On the other hand there are occasions when an approximate measure of pH is sufficient and when an indicator-paper is to be preferred. On such an occasion it is desirable to know the difficulties to be encountered. We are indebted to Walpole (1913) and others but particularly to Kolthoff (1919, 1921) for investigations on this subject. Kolthoff has given particular attention to the sensitivity of indicator papers when used in titrations, a situation where there is generally but little buffer action near the end-point. Under such circumstances there are to be regarded a number of details which are described at length in Kolthoff's papers. Several of these details will be perceived if we describe some of the more important aspects of the indicator-paper method of determining pH.

In general one must ride either horn of the following dilemma. The paper is sized, in which case the buffer action of the tested solution must be strong enough and allowed time enough to overcome the buffer action of the sizing. Or the paper has the qualities of filter paper, in which case the solution tested will spread and leave rings of different composition formed by the adsorptive power of the capillaries.

Kolthoff found that various treatments and selections of filter paper are of secondary importance, so the choice lies between sized and unsized paper. Now certain coloring matters are adsorbed by filter paper so that a separation is possible and the clear solution can be found in a ring about the point of contact between a tested colored solution and the indicator paper. But beyond this ring is a much more dilute one and unless one knows the properties of the system under examination it is not easy to correctly estimate the pH of the solution from the appearances of the paper.

Although coated paper may lose in sensitivity by not taking

up so much indicator as filter paper and must be used with strongly buffered solutions it is the more convenient. In any case the paper should be left in contact with the tested solution a generous length of time, for the establishment of equilibrium may be very slow (Walpole), and there must be instinctively exercised a mental plotting of the time curve.

If, after having exhausted all other methods, it is found that the indicator-paper method is the better adapted to a particular set of circumstances, the procedure should be calibrated to the purpose at hand rather than forced to render accurate pH values.

Dilution. As indicated in Chapter II a well buffered solution may often be moderately diluted without seriously altering the pH.

When dealing with complex solutions which are mixtures of very weakly dissociated acids and bases in the presence of their salts, and especially when the solution is already near neutrality dilution has a very small effect on pH, so that if we are using the crude colorimetric method of determining pH a five-fold dilution of the solution to be tested will not affect the result through the small change in the actual hydrogen ion concentration. Differences which may be observed are quite likely to be due to change in the protein or salt content. For this reason as well as for other reasons Clark and Lubs (1917) considered it wise to use M/20 standard comparison solutions instead of more concentrated standards for bacteriological media where dilution is often advantageous. The salt content of the standards undoubtedly influences the indicators and should be made as comparable as is convenient with the salt content of the solutions tested when these are diluted to obtain a better view of the indicator color.

The conclusion that dilution has little effect on the hydrogen ion concentrations of many solutions has long been recognized. Michaelis (1914) found little change in the pH of blood upon dilution, and Levy, Rowntree, and Marriott (1915) depended upon this *in part* in their dialysis method for the colorimetric determination of the hydrogen ion concentration of blood. Henderson and Palmer (1913) have used the dilution method in determining the pH of urines, and Paul (1914) records some experiments with wines the pH values of which were affected but little by dilution. The legitimacy of dilution has been tacitly admitted

by bacteriologists in their procedure of diluting media to be titrated to what is in reality a given pH as indicated by phenolphthalein.

In the examination of soil extracts colorimetrically little could be done were the "soil-solution" not diluted. Whatever may be the effect it is certain that the correlations between the pH values of such extracts and soil conditions is proving of great value (see Chapter XXI). Wherry has developed a field kit of the sulfon phthalein indicators with which he has found some remarkable correlations between plant distribution and the pH of the native soils. This field kit is now on the market.

The use of indicators in bacteriology. Perhaps no other science requires such continuous routine use of indicators as does bacteriology. This use is chiefly in the adjustment of the reaction of culture media and in the testing of the direction and limits of fermentation. While these are but examples, the frequency with which they become matters of routine warrant a brief outline of special procedures.

If experience has shown that the pH of the medium may lie within a zone about 0.5 units of pH wide, it is sufficient to add unstandardized acid or alkali, as the case may be, until a portion of the medium tested with the proper indicator in proper concentration approximately matches that color standard shown in the color chart of page 50 corresponding to the pH value to be approximated. This requires experience in overcoming the confusing effect of the natural color of the medium and also a well established sense of color memory. The beginner should proceed in some such way as the following.

It is desired, for instance, to adjust a colorless medium to pH 7.5. The medium as prepared is somewhat below the final volume. A quick, rough test at room temperature shows that the pH value lies between 6.0 and 6.5. Therefore, alkali must be added. The alkali solution to be used need not be standardized but may be about 1 normal. An exact one-in-ten dilution of this is run into 5 cc. of the medium to which has been added 5 drops of phenol red solution. Titration is continued until the color nearly matches 10 cc. of standard buffer "7.5." The tube of medium is now diluted to 10 cc. so that a color comparison may be made between test solution and standard, each contain-

ing the same concentration of indicator. The tubes are viewed through equal depths of solution. If the desired point 7.5 has been overstepped another trial is made. If 7.5 is not reached a moderate addition of alkali may be made without serious violation of volume requirements, and a second reading is taken.

Having made estimates of the pH values in the two readings an interpolation is made of the amount of dilute alkali required to bring the medium to exactly pH 7.5. From this is calculated the amount of the stronger alkali required for the main batch. Having added this a check determination is made. When finally adjusted the medium is diluted to its final volume. Most culture media suffer alterations of their pH values during sterilization and consequently allowance must be made if the final pH value is to be exact. This allowance will vary with the medium but can easily be determined for a standard medium prepared under uniform conditions.

When the color or turbidity of a tested solution interferes with direct color comparisons proceed as above but make comparisons by means of the Walpole compensation method described on page 70.

It need hardly be said that if the requirements of an organism are such that the desired pH value cannot be obtained with phenol red a more suitable indicator is selected from table 6 and if the medium requires the addition of acid an unstandardized acid solution approximately normal (or stronger) and an exact 1:10 dilution of this are substituted for the alkali solutions mentioned above.

In testing fermentations it often happens that the final hydrogen ion concentration is of more significance than the amount of acid or alkali formed; but the two distinct types of data are not to be confused to the complete displacement of either. It is sometimes convenient to incorporate the indicator with the medium provided the indicator is not reduced or destroyed by the bacterial action. The sulfon phthaleins are particularly useful for they are not reduced except by the more active anaerobes. Brom cresol purple replaces litmus in the common milk-fermentation tests without introducing that confusion which the reduction of litmus causes. It reveals differences in pH even beyond the range of its usefulness for ordinary measurements, passing from a

greyish blue in normal milk to more and more brilliant yellows with the development of acidity, and to a deep blue with the development of alkalinity. For further details see Clark and Lubs (1917).

In the method of Clark and Lubs (1915, 1916) for the differentiation of the two main groups of the coli-aerogenes bacteria, as well as in the similar method of Avery and Cullen (1919) for separating streptococci, the composition of the medium is so adjusted to the metabolic powers of the organisms, that the reaction is left acid to methyl red in one case, and alkaline in the other. No exact pH measurements are necessary. In cases where large numbers of cultures falling within the range of one indicator are to be tested, the cultures may be treated with the indicator and compared by grouping. A careful determination made upon one member of a homogeneous group will serve for all. In this way large numbers of cultures may be tested in a day.

Special uses. While on the subject of approximate determinations with indicators a word may be said about the usefulness of the quick, rough test.

Almost every investigator has come to realize that among the mistakes in labeling chemicals the more frequent are cases in which a basic salt is labeled as an acid salt and vice versa. Now a solution of Na_2HPO_4, for example, has a pH value, which, while easily displaced (see figure 5), distinguishes it definitely from a solution of NaH_2PO_4 or Na_3PO_4. Indeed some idea may be obtained of the degree of impurity if the pH value of a dilute solution is displaced definitely from about pH 8.5. Some serious accidents are said to have occurred in medical practice by the use of sodium citrate purported to be neutral and too late found to be acidic. One short, swift test with an indicator could have revealed the situation, and averted the accident.

Indeed there are a great many substances solutions of which have either a buffered and definite pH value, as in the case of acid potassium phthalate, or else a pH value easily displaced by impurities from that of the absolutely pure substance but still falling within limits, as in the case of the primary and secondary phosphates. When properly defined, such values can be made part of the specifications for purity. Especially in the case of drugs which are often used beyond the reach of the assay

laboratory a simple indicator test should prove helpful as suggested by Evers (1921) and especially emphasized by Kolthoff (1921).

In the case of milk it is quite impossible to define the pH by a comparison of the color of an indicator in the milk with the color of the indicator in a clear standard; yet differences are made distinctly evident, and, if taken only for what they actually mean, are helpful in the grading of milk and in the study of the conduct of different bacteria inoculated into sterile milk. Clark and Lubs (1917) called attention to the superiority of the sulfonphthalein indicators, especially brom cresol purple, for this purpose.

Spotting. When only small quantities of solution are available or when highly colored solutions are to be roughly measured, their examination in drops against a brilliant white background of "opal" glass is often helpful. In the examination of colorless solutions comparisons with standards may be made as follows. A drop of the solution under examination is mixed with a drop of the proper indicator solution upon a piece of opal glass. About this are placed drops of standard solutions and with each is mixed a drop of the indicator solution used with the solution under examination. Direct comparison is then made. Felton who developed details in this method for the examination of small quantities of solutions used in tissue-culture investigations found mixtures of indicators of particular value for orientation. Equal parts of methyl red and brom thymol blue, for instance, give brilliant color contrasts in this drop method between about pH 4.6 and 7.6; but with an unsatisfactory zone between 5.6 and 6.2. Methyl red and brom cresol purple are used between pH 4.6 and 7 while for rough work between 1.2 and 9 methyl red and thymol blue are used. These mixtures are used only as "feelers." The opal glass or porcelain upon which the tests are to be made should be carefully tested for the absence of material seriously affecting the acid-base equilibria of the material under examination. Errors due to unequal drops, evaporation and impurity of indicator are to be watched for. For further details see Felton (1921).

CHAPTER IX

Outline of the Electrometric Method

A noble metal coated with platinum black, which will hold large quantities of hydrogen, may be made to serve as a hydrogen electrode. When it is laden with hydrogen and immersed in a solution containing hydrogen ions, there is exhibited a difference of electrical potential between solution and electrode which is dependent upon the concentration of the hydrogen ions; just as the potential difference between a silver electrode and a solution of silver ions is dependent upon the concentration of the silver ions.

We have no reliable means of measuring this single potential difference; but when we join two hydrogen electrodes, as shown in figure 13, we can not only measure the difference between the aforementioned differences of potential, i.e., the total electromotive force (E. M. F.) of the "gas chain" as it is called, but we can also derive an equation showing how this E. M. F. will vary with the *ratio* of the concentrations of the hydrogen ions about the two electrodes. If C is the concentration of the hydrogen ions in one solution and C′ the concentration in the other, the E. M. F. of the combination will be related to the ratio of the concentrations by the following equation expressed in numerical form for a temperature of 25°C.

$$\text{E. M. F.} = 0.059 \log \frac{C}{C'}$$

We shall leave to the next chapter the derivation of the equation and shall now put it in a form not restricted to the particular temperature of 25°C. assumed above.

$$\text{E. M. F.} = 0.000{,}198{,}37\ T \log \frac{C}{C'}$$

Here T is the absolute temperature, the zero point of which is 273.09° below 0°C. A table giving the values of 0.000,198,37 T for various temperatures centigrade is given in the Appendix. Thus if we join two hydrogen electrodes as illustrated in figure 13 measurements of the electromotive force of the chain and of

the temperature allow us to calculate the ratio of the one hydrogen ion concentration to the other. Then if one hydrogen ion concentration is known we may derive the other.

As the "known" there may be used any one of the buffer solutions described in Chapter VI. The reader should note, however, that the values of these "known" solutions are derived from

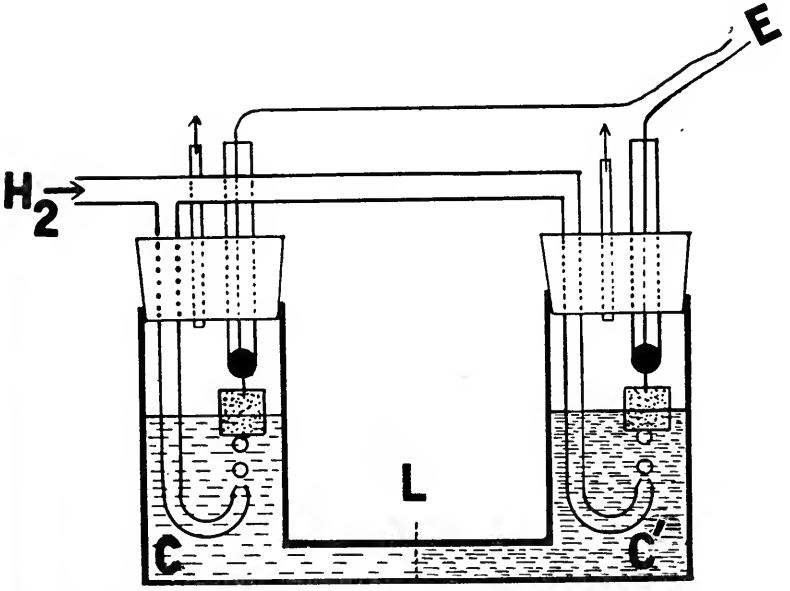


FIG. 13. DIAGRAM OF CONCENTRATION CHAIN OF HYDROGEN ELECTRODES

hydrogen electrode measurements which, as we have just seen, furnish ratios only. Some ultimate standard is therefore implied. This is discussed in Chapter XIX.

If there be no means at hand for measuring the electromotive force but there is available a galvanometer or a home-made capillary electrometer for detecting small currents, the following procedure may be used. Two hydrogen electrodes are set up as in figure 13. By means of the buffer solutions described in Chapter VI the hydrogen ion concentration in one electrode vessel is varied until no difference of potential occurs between the two electrodes. This point is determined by absence of deflection

in the galvanometer or by no change in the meniscus of the capillary electrometer. Then $C' = C$ in the above equation.

Instead of setting up two hydrogen electrodes, one of which is a known standard, it is generally more convenient to replace the standard hydrogen electrode by a more permanent "half cell" such as the "calomel electrode." This is an electrode of mercury covered with calomel in the presence of a definite KCl solution, for example saturated KCl solution. If this so-called "saturated calomel electrode" is used, a tube containing saturated KCl is led directly to the solution in the hydrogen electrode vessel.

Now suppose that in the first place there were used two hydrogen electrodes as in figure 13, and let it be assumed that one of these was immersed in a solution normal with respect to hydrogen ions. Let C be identified as 1 normal and C', the unknown be less than 1 normal. Then E. M. F. $= 0.000,198,37\ T \log \frac{1}{C'}$.

Now suppose that the normal hydrogen electrode is connected with a "saturated calomel electrode." We might then have an arrangement as follows:

$$\left.\begin{array}{ll} \text{II} & \left\{\begin{array}{l}\text{saturated calomel electrode}\\ \\ \text{normal hydrogen electrode}\end{array}\right. \\ \text{III} & \left\{\begin{array}{l}\\ \\ \text{hydrogen electrode in } [H^+] = C'\end{array}\right. \end{array}\right\} \text{I}$$

Suppose the difference II has already been determined and that I is measured in the immediate experiment. Then I − II = III. Having found III, we can use the equation for two hydrogen electrodes, one of which is the "normal," and so solve directly for C'.

At 25°C. the mercury of the calomel electrode is 0.246 volt more positive than the platinum of the normal hydrogen electrode.

Hence: observed E. M. F. $- 0.246 =$ III

$$\text{I} - \text{II} = \text{III}$$

$$\text{III} = 0.000,198,37\ T \log \frac{1}{C'}.$$

At 25°C., $T = 273.09 + 25 = 298.09$.

Then observed E. M. F. $-$ 0.246 = 0.0591 log $\frac{1}{C'}$.

But log $\frac{1}{C'}$ = pH.

$$\frac{\text{Observed E.M.F.} - 0.246}{0.0591} = \text{pH.}$$

If the observed E. M. F. is 0.648, pH = 6.80.

Although it is impracticable to describe at this point the details of a complete system for the measurement of hydrogen ion concentration an outline may be given with which to coordinate the main features as they will develop in subsequent chapters.

Figure 14 illustrates a simple system which may be put together from inexpensive material. It is not a system which can be recommended for precise or even routine measurements, but it will work and is well adapted to show the principles concerned.

Hydrogen, prepared by one of the methods described in Chapter XV, passes into the hydrogen electrode vessel A and escapes at B. Connected with this vessel by the siphon S, filled with a saturated KCl solution, is the calomel electrode M consisting of a layer of mercury covered by calomel under a saturated solution of KCl. The hydrogen electrode H consists of a piece of platinum foil covered with platinum black. It is welded to a platinum wire which is sealed into the glass tube.

Hydrogen is bubbled through the solution in A until solution and electrode are thoroughly saturated with the gas.

The difference between the potential difference at the mercury-calomel junction and the potential difference at the hydrogen electrode-solution junction is now measured by means of a potentiometer. A simple form of this is illustrated.

A storage battery P sends current through the rheostat R, the calibrated resistance-wire K-L and the fixed resistance L-F. By properly setting the switch O a Weston cell W having an electromotive force of 1.018 volts can be connected to K and F, the + pole of the Weston cell being connected to the + side of the battery current. The rheostat R is now varied until there is no deflection of the galvanometer or electrometer E. Then the difference of potential between K and F is equal to the E. M. F. of the Weston cell. The resistance K-L is such that when the

above adjustment is made the difference of potential between K and L is one volt. A scale properly divided is placed beside the wire K-L. When the sliding contact X is at K there will be no difference of potential between X and K. When X is at L the difference of potential between X and K will be one volt. When X is at some intermediate position the difference of potential between X and K will be that fraction of one volt indicated by the scale.

Having first carefully adjusted the potentiometer by means of the standard Weston cell the switch O is thrown to connect the calomel electrode-hydrogen electrode system and X is slid in one direction or the other until the galvanometer E shows no deflection. Then the difference of potential between X and K is equal to the difference of potential between mercury and platinum.

The temperature is read and the data put into the equations given above.

Neither measured E. M. F. nor Weston cell should be left in circuit for more than an instant. While switch O can be used for this momentary completion of circuit, it is more convenient to use a telegraph key in the galvanometer circuit.

If care be taken to maintain the hydrogen at barometric pressure, the effects of minor variations of the barometer from sea level conditions and of displacement of hydrogen by water vapor may be neglected in rough measurements. A discussion of the barometric pressure is found in the next chapter.

In all cases where two unlike solutions are joined as in figure 13, there will develop a local potential difference at the liquid junction. To deal with this precisely is the most difficult of the problems encountered. The subject is discussed in Chapter XI. In very many instances, however, the employment of a saturated solution of KCl as is specified in the apparatus illustrated in figure 14, reduces the liquid junction potential difference to an order of magnitude which is negligible.

Since variations may occur in the calomel electrode or in the reliability of the hydrogen electrode it is well to check the system frequently by means of measurements made with standard solutions. Some of these are described in Chapter XVIII.

In the use of the potentiometer the elementary principles must be understood lest standard cells or half-cells be injured

or quite erroneous results obtained. Therefore, these principles are discussed in Chapter XIV.

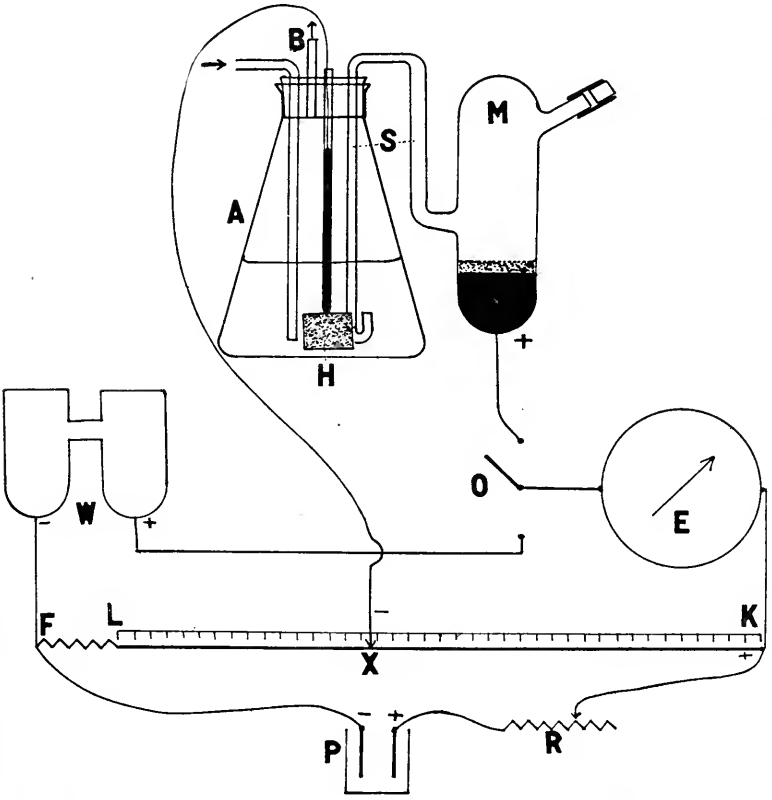


FIG. 14. A SIMPLE ARRANGEMENT FOR ELECTROMETRIC MEASUREMENT OF pH.

Were it not for the fact that several experimenters have tried to make hydrogen electrode measurements by use of conductivity instruments, it would seem hardly necessary to say that the measurement of conductivity or its reciprocal, resistance, is a procedure entirely different from the measurement of electromotive forces or potential differences.[1]

[1] The surprising number of cases in which this confusion has been revealed may be an interesting psychological result of the emphasis hitherto placed upon conductivity measurements, sometimes to the entire exclusion of any reference to potentiometric measurements.

If the beginner is puzzled by the array of apparatus described in the following pages he may welcome the following suggestion. The main outline of a problem can often be defined by the use of the Hildebrand electrode used in connection with the saturated calomel half-cell and by using as a potentiometer the voltmeter system. This set of apparatus is illustrated in figure 28. It not infrequently happens that the outlining of a problem with this or a comparable system will indicate that further refinement would be useless or confusing. It also frequently happens that the errors suggest phantom relations or obscure existing relations of importance. It is, therefore, advisable whenever possible to keep the accuracy of measurements just ahead of the immediate demands. To meet this requirement the investigator must gain the ability to judge for himself the apparatus required and it is to contribute toward this and the pleasure of work that the following chapters are written in some detail. If the reader does not care to work out the peculiar requirements of his problem he is advised, after having outlined his problem with the system mentioned above, to obtain a reliable potentiometer of standard, not unique, design and to use the system illustrated in figure 19. In the first instance accurate temperature control is unnecessary. In the second instance it is advisable if for no other purpose than the avoidance of vexatious uncertainties.

CHAPTER X

Theory of the Hydrogen Electrode

In treating the theory of the hydrogen electrode we shall first consider Nernst's (1889) conception of electrolytic solution tension as a useful way of remembering certain important relations and then pass to the thermodynamic derivation of the E. M. F. of a concentration cell.

If a metal be placed in a solution of its salt there will be a difference of electrical potential between metal and solution which will vary in an orderly manner with the concentration of the metal ions. To account for the difference of potential Nernst assumed that a metal possesses a characteristic solution tension comparable with the vapor pressure of a liquid, or, better, with the solution pressure of a crystal of sugar—but with the important qualification that it is the metal ions which pass into solution. Imagine first that the metal is in contact with pure water. The metal ions passing into solution carry their positive charges and leave the metal negative. Thus there is established a so-called double layer of electrical charges at the interface between metal and solution, the solution being positively and the metal negatively charged relative to one another. This potential difference forcibly opposes further dissolution of metallic ions, for the relative positive electrical field in the solution and the relative negative field in the metal restrain any further migration of positively charged bodies from the metal to the solution. Equilibrium is established when the electrostatic control equalizes the solution pressure.

If now there are already in the solution ions of the metal, the relative electrostatic field in the solution has already been partially established, fewer ions will escape from the metal and the metal is left more positive.

Therefore the higher the concentration of the positive metallic ions in the solution the more positive will be the charge on the metal and, conversely, the lower the concentration of the metallic ions in the solution the more negative will be the charge on the metal.

Not only metals but various gases are found to act in a similar way when means are devised to bring them into a situation as easily handled as are metal electrodes. Hydrogen is one of these gases and the means of handling it as an electromotively active gas is to take it up in one of those metals such as platinum, palladium or iridium which in a finely divided condition hold large quantities of hydrogen. Platinum black deposited upon platinum and laden with hydrogen forms a hydrogen electrode. It can be brought into equilibrium with hydrogen ions as silver is brought into equilibrium with silver ions; and the more positive it becomes the higher must be the concentration of the positively charged hydrogen ions in the surrounding solution.

It remains however to formulate with mathematical precision the way in which the potential of the hydrogen electrode changes with the concentration of the hydrogen ions; and for this purpose the energy relations must be considered.

Suppose a metal electrode dips into a solution of ions of the same metal. Let the concentration of these ions be such that their partial pressure, which would be manifest in an arrangement for producing osmotic pressure, is P in the volume V.

Let the electrode be of such a size that one gram mol of ions, carrying nF faraday of electricty, can pass from electrode to solution to there raise the partial pressure by dP. The increase of the difference of potential between electrode and solution will be dE. The electrical work expended will then be nFdE and the work taken up by the material system will be VdP. If the process is reversible, and the system is allowed to return to the original state,

$$nFdE - VdP = 0$$

or

$$dE = \frac{VdP}{nF}. \tag{26}$$

We shall now assume that we are dealing with an "ideal solution" by which we mean a solution in which the pressure-volume relation of the ions conforms to the gas law for a "perfect gas," then $PV = RT$ or $V = \frac{RT}{P}$.

Substituting this equivalent of V in equation (26) we have

$$dE = \frac{RT}{nF}\frac{dP}{P}$$

On integration this becomes

$$E = \frac{RT}{nF} \ln P + K \qquad (27)$$

where ln is the symbol for natural logarithm to the base e and K is an integration constant.

The integration constant is the point of reference for the general relation $E = \frac{RT}{nF} \ln P$. It is the potential difference between electrode and solution when some arbitrary unit of pressure is chosen and P = 1. Then in accordance with the unit chosen E = K.

LeBlanc (1907) and others have substituted for K an equivalent constant of the form $-\frac{RT}{nF} \ln p$, called p the electrolytic solution tension of Nernst and so obtained the relation

$$E = \frac{RT}{nF} \ln \frac{P}{p}.$$

But it is of doubtful value to postulate the physical significance of K in this manner. For present purposes we can afford to leave K as it stands, a pure integration constant.

Let us consider now the arrangement known as a concentration cell. Let the two vessels of figure 13 contain the same metal ion in concentrations C and C′ corresponding to "osmotic pressures" P and P′. Let there dip into each solution an electrode of the metal. Let the two solutions be connected by a siphon, and the electrodes by a device for measuring the E. M. F.

Using the equation (27) developed above we know that at electrode 1 there will be a difference of potential $E = \frac{RT}{nF} \ln P + K$ and at electrode 2 a difference of potential $E' = \frac{RT}{nF} \ln P' + K$. The

total E. M. F. will be the algebraic sum of these potential differences. If P′ be less than P, the electrode in contact with the ions at partial pressure P′ will be negative to the electrode in contact with the ions at partial pressure P. Hence

$$\text{E. M. F.} = E - E' = \frac{RT}{nF} \ln P + K - \left[\frac{RT}{nF} \ln P' + K\right] = \frac{RT}{nF} \ln \frac{P}{P'}.$$

Since the ratio of the pressures may be considered equal to the ratio of the ion concentrations,

$$\text{E. M. F.} = \frac{RT}{nF} \ln \frac{C}{C'} \tag{28}$$

This is the fundamental equation for the E. M. F. of a concentration chain.

R is the gas constant, T the absolute temperature, (273.09+ t centigrade), n the valency of the ion and F the faraday or the quantity of electricity associated with 1 gram molecule equivalent.

To put this equation into working form there have to be found the electrical equivalents for R and F. Since measurements of potential are to be made in terms of the international volt this and the related units will first be defined as they are given in Bureau of Standards Circular No. 60 (1916), "Electrical Units and Standards."

International ohm. The international ohm, which is generally referred to as the ohm, but which is to be distinguished as are other international units from the "absolute" units, is defined as "the resistance offered to an unvarying electric current by a column of mercury at the temperature of melting ice, 14.4521 grams in mass, of a constant cross-sectional area and of a length of 106.300 cm."

International ampere. The international ampere, generally referred to as the ampere, is defined as "the unvarying electric current which, when passed through a solution of nitrate of silver in water in accordance with specification II (of the 1908 London Conference), deposits silver at the rate of 0.00111800 of a gram per second."

International volt. The volt is derived from current and resistance in accord with Ohm's law, $C = \frac{E}{R}$. The international

volt is therefore defined as "the electrical pressure (electromotive force) which, when steadily applied to a conductor the resistance of which is one international ohm, will produce a current of one international ampere."

F, the faraday, is derived for the international system as follows. The international ampere deposits silver at the rate of 0.00111800 of a gram per second. Since the atomic weight of silver is 107.88, a gram equivalent would be deposited in one second by 96494 amperes. The coulomb (international) is the quantity of electricity transferred by a current of one international ampere in one second. Hence 96494 coulombs are carried by a gram equivalent of silver and this is the value of the faraday in the international system.[1]

Returning now to equation (28) we know that R, the gas constant, is derived from the gas equation

$$PV = \frac{P_oV_o}{273.09}T, \text{ where } \frac{P_oV_o}{273.09} \text{ is R.}$$

V_o, the volume of 1 gram molecule of an ideal gas at one atmosphere pressure and 0°C. is 22412 ± 2 cc. (Berthelot, 1904). P_o = one atmosphere or 76 cm. of mercury at 0°C. and 45° latitude. Since the acceleration of gravity at 45° latitude was taken to be 980.665 cm. per second when the "atmosphere" was defined, and, since 1 cc. mercury under the action of such a gravitational pull weighs 13.59545 grams, $P_o = 980.665 \times 76 \times 13.59545$ or 1013276 dynes per square centimeter.

$$\text{Hence R is } \frac{1013276 \times 22412}{273.09} = 83157719.8 \text{ ergs.}$$

10^7 ergs = one joule absolute. One joule, absolute = 0.99966 international joule. Hence R = 8.3129446 international joules, or volt coulombs.

From the derivations outlined above our equation reduces to the numerical form

$$E = \frac{8.3129446}{96494}\frac{T}{n}\ln\frac{C}{C^1}$$

[1] The absolute value is approximately 96,500 (Vinal and Bates, 1914).

Transposing to Briggsian logarithms (to the base 10) by dividing by 0.43429 we have

$$E = 0.00019837 \frac{T}{n} \log \frac{C}{C^1} \qquad (29)$$

In the case of the hydrogen electrode, where the valence of the ionic hydrogen concerned is one, n is generally not written.

A table of the values of 0.00019837 T for various temperatures is given in the Appendix.

The significance of the equation for the concentration chain is that, if T is known, and if the concentration of the ions in the other solution is known, then the concentration of the ions in one solution can be determined from the E. M. F. of the chain. Fundamentally there is no other way of applying electromotive force determinations to the estimation of ion concentrations, unless there can be brought to bear mass action relations. This makes it necessary to start somewhere in the system with a solution whose hydrogen ion concentration has been determined by an independent method.

Let us assume for the moment that the conductivity method will give us correct information upon the hydrogen ion concentration of some simple solution such as that of HCl.

It will be remembered that hydrogen ion concentrations are expressed in terms of normality, a solution normal with respect to hydrogen ions being one which contains in one liter of solution 1 gram[2] of hydrogen ions.

If, then, the normality of the hydrogen ion concentration in any unknown solution is to be determined it would seem that the most convenient solution with which to compare the unknown would be a solution of normal hydrogen ion concentration. Between a hydrogen electrode in this standard and a hydrogen electrode in the unknown solution of hydrogen ion normality Cx there would be a difference of potential, E, given by the equation

$$E = 0.000,19837\ T \log \frac{1}{Cx} \qquad (30)$$

[2] It makes little difference whether we regard the atomic weight of hydrogen as 1.0 or as 1.008 for the purpose at hand.

A measurement of E and T would give Cx. Now E in the above equation is the difference between the potential difference at the one hydrogen electrode and the potential difference at the other hydrogen electrode. Nothing need be known about the value of either single potential difference and very little is known. If the electrode in the normal solution is made the standard it is obviously convenient for present purposes to call the potential difference between this electrode and the solution zero. Thus arises the definition: *The potential difference between a hydrogen electrode under one atmosphere pressure of hydrogen and a hypothetical solution normal with respect to the hydrogen ion shall be considered to be zero at all temperatures.*[3]

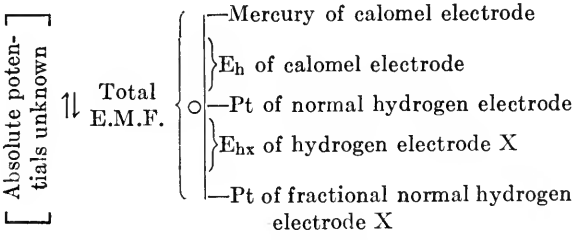
Having established by definition the value of the potential difference at the normal hydrogen electrode it becomes convenient to speak of the potential difference at any other hydrogen electrode as the hydrogen electrode *potential,* thus abbreviating the term "potential difference." It is, of course, implied that such a "potential" is referred to the potential difference at the normal hydrogen electrode. To indicate this the symbol E_h is used.

Unfortunately it has been necessary to introduce into the definition of the normal hydrogen electrode the phrase "*hypothetical* solution normal with respect to the hydrogen ion." This is because that very desirable standard solution would have to be prepared by means of "strong" acids and the estimation of the hydrogen ion concentration would fall under those uncertainties which have already been mentioned in a previous chapter. The difficulty is not entirely obviated by making the experimental standard a more dilute solution of a strong acid as has been done; but we shall leave to Chapter XIX further discussion of this problem, and, for the moment, we shall assume that there can be constructed from measurements such as those of the conductivity method a solution having a definite, known hydrogen ion concentration. We could proceed with this, using it as one of two solutions in a hydrogen gas cell, and applying to this cell the

[3] In various places, notably in the report of the Potential Commission of the Bunsen-Gesellschaft (Abegg, Auerbach and Luther, 1910) it is not specifically stated that this difference of potential shall be *zero at all temperatures,* but it seems to have been so understood and is specifically stated by Lewis (1913).

formula relating the E. M. F. to the ratio of the known to the unknown hydrogen ion concentration. But it is more convenient to use as a working-standard a calomel half cell (see Chapter XIII). When this is joined to a hydrogen electrode to form a calomel-hydrogen cell we need to know the difference of potential between the calomel half cell and some known hydrogen electrode. Then we can correct the observed E. M. F. by this difference and consider the corrected E. M. F. to be as if it were that between two hydrogen electrodes.

Remembering that the mercury of the calomel half cell is positive to the platinum of the normal hydrogen electrode and that the platinum of a hydrogen electrode becomes more negative the more dilute the hydrogen ion concentration, we have the scheme shown below



If E. M. F. is measured and E_h is known, the value of E_{hx} is at once obtained. This is the difference of potential between two hydrogen electrodes and equation (29) applied. In its working form this equation is:

$$\frac{\text{E.M.F. (observed)} - E_h \text{ (of calomel half cell)}}{0.000,198,37\,T} = \log \frac{1}{[H^+]} = \text{pH} \quad (31)$$

The above equation is still incomplete because we have not taken into consideration the liquid junction potential differences which exist wherever two unlike solutions are brought into contact. Nor have we yet considered the effect upon the potential difference at a hydrogen electrode of a change in the pressure of hydrogen from the one atmosphere partial pressure specified for the normal hydrogen electrode. These two will be considered from the point of view of corrections to be made. Liquid junction potential differences, because of their distinct importance, will be treated in a separate chapter.

BAROMETRIC CORRECTION

The potential difference between a metal and solution will vary somewhat with the condition of the metal. A hammered, twisted or scratched electrode may show a different potential against a given concentration of its ions than will an electrolytically deposited metal. In the case of the hydrogen electrode it seems to make little difference whether the hydrogen be held in platinum, palladium or iridium but it does make a considerable difference if the surrounding pressure of hydrogen varies. If we have two hydrogen electrodes immersed in the same solution at the same temperature but under different pressures of gaseous hydrogen, we may assume that the concentration of the hydrogen in one electrode is different from that in the other electrode, and that the potential difference may be expressed as

$$E = E_1 - E_2 = \frac{RT}{nF} \ln \frac{[H]_1}{[H]_2} \qquad (32)$$

in which equation R, T, n, and F have their customary significances and $[H]_1$ and $[H]_2$ are concentrations of *atomic* hydrogen in the electrodes (platinum black). Since n, the valence of hydrogen, is 1, it may be omitted.

We may now assume that there is an equilibrium between the molecular hydrogen about the electrode and the atomic or ionic hydrogen in, or issuing from, the electrode. This equilibrium may be expressed in accordance with the mass law as follows:

$\frac{[H] \times [H]}{[H_2]} = K_t$ where $[H]$ = concentration of atomic hydrogen

and $[H_2]$ = concentration of molecular hydrogen

Whence,

$$[H] = \sqrt{K_t[H_2]} \qquad (33)$$

Substituting (33) in (32), we have

$$E = \frac{RT}{F} \ln \frac{\sqrt{K_t [H_2]_1}}{\sqrt{K_t [H_2]_2}} = \frac{RT}{2F} \ln \frac{[H_2]_1}{[H_2]_2}$$

It should be noted that the factor 2 in this equation does not come from giving hydrogen an effective valence of 2, as has often been stated, but from the introduction of equation (33). We

might however derive the equation more directly by the energy relations and then the factor 2 would enter by reason of the volume change involved.

If the ratio of pressures is equal to the ratio of gas concentrations

$$E = \frac{RT}{2F} \ln \frac{P'_{H_2}}{P_{H_2}}$$

If P'_{H_2} be one atmosphere and P_{H_2} be expressed in atmospheres

$$E = \frac{RT}{2F} \ln \frac{1}{P_{H_2}} \tag{34}$$

This is the equation for the difference of potential between a hydrogen electrode under one atmosphere pressure of hydrogen (e.g., the normal hydrogen electrode) and a hydrogen electrode under pressure P_{H_2}.

Experimental justification of this equation is found in the experiments of Czepinski, Lewis and Rupert, Lewis and Randall, Lewis and Sargent, Ellis, Loomis and Acree and others.

Hainsworth and MacInnes have studied the effect of pressures up to 400 atmospheres and taking account of the volume changes of Hg, calomel, etc. which are negligible for smaller differences in pressure, they find a linear relation except for a slight deviation at the higher pressures.

Several writers have felt constrained to emphasize the fact that in determining the hydrogen pressure from barometer readings they have subtracted the vapor pressure of the solution. The emphasis is still advisable, for a considerable number of precise hydrogen electrode data are published with corrections for barometric pressure on the basis that the normal hydrogen electrode pressure is one atmosphere *including* the vapor pressure of the solution. Corrections should be made to one atmosphere pressure of *hydrogen*, or else the standard used should be distinctly specified.

Clark and Lubs (1916) have suggested that a more consistent standard than that now recognized for the normal hydrogen electrode would be obtained by defining a standard *concentration* of hydrogen rather than a standard pressure. They used the commonly accepted "standard condition" of a gas which is the con-

centration at 0°C. and 760 mm. pressure. This would bring both the hydrogen and the hydrogen ions to a concentration basis, whereas now the one is expressed in terms of concentration and the other in terms of pressure.

In applying the correction,

$$E_{bar.} = \frac{RT}{2F} \ln \frac{1}{P_{H_2}},$$

it will be remembered that a decrease of the hydrogen pressure may be considered as a decrease of the electrolytic solution tension of the hydrogen. Hence under decreased hydrogen pressure the electrode is left more positive.

In the cell

$$Hg \mid Hg_2Cl_2KCl \mid H^+ \mid Pt \mid H_2$$

if the hydrogen is under diminished pressure the E. M. F. of the cell is too low. Hence the correction must be applied to make the E. M. F. larger than observed. Equation (31) becomes:

$$\frac{E.M.F. + E_{(bar.)} - E_{(calomel)}}{0.00019837\ T} = pH \qquad (35)$$

To aid in the calculation of pressure corrections it is convenient to plot a curve giving the millivolts to be added to the observed E. M. F. for various corrected partial pressures. Tables of corrections from which a chart may be plotted are given in the Appendix. In these tables the barometer pressures given are the corrected pressures. If hydrogen escapes from about the hydrogen electrode through a trap or other device which exerts back pressure, this pressure must be taken into consideration. Otherwise it is assumed that the pressure of the hydrogen is that of the barometer less the vapor pressure of the solution. To obtain the corrected barometer reading the instrumental calibration of the instrument is first applied, then the temperature correction (a table of which is given in the Appendix) necessary to bring the height of the mercury column at temperature t to its height at temperature 0°C. Then there remains the correction for locality (see tables in Landolt-Börnstein) in order that the pressure may be reduced to the common basis of the "atmosphere," namely, the pressure of 760 mm. mercury where the acceleration of gravity is

980.665 cm. per second. The last correction is of significance only for very accurate measurements and exceptional localities.

For all ordinary cases it may be assumed that the vapor pressure is that of pure water at the temperature indicated.

If the unit pressure is one atmosphere, the partial pressure must be reduced to atmospheres.

As inspection of the table in the Appendix will indicate, the barometric correction may be neglected in rough measurements.

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CHAPTER XI

Potential Differences at Liquid Junctions

When two unlike solutions of electrolytes are brought into contact there develops at the junction a potential difference. Since no important chain can be constructed without involving such a liquid junction potential, it is of great importance to know the cause so that the magnitude of the potential may be calculated or ways devised for its reduction.

The principal cause of the potential difference was attributed by Nernst to unequal rates of diffusion of ions across the plane of junction.

It has been found in the study of electrolytic conduction that under uniform potential gradient different ions move through a solution with different velocities. The following table taken from Lewis' *A System of Physical Chemistry* shows the velocities of a number of ions in aqueous solution under a potential gradient of one volt per centimeter. Since in each case the potential gradient is the same and the ionic charge the same it is evident that the order in which the velocities stand in the table is the order in which the velocities of free movement will stand.

ION	ABSOLUTE VELOCITY IN CENTIMETERS PER SECOND. 18°C.	ION	ABSOLUTE VELOCITY IN CENTIMETERS PER SECOND. 18°C.
H	$32.50\ 10^{-4}$	OH	$17.80\ 10^{-4}$
K	$6.70\ 10^{-4}$	Cl	$6.78\ 10^{-4}$
Na	$4.51\ 10^{-4}$	NO_3	$6.40\ 10^{-4}$
Li	$3.47\ 10^{-4}$	CH_3COO	$3.20\ 10^{-4}$
Ag	$5.70\ 10^{-4}$		

Let it now be assumed that a solution of hydrochloric acid is placed in contact with pure water of negligible ion content at an imaginary plane surface. Independently of one another the chlorine and the hydrogen ions will *tend* to migrate across the interface and into the water. As shown in the above table the velocity of the hydrogen ion under the influence of a potential gradient

is much greater than the velocity of the chlorine ion under the same gradient, and the relative velocities of free movement must therefore be in the same proportion. Consequently there will be established on the water side of the plane an excess positive charge. This charge will increase until the electrostatic attraction dragging the slower moving chlorine ions brings them to the velocity of the hydrogen ions. When this state is reached, as it is almost instantaneously, there is established a steady potential difference at the liquid junction. If the water is replaced by a solution of an electrolyte, we have not only the chlorine and the hydrogen ions migrating across the boundary into this new solution, but the ions of this solution migrating into the hydrochloric acid solution.

In the comparatively simple case where two solutions of different concentration of the same binary electrolyte are placed in contact the following elementary treatment may be used. Let the concentration of the ions on one side of the interface be C and on the other side be a lesser concentration C′.

When migration has established the steady potential E let it be over an interface of such extent that E is due to the separation of one faraday. If that fraction of the separated charge which is carried by the anion is n_a the work involved in the transport of n_a equivalents from C to C′ is $n_a RT \ln \frac{C}{C'}$. Likewise if that fraction of the charge carried by the cations is n_c the work involved in the transport of n_c equivalents from C to C′ is $n_c RT \ln \frac{C}{C'}$. The work involved in the *separation* of the ions as they migrate from the high to the low concentration is

$$n_a RT \ln \frac{C}{C'} - n_c RT \ln \frac{C}{C'} = EF$$

Whence

$$E = (n_a - n_c) \frac{RT}{F} \ln \frac{C}{C'} \text{ or } (n_c - n_a) \frac{RT}{F} \ln \frac{C}{C'} \qquad (36)$$

according to which ion moves the faster.

Now the ions being univalent, n_a, the fraction of the charge carried by the anion, is equal to the fraction N of one equivalent of anions transported from the cathode to the anode section. Like-

wise n_c is 1-N. The ratio of N to 1-N is equal to the ratio of the absolute velocities of the ions.

$$\frac{N}{1-N} = \frac{\text{velocity of anion } (V_a)}{\text{velocity of cation } (V_c)}$$

Whence

$$N = \frac{V_a}{V_a + V_c}, \text{ transport number of anion,}$$

and

$$1 - N = \frac{V_c}{V_a + V_c}, \text{ transport number of cation.}$$

Substituting N for n_a and $1 - N$ for n_c in equation (36)

$$E = \frac{(V_a - V_c)}{(V_a + V_c)} \frac{RT}{F} \ln \frac{C}{C'} \tag{37}$$

Lewis and Sargent (1909) have treated the special case of two equally concentrated solutions of two binary salts having one ion in common. Substituting equivalent conductivities as proportional to mobilities they obtain

$$E = \frac{RT}{F} \ln \frac{\lambda_1}{\lambda_2} \tag{38}$$

where λ_1 and λ_2 are the equivalent conductivities of two solutions. Applying this equation they obtain the following correspondence between calculated and observed values of E, the liquid junction potential.

SOLUTIONS IN CONTACT	E (OBSERVED)	E (CALCULATED)	E (OBS.)-E (CALC.)
0.2N KCl–0.2N $KC_2H_3O_2$.....	−0.0080	−0.0082	0.0002
0.1N KCl–0.1N $KC_2H_3O_2$.....	−0.0074	−0.0077	0.0003
0.2N KCl–0.2N KOH.....	+0.0170	+0.0168	0.0002
0.1N KCl–0.1N KOH.....	+0.0165	+0.0165	0.0000
0.2N KCl–0.2N KBr.....	+0.0004	+0.0004	0.0000
0.2N NaCl–0.2N NaOH.....	+0.0192 ± 0.0003	+0.0187	
0.1N KCl–0.1N HCl.....	−0.0286	−0.0286	0.0000

In the more general case limited chiefly by the condition that

all the ions shall have the same valency Planck (1890) deduced the equation:

$$E = \frac{RT}{wF} \ln \xi \tag{39}$$

where E is the contact difference of potential in volts and ξ is defined by the equation:

$$\frac{\xi U_2 - U_1}{V_2 - \xi V_1} = \frac{\ln \frac{c_2}{c_1} - \ln \xi}{\ln \frac{c_2}{c_1} + \ln \xi} \cdot \frac{\xi c_2 - c_1}{c_2 - \xi c_1} \tag{40}$$

c_1 is the sum of the concentrations of cations and anions in the more dilute solution and c_2 the sum in the more concentrated solution. w is the valency, R the gas constant, F the faraday, and

$$U_1 = u'c' + u''c'' + \ldots.$$
$$V_1 = v'c' + v''c'' + \ldots.$$

and U_2 and V_2 are similar sums for the second solution. The u′ and v′ symbols represent the ion mobilities and the c′ symbols the corresponding ion concentrations.

Besides the limitation noted above this equation is strictly applicable only to very dilute solutions where dissociation is complete and it was deduced for the condition of a sharp boundary such as is not realized in experimental work.

P. Henderson (1907, 1908) therefore considered the connecting boundary as a series of mixtures of the two solutions in all proportions and deduced a somewhat simpler equation which Cumming (1912) has modified by introducing the mobilities at the different concentrations used.

It is of course obvious that the equations given above and many others of like nature are inapplicable when the solutions placed in contact are of unknown composition or are very complex. Brønsted (1922) has proposed a novel method of approach which may prove to have some value, but as yet it is untried, and we are forced to get such comfort as we can find in a deduction from the above treatment which will be considered presently. But even in the simple cases where one or another of the equations

apply the experimenter must face the difficulty of maintaining experimentally the conditions for which they were set up. For instance Chanoz (1906) constructed the symmetrical arrangement:

$$\text{Electrode} \;\|\; \underset{A}{\text{MR}} \;|\; \underset{B}{\text{M'R'}} \;|\; \text{MR} \;\|\; \text{Electrode},$$

and then, by maintaining a more or less sharp boundary at A by renewal of the contact, and allowing diffusion to occur at B, he obtained very definite E. M. F.'s instead of the zero E. M. F. which the symmetrical arrangement demanded. This time effect has been noted by Weyl (1905) and has since been frequently reported, for instance, by Bjerrum (1911), Lewis and Rupert (1911), Cumming and Gilchrist (1913), Walpole (1914) and Fales and Vosburgh (1918).

Since the change of potential has been ascribed to the diffusion and mixing which alter the distribution of the contending, migrating ions, it has seemed to many that the effect could be made more uniform and conditions more reproducible if the solutions were brought into contact at a membrane. This would tend to prevent mixing. Sand or other material would also delay the mixing and the diffusion. Cumming and Gilchrist (1913) used a symmetrical chain such as that of Chanoz (see above), and found that when a membrane was introduced at A while ordinary contact was allowed at B the symmetry of the chain was destroyed. Prideaux (1914) also found a difference when the contact was made in the one case with, and in the other case without, a parchment membrane. On considering this case and others in which the constituents of the membrane may take part in the establishment of the potential, he came to the conclusion that there were phenomena concerned which made the application of the ordinary equations of dubious value. See also Beutner (1913).

Lewis, Brighton and Sebastian (1917) using Bjerrum's (1911) suggestion of a layer of sand in which to establish the liquid contact found that "at no time were reproducible results obtained nor results which remained constant to 0.0001 volt for more than a minute or two. The potential of the liquid junction first established was surprisingly high (0.030 volt) and fell rapidly with-

out reaching any definite limiting value." The liquids placed in contact in this experiment were 0.1M HCl and 0.1M KCl. These authors abandoned the sand method.

On the other hand Myers and Acree (1913) report satisfaction with Bjerrum's "Sandfüllung."

Other devices such as the use of a wick have been resorted to, but on the whole direct liquid contact is considered the best.

Recently Lamb and Larson (1920) have described the "flowing junction" which they find to be much more reproducible than the junctions usually made. They conclude "that a 'flowing' junction, obtained simply by having an upward current of the heavier electrolyte meet a downward current of the lighter electrolyte in a vertical tube at its point of union with a horizontal outflow tube, or by allowing the lighter electrolyte to flow constantly into a large volume of the heavier electrolyte, even with N solutions, gives potentials constant and reproducible to 0.01 of a millivolt." The device used by Lamb and Larson is illustrated in figure 15.

MacInnes and Yeh (1921) have improved the system of Lamb and Larson and have confirmed the principle that reproducible liquid junction potentials may be thus obtained, but they find most interesting effects with different rates of flow. Of particular importance is the observation that the reproducible potentials are not the highest that can be obtained.

It is encouraging to see experimental work of this type being done for those who are interested in the general applications of electrode measurements cannot escape the feeling that the experimental side of the problem has been too much neglected.

A most important contribution to experimental methods of handling liquid junction potential differences arose from the theory of Nernst that the potential is due to the unequal migration of ions. The table of velocities given on page 163 will show that if KCl is concerned no large potential can arise from the participation of its ions, because they move with about the same velocity. If such a salt be present in high concentration upon both or even one side of the interface, the electrostatic fields of its ions will dominate the situation, and, migrating at equal velocities, will tend to maintain zero junction potential difference. Bjerrum (1911) studied the potential differences developed when concentrated so-

lutions were thus employed and estimated the theoretical values with the aid of Planck's formula and with that of Henderson, which purports to take into account the effect of the destruction of a sharp boundary. He came to the conclusion that the use of a 3.5M KCl solution would not completely eliminate the potential against hydrochloric acid solutions but he suggested a more or less empirical extrapolation which would, he thought,

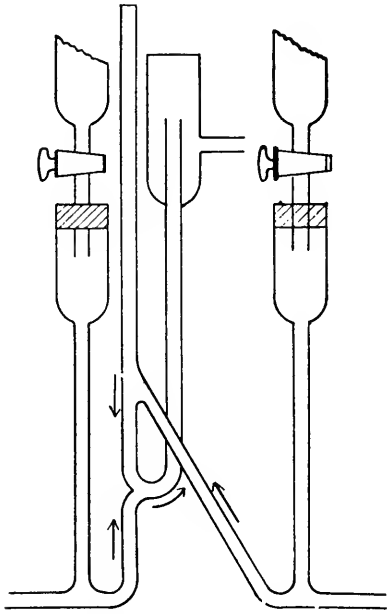

FIG. 15. LAMB AND LARSON'S DEVICE FOR THE FLOWING JUNCTION

give the proper correction. The correction is the difference in the E. M. F.'s of a chain found when first 3.5M KCl is used and then when 1.75M KCl is used to connect two electrodes.

More recently Fales and Vosburgh (1918) have made an extensive comparison of various chains, and with the aid of Planck's formula to give the order of magnitude of various contact potentials, thay have attempted to assign values which will lead to a general consistency. They concur with others in finding Planck's formula invalid in the assignment of accurate values to liquid junctions, such as:

"xM KCl − 1.0M HCl and xM KCl − 0.1M HCl where x ranges from 0.1 to 4.1 and the temperature is 25°C."

They conclude that "there is no contact potential difference at 25° between a saturated solution of potassium chloride (4.1M) and hydrochloric acid solutions ranging in concentrations from 0.1 molar to 1.0 molar," confirming the suggestion of Loomis and Acree (1911).

Because of the great detail concerned in the reasoning of Fales and Vosburgh it is impossible to briefly summarize their work, but before their conclusion can be considered valid it must be noted that they themselves point out that "in an electromotive force combination having a contact potential difference as one of its component electromotive forces, the diffusion across the liquid junction of the one liquid into the other brings about a decrease in the magnitude of the contact potential difference, and this decrease may amount to as much as one-tenth of the initial magnitude of the contact potential difference." This experimental uncertainty undoubtedly renders questionable the *comparability*, if not the precision of measurements by different experimenters. If so there may lurk in the data used by Fales and Vosburgh in their argument of adjustment to consistency an indefinite degree of incomparability.

Indeed the whole subject of contact potential is still in an unsatisfactory state. The experimental uncertainties which have been revealed have sometimes been overlooked in the calculation of important electrode values. Some of these values will be discussed in Chapter XIX.

In writing the components of a chain it is customary to designate the situation of a potential difference by a single line and the position of a potential difference which is to be left out of consideration by a double line. Thus

$$\mathrm{Pt\ H_2\ |\ N/10\ HCl\ |\ N/10\ KCl\ Hg_2Cl_2\ |\ Hg}$$

indicates that there are potential differences at the positions shown by the lines; while if the above chain is written as

$$\mathrm{Pt\ H_2\ |\ N/10\ HCl\ ||\ N/10\ KCl\ Hg_2Cl_2\ |Hg}$$

the double line indicates that the liquid junction potential difference is to be left out of consideration in formulating the E.M.F.

It now remains to determine if possible the order of magnitude of the contact differences of potential entering into chains used in the study of physiological solutions and the buffer solutions of the colorimetric method.

Since the concentrations of the hydrogen and the hydroxyl ions, which are the most mobile of all ions, are very low in most of these solutions, the contact potential difference may be expected to be much less than that found in hydrochloric acid solutions and similar solutions of high hydrogen or hydroxyl ion concentrations. It is the customary practice to employ saturated KCl in making the junction or to make the junction first with 3.5M, then with 1.75M KCl and extrapolate according to Bjerrum. The extrapolation so indicated generally amounts to only a few tenths of a millivolt, and in certain cases such as "standard acetate" to only 0.1 millivolt. Although such an extrapolation may be too low or too high its magnitude indicates that the error is not large. Furthermore there is found experimentally a drift in contact potential difference with time which is very much less than that found, for instance, at the junction sat. KCl—0.1M HCl. There can be no doubt that this is indicative of a low potential difference.

As pointed out by Clark and Lubs (1916), it is the difficulty in dealing with the contact potential of hydrochloric acid solutions that renders them unsuitable for routine standardization of hydrogen electrodes.

Practical conclusions reached by experimentation are:

1. For precise E. M. F. measurements combinations having small liquid junction differences of potential should be used as far as is practicable.

2. It should be recognized that the E. M. F. of a cell which derives part of its E. M. F. from a liquid junction potential difference varies with the time elapsing after the formation of the liquid junction. Consequently this time should become a part of the data to be recorded.

3. It is preferable that measurements of E. M. F. be made directly after the formation of or the renewal of the liquid junction.

4. Since the liquid junction potential difference may vary with the manner of its formation the effort should be made to effect this junction in a reproducible way.

5. Reproducible potential differences are given by the flowing junction in the cases so far tried.

6. Narrow or capillary tubes at the point of liquid junction should be avoided.

7. An apparatus which permits the renewal of a junction and its complete removal when cells are left set up together for some time is preferable to any device such as membranes to protect the diffusion of solutions into electrode spaces.

8. Membranes at the liquid junction are to be avoided.

9. Wherever permissible saturated KCl solution should form one side of a liquid junction.

10. When a concentrated KCl solution is used to make liquid junction it should be stated whether the Bjerrum extrapolation with the use of 3.5M and 1.75M KCl has been employed or whether *saturated* KCl was used without the Bjerrum extrapolation.

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CHAPTER XII

Hydrogen Electrodes and Electrode Vessels

For the most part the base of a hydrogen electrode is simply a piece of platinum foil or wire. If wire is used an end is fused into a glass tube carrying mercury to form a convenient means of making contact with the lead of the potentiometer circuit. The wire may then be wound upon a machine screw to give it a neat form. If foil is used a piece about 1 sq. cm. is first welded to a short piece of No. 30 B. S. gauge platinum wire by tapping the two smartly with the flat end of a punch while they are laid upon a flat hard surface in the white heat of a blast lamp. Next draw off a glass tube to a *thin*, blunt point and break away the capillary until the wire will enter. Slip the wire in until the foil touches the glass. Then, holding the wire with foil uppermost, rotate the tube with the junction in the tip of a fine flame. Let the glass fuse until a perfect seal is made and a little of the glass fuses to the edge of the foil. The steps are illustrated in figure 16. It is important to avoid a seal with too thin a glass junction, for such a seal will easily crack. It is likewise important

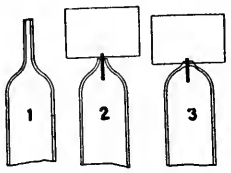

Fig. 16. Construction of Simple Electrode

to avoid too heavy a junction for proper annealing then becomes difficult. To anneal hold the electrode directly after its construction in a smoky flame and gradually remove to cooler and cooler parts of the flame. If a light but substantial junction is made with the edge of the *foil* the electrode will be rugged.

In place of the platinum foil gauze is sometimes successfully used. The advantage is a larger surface; but gauze will make a careful technician nervous over the problem of thoroughly cleaning the crevices.

It is sometimes assumed that complete equilibrium can be attained only when the hydrogen in the interior of the metal supporting the platinum black is in equilibrium with that on the surface. To reduce the time factor of this soaking-in process it is considered advantageous to use as the supporting metal a very thin film of platinum or iridium deposited upon glass. Doubtless the finest of such films could be deposited by holding the glass tangent to the Crookes' dark space of a vacuum discharge and spattering the metal on from electrodes under 5000 volts difference of potential. The method practiced is to burn the metal on from a volatile solvent. The recipe given by Westhaver(1905) is as follows: 0.3 gram iridium chloride moistened with concentrated HCl is dissolved in 1 cc. absolute alcohol saturated with boric acid. To this is added a mixture of 1 cc. Venetian turpentine and 2 cc. lavender oil. The glass after being dipped in this solution is rotated while drying to give an even deposit. It should then be very carefully dried to prevent blistering during the ignition. On gradually heating over an alcohol flame there is at last produced a very thin film of iridium. The process should be repeated until a good conducting film is obtained.

Gooch and Burdick (1912) have better success with a viscous mixture of pure chloroplatinic acid and glycerine. This is applied with an asbestos swab to the glass which has previously been heated to a temperature which will instantly volatilize the glycerine. The resulting film is heated until it adheres well to the glass.

Meillère (1920) gives the following recipe. 0.5 gram dry platinum chloride is triturated with 10 or 15 grams of essence of camomile. The mixture is thinned with about an equal volume of methyl alcohol.

If after some practice it is found that even deposits can be formed by one or another of the methods, the next difficulty met is in obtaining good adherence of the film to the glass. This must be done by heating sufficiently but at the same time there must be avoided a fusion of such extent that the continuity of the metallic film will be destroyed. Such a fusion will be more easily avoided and at the same time volatilization of impurities in the film will be made easier because of the higher temperature

permitted, if the glass support is made of a "hard" glass. However, in the selection of such a glass one with a temperature coefficient of expansion approximately equal to the platinum should be selected,—chiefly as a provision for the next step which will now be described.

The chief technical difficulty in the preparation of electrodes with the films described is in establishing the necessary electrical connection. An exposed platinum wire contact destroys the object in using the film. Ordinarily the electrode is made by first coating a bar of glass in the end of which there is sealed a platinum wire and then fusing this bar into the end of a glass tube so that the platinum contact is exposed within the tube where mercury contact may be made. Connection with the film is made by the film of metal that runs through the glass seal. It is less clumsy to seal the wire into the end of a glass tube, break off the wire flush with the glass, coat the tube with the film and then close over the exposed wire with a drop of molten glass.

A scheme which is said to partially accomplish the purpose of a thin film of supporting metal is to cover a platinum support with a gold-plate, gold being relatively impervious to hydrogen. It is doubtful whether this reason has much practical weight. However a gold-plate is of great advantage. If offers a surface upon which deposits of "black" adhere well. It forms a support easily dissolved by electrolysis in hydrochloric acid, thus providing an easy means of removing old deposits. And the character of the gold deposit gives an additional means of testing the cleanliness of the electrode prior to blackening.

For the gold plating of electrodes the following recipe may be used. Dissolve 0.7 gram gold chloride in 50 cc. water and precipitate the gold with ammonia water, taking care to avoid an excess. Filter, wash and dissolve immediately in a KCN solution consisting of 1.25 grams KCN in 100 cc. water. Boil till the solution is free from the odor of ammonia.

DEPOSITION OF "BLACK"

According to the work of earlier investigators and the consensus of opinion among more recent investigators there seems to be no difference under equilibrium conditions between coatings of platinum-, iridium- or palladium-black. No recent detailed data

are available however. Of the three, iridium is recommended by Lewis, Brighton and Sebastian because of its higher catalytic activity, and palladium by Clark and Lubs (1916) for use in the study of physiological solutions because of the relative ease with which one deposit may be removed before the deposition of the next in the frequent renewals which are often necessary. Palladium black is easily removed by electrolysis in HCl. Deposits of platinum or iridium may be removed by electrolysis in HCl solution, if they are deposited upon a gold plate.

One of the essentials for making good deposits is a very high degree of cleanliness of the electrode. A good test is the evenness with which bubbles of hydrogen escape from the surface during electrolysis. Another essential in the preparation of a good electrode is that the deposit of black metal be not only even but of proper thickness. The inclination is to make the deposit too thick, with the production of a sluggish electrode. To obtain evenness of deposit it is necessary to hold down the dimensions of the electrode, provide more than one lead, or modify the rate of deposit. With this much said there remains very little systematized information upon the composition of solutions and the current densities which are best for the deposition of the finely divided metal required.

For the deposition of platinum black Ellis (1916) uses a solution of pure chloroplatinic acid containing 1 per cent Pt. He cautions against the use of the lead acetate which has come down to us in recipes for the deposition of platinum black upon electrodes for conductivity measurements. For the deposition Ellis uses a small auxiliary electrode and a current large enough to liberate gas freely at both electrodes. He continues the deposition with five-minute reversals of current for two hours and obtains a very thick coating. The author prefers an adherent, even, thin deposit sufficient to just cover the glint of metal beneath. In comparison of one against another in the same solution such thin deposits are found to agree within 0.02 millivolt. They may be deposited within a minute from the solutions used by the author.

For the deposition of iridium Lewis, Brighton and Sebastian (1917) make the gold or gold-plated electrode the cathode in a 5 per cent solution of iridium chloride. "The best results were obtained with a very small current running for from twelve to

twenty-four hours. Too large a current gives a deposit which appears more like platinum black and which is easily rubbed off."

The author has used deposits of platinum, iridium and palladium upon platinum, upon gold-plated platinum and upon "rhotanium" alloy. Acidified (HCl) 3 per cent solutions of the chlorides of each metal are used without much attention to the exact strength. The current from a four-volt storage battery is allowed to produce a vigorous evolution of gas. The electrode is plunged, immediately after the deposition, into a dilute sulfuric acid solution and electrolyzed. It is required that the bubbles of hydrogen then escaping come off evenly, that the electrode be evenly covered with the deposit in thickness sufficient to cover the glint of polished metal, and that the deposit shall adhere under a vigorous stream of water. No electrode is ever subjected to blast lamp treatment as is sometimes recommended. Instead, renewals are made by removing the old deposit by electrolysis in HCl solution, and, if any defect whatsoever develops to prevent a good redeposition after such electrolysis, the electrode is retired from duty.

It must be admitted that the above description is loose. This is because the rush of experimental application has prevented a detailed examination of conditions, and experience has taught details difficult to formulate in exact language. No detailed descriptions have been found in the literature and those that are found are quite inadequate to account for the varied deposits sometimes formed. One item which it would be interesting to investigate is the influence of the hydrogen ion concentration of the solution upon the character of the deposit. Since there is a simultaneous deposit of metal and hydrogen and, since the character of the platinum, palladium or iridium black is undoubtedly due to the vigor of the hydrogen evolution, it is evident that the pH of the solution constitutes a part of the conditions.

It may be said however, that ordinarily there is little difficulty in obtaining an active deposit if the metal concentration is maintained as the solution is used, if electrodes are kept thoroughly clean and if the solutions are kept free from even those impurities which collect as a film upon exposed solutions. To remove these films suck them off with a clean tube attached to a filter pump.

The system used by the author for deposition of "black" is

as follows. A row of small vessels, such as weighing bottles about 2 cm. diameter and 5 cm. deep are fitted with electrodes. These electrodes are all attached through binding posts mounted on a wooden rail. These in turn are connected to one pole of a double-pole, double-throw switch. The opposite pole is connected with a flexible lead tipped with platinum. This lead is used to connect with the electrodes to be treated. The middle connections of the double-throw switch are connected with a 4-volt storage battery. The other connections are cross-wired. One of the vessels is filled with hydrochloric acid made by a one-to-one dilution of ordinary 37 per cent acid. This is used to dissolve previous deposits with the aid of electrolysis (switch reversed, treated electrode +). Another vessel is filled with 10 per cent sulfuric acid for preliminary direct and counter-electrolysis in testing the cleanliness of the electrode. Another vessel is filled with the platinum, palladium or iridium chloride solution. When using palladium so-called reagent palladium is used as + electrode and this is removed from the solution when not in use. After deposition of the black the electrode under treatment is quickly placed under a vigorous stream of water and then electrolyzed in a another vessel of freshly prepared ten per cent sulfuric acid until thoroughly charged with hydrogen.

When used with inorganic solutions which undergo no decomposition electrodes may often be used repeatedly, provided they are kept clean and *not allowed to dry.* When there is any sign or suspicion of an electrode becoming clogged, poisoned, worn, dry or in any way injured, there should be not the slightest hesitation in reblackening or even rejecting it. It is therefore not good practice to so tie up a particular electrode with an expensive stopper or vessel that there will be hesitation in rejecting it.

HYDROGEN ELECTRODE VESSELS

So many types of vessel have been published that it is difficult to do justice to the advantages of each. The selection must depend in some instances upon the material to be handled, but in any case there are a few principles which it is hoped will be made clear by a discussion of a few of the more widely used vessels.

The general method of operation is to partially or wholly im-

merse the electrode in the solution to be measured and then to bubble hydrogen through the vessel till constant potential is attained. The vessel described by Lewis, Brighton and Sebastian (1917) and illustrated in figure 17 is representative of the general type of vessel used for what may be called the classic mode of operation. The following is the quoted description of this vessel:

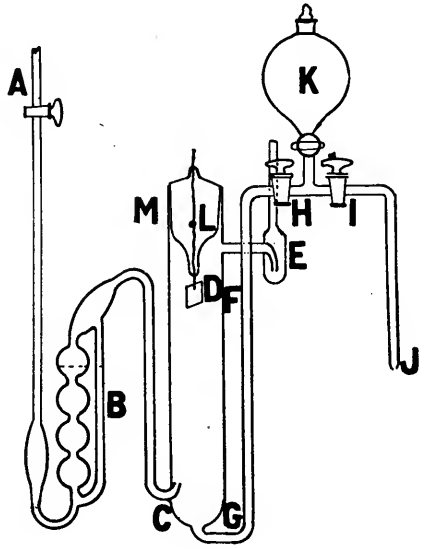

FIG. 17. HYDROGEN ELECTRODE VESSEL OF LEWIS, BRIGHTON AND SEBASTIAN

Hydrogen from the generator enters at A, and is washed in the bubbler B with the same solution that is contained in the electrode vessel. This efficient bubbling apparatus saturates the gas with water vapor, so that the current of hydrogen may run for a long period of time without changing the composition of the solution in the main vessel. The gas rises from the tip C, saturating and stirring the whole liquid from G to F, and leaves the apparatus through the small trap E, which also contains a small amount of the same solution. The platinum wire attached to the electrode D is sealed by lead glass into the ground glass stopper M. L is a joint made by fusing together the end of the platinum wire and the connecting wire of copper. The surface of the solution stands at the height F so that the iridium electrode is about one-half immersed. The apparatus from F through G, H, I to J is filled with the solution. With the form of construction shown it is an easy matter to fill the tube without leaving any bubbles

of air. The reservoir K filled with the same solution serves to rinse out the tube I, J from time to time. The whole apparatus may be mounted upon a transite board, or for the sake of greater mobility, may be held in a clamp, the several parts being rigidly attached to one another to avoid accidental breakage. The whole is immersed in the thermostat about to the point L.

The tube J dips into an open tube through which communication is made to other electrode vessels. This connecting tube may be filled with the same solution as is contained in the hydrogen electrode vessel or with any other solution which is desired. All measurements with acids are made with one of the stopcocks H, I, closed. These stopcocks are not greased and there is a film of acid in the closed stopcock which suffices to carry the current during measurement. In order to make sure that no liquid potential is accidentally established, the second stopcock may be closed up and the first opened. No difference of potential in acid solution has ever been observed during this procedure (but this is not true for solutions of salt and alkalies). If it is desired that both stopcocks be open, the same liquid that is in the electrode vessel is placed in the connecting tube at J and the stopcocks H and I are opened after the current of hydrogen has been cut off by the stopcock A, and the opening of the trap E has been closed.

If hydrogen enters the cell at the rate of one or two bubbles per minute several hours are required for the saturation of the solution and for the removal of air. After this time the potential is absolutely independent of the rate of flow of hydrogen and the generator may be entirely cut off for many hours without any change.

For some biochemical studies such a vessel is unsuitable. It is sometimes absolutely essential that equilibrium potentials be established rapidly. The necessity is perfectly apparent when one is dealing with an actively fermenting culture. It is not always so apparent when dealing with other solutions, but it is suspected that absolutely complete equilibrium is never attained in some complex biochemical solutions and that we have to depend upon speeding up the reaction between hydrogen and hydrogen ions till a virtual equilibrium point is attained (see Chapter XVII).

It was shown by Michaelis and Rona (1909) that a fairly constant E. M. F. is quickly attained, even in blood, if the platinized electrode, previously saturated with hydrogen, is allowed to merely touch the surface of the solution. This is probably due, as suggested by Hasselbalch (1913) and again by Konikoff (1913), to a rather sharply localized equilibrium at the point of contact. Reductions and gas interchanges having taken place within the small volume at the point of contact, diffusion from the remaining body of the solution is hindered by the density of the surface layer with which alone the electrode comes in contact.

In exploring new fluids it appeared hazardous to the writer to rely upon such a device, which appears to take advantage of only a localized and hence a pseudo-equilibrium, and which makes no allowance for a possible difference between the solution and surface film in the activity of the hydrogen ions. Hasselbalch's (1911) principle seemed therefore to be more suitable.

Hasselbalch found that a very rapid attainment of a constant potential can be obtained by shaking the electrode vessel. Under these conditions there should be not only a more rapid interchange of gas between the solution, the gaseous hydrogen, and the electrode, an interchange whose rapidity Dolezalek (1899) and Bose (1900) consider necessary, but the combined or molecular oxygen, or its equivalent, in the whole solution should be more rapidly brought into contact with the electrode and there reduced. Furthermore, by periodically exposing the electrode the hydrogen is required to penetrate only a thin film of liquid before it is absorbed by the platinum black. The electrode should therefore act more rapidly as a hydrogen carrier. For these reasons a true equilibrium embracing the whole solution should be rapidly obtained with the shaking electrode; and indeed a constant potential is soon reached.

Eggert (1914–1915) in Nernst's laboratory made a study of the rapidity of reduction by hydrogen electrodes in which he compared the effect of alternate immersion and exposure to the hydrogen atmosphere with the effect of continued immersion. In the reduction of metal salt solutions such as ferric salts he obtained a much greater velocity of reduction when the electrode was periodically removed from the liquid carrying a thin film of solution to be exposed to the hydrogen. The maximum velocity was proportional to the platinum surface and the time of contact with the gas. It was independent of the number of times per minute the electrode was raised and lowered. As the reaction neared completion the decrease in velocity of reaction became exponential.

Making use of the principles brought out in the preceding discussion and also certain suggestions noted in the chapter on liquid junction potentials Clark (1915) designed a vessel which appears to have found favor for general use. A working drawing of this vessel is shown in figure 18. If solutions more viscous than fresh

milk are to be used, the bores of the inlet and outlet tubes should be made larger. If only very small quantities of the solutions to be tested are available, the dimensions of the vessel may be reduced. In figure 19 is a diagrammatic sketch of the complete system now in use by the author for ordinary work.

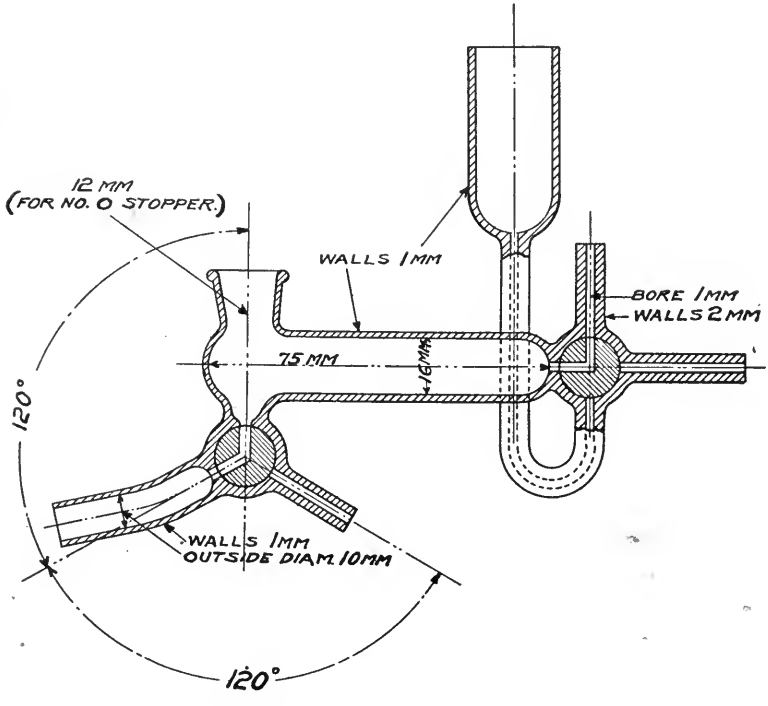

FIG. 18. A HYDROGEN ELECTRODE VESSEL (CLARK, 1915)

Notes. In submitting this working drawing to a glass blower it shall be specified that: (1) Cocks shall be joined to chamber with a neat and wide flare that shall not trap liquid. (2) Cocks shall be ground to hold high vacuum. (3) Bores of cock keys shall meet outlets with precision. (4) The handles of keys shall be marked with colored glass to show positions of bores. (5) The handles of both keys shall be on the same side (front of drawing). (6) Vessel shall be carefully annealed. (7) Opening for no. 0 rubber stopper shall be smooth and shall have standard taper of the standard no. 0 stopper. (8) Dimensions as given shall be followed as closely as possible. (9) No chipped keys or violation of the above specifications shall be accepted.

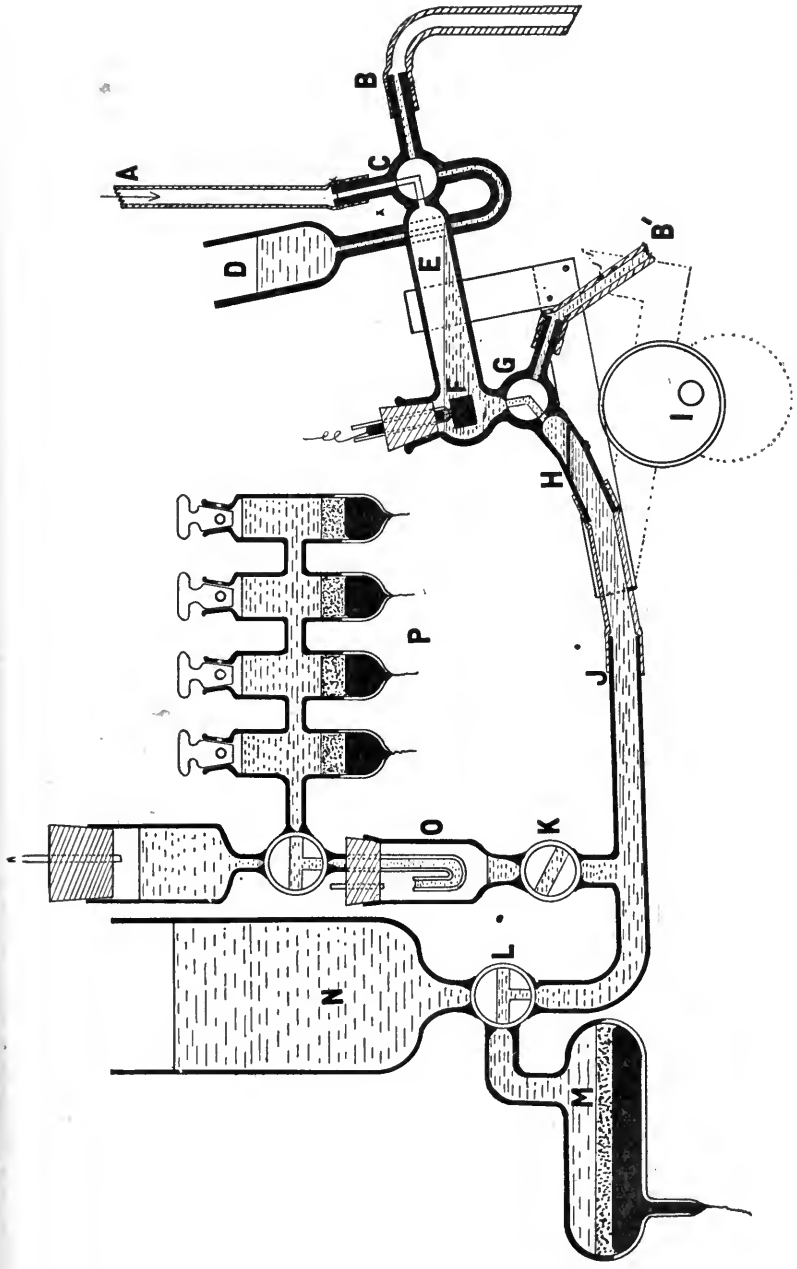

FIG. 19. A SYSTEM FOR THE MEASUREMENT OF HYDROGEN ELECTRODE POTENTIALS

The electrode vessel is mounted in a clamp pivoted behind the rubber connection between J and H. This clamp runs in a groove of the eccentric I, the rotation of which rocks the vessel. In the manipulation of the vessel, the purpose is, first, to bring every portion of the solution into intimate contact with the electrode F and the hydrogen atmosphere, to make use of the principle of alternate exposure and immersion of electrode and then, when equilibrium is attained, to draw the solution into contact with concentrated KCl solution and form a wide contact at H in a reproducible manner. The E. M. F. is measured directly after the formation of this liquid junction.

The vessel is first flooded with an abundance of hydrogen by filling the vessel as full as possible with water, displacing this with the hydrogen, and then flushing with successive charges of hydrogen from the backed-up generator. Water or solution is run into the vessel from the reservoir D which can be emptied through the drain B by the proper turning of the cock C. Solution or hydrogen displaced from the vessel is drained off at B′. These drains when they leave the electrical shielding (see p. 231) should hang free of any laboratory drain.

With the vessel rocked back to its lowest position the solution to be tested is run in from D (after a preliminary and thorough rinsing of the vessel with the solution) until the chamber E is about half full. Cock G is closed and cock C is turned so as to permit a constant pressure of hydrogen from A to bear upon the solution. For very careful work it is well to bubble hydrogen through the solution. The rocking is then commenced and is continued until experience shows that equilibrium is attained with the solution of the type under examination. The eccentric I should give the vessel an excursion which will alternately completely immerse the electrode F and expose it all to the hydrogen atmosphere. The rate of rocking may be adjusted to obtain the maximum mixing effect without churning.

To establish the liquid junction the rubber tube between J and H is pinched while G is turned to allow KCl solution to escape at B′. Then a turn of G and the release of the pinch draws the solution down through the cock to form a broad mixed junction at H. For a new junction the old is flushed away with fresh KCl from the reservoir N by properly setting cock L.

With the closed form of calomel electrode, M, shown in the figure no closed stopcocks need be interposed between the terminals of the chain. With the customary calomel electrode vessel it is necessary to use a closed cock somewhere and since this must be left ungreased it is well to have it a special cock[1] at J.

If a tube be led out from J and branched, several hydrogen electrode vessels may be joined into the system. At all events it is well to work with two vessels in parallel so that one may be flushing with hydrogen while the other is shaking.

The electrode F is supported in a sulfur-free rubber stopper. A glass stopper may be ground into place but is seldom of any advantage and may prove to be a mistake. In the first place it is advisable to be free with electrodes and to instantly reject any which fail to receive a proper coating of metal. The inclination to do this is less if it entails the rejection of a carefully ground stopper. Unless the stopper is accurately ground into place it is worthless. Furthermore it is very difficult to so grind a glass stopper that there will be left no capillary space to trap liquid. A rubber stopper can be forced into place without leaving such a space. The rapidity with which measurements are usually taken makes it improbable that a rubber stopper, if made sulfur free, can have any appreciable effect. If the rubber must be protected a coating of paraffine will do.

The calomel electrode M is of the saturated type so that no particular care need be taken to protect it from the saturated KCl used in making junctions. This is the working standard for the accurate standardization of which there is held in reserve the battery of accurately made, tenth-normal, calomel electrodes P. This battery may be connected with the system at any time by making liquid connection at O and opening K. After a measurement the liquid junction is eliminated, the space rinsed with the tenth normal KCl, and liquid contact *left broken.*

The design of this system is obviously for an air bath. The necessity of raising cocks out of an oil bath would not permit such direct connections as are here shown.

[1] To make an easily turning cock out of which KCl will not creep, grease the narrow part of the socket and the wide part of the key. When the key is replaced there will be two bands of lubricant on which the key will ride with an uncontaminated zone between for the film of KCl solution.

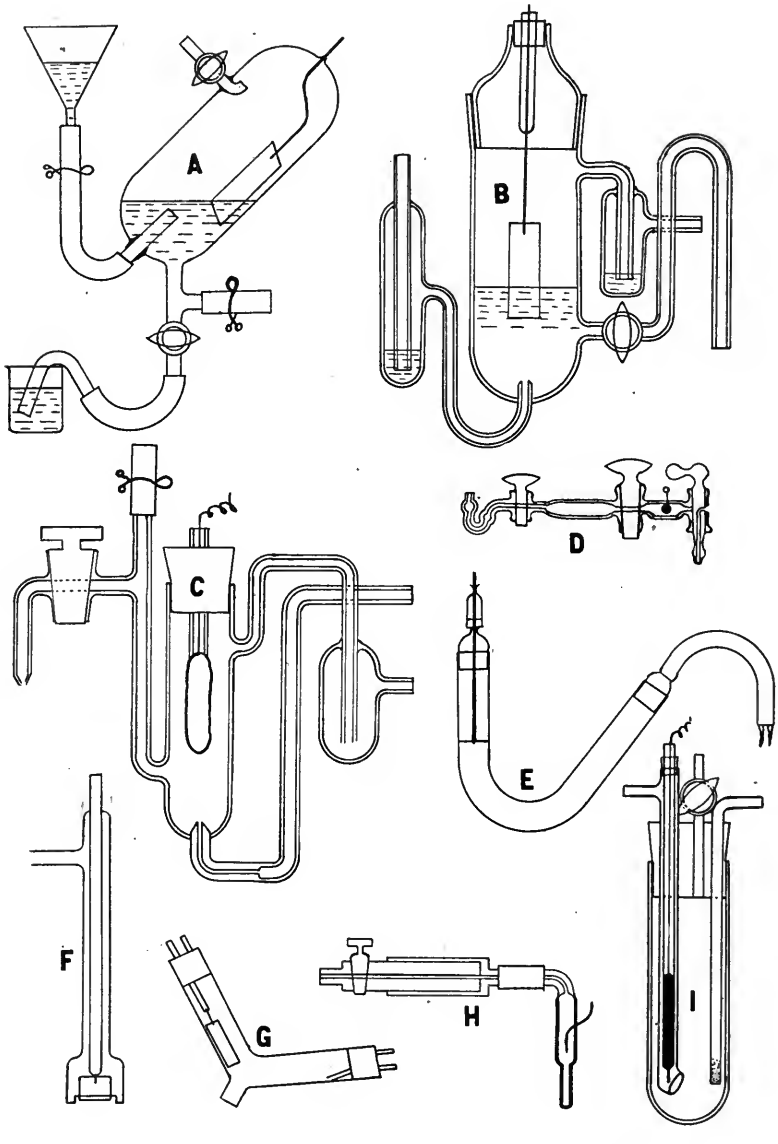

FIG. 20. TYPES OF HYDROGEN ELECTRODE VESSELS

In figure 20 are shown several other designs of electrode vessels. A is one of the original Hasselbalch vessels which have since been modified for the use of replaceable electrodes. B (Sørensen), (Ellis) and C (Walpole), are operated in a manner similar to the vessel shown in figure 18. Walpole's vessel was made of silica and the electrode was of platinum film as described on page 174. D (McClendon and Magoon) was designed for determinations with small quantities of blood. E (Michaelis), employs a stationary hydrogen atmosphere and a wick connection for the liquid junction. G (Long) is a simple device which the designer thought applied the essential principles of Clark's vessel. Barendrecht's vessel, H, is designed for immersion in an open beaker for estimations during titrations. It is similar to a design of Walpole's (1914), but is provided with a plunger the working of which permits the rinsing of the bulb and the precise adjustment of the level of the liquid. Another immersion electrode is Hildebrand's, F, the successful operation of which depends upon a vigorous stream of hydrogen, which, on escaping from the bell surges the solution about the electrode. A modification which provides better protection of the electrode from oxygen is Bunker's design, I.

At this point it may be of interest to note that Wilke (1913) attempted to make a hydrogen electrode by using a thin tube of palladium on the interior of which hydrogen was maintained under pressure. One of the difficulties with such an electrode is the estimation of the hydrogen pressure at the solution-electrode interface. Wilke's idea has never been developed to a practical point so far as we know, but it is worthy of study as an immersion electrode for industrial use.

For titrations where exact control of liquid junction potential differences is of relatively less importance than control of wastage of the material titrated, the system illustrated in figure 21 is useful. Titrations are carried on in the Erlenmeyer flask which is held in place by the plate P. The arm carrying the spring may be attached to the support at A in a variety of ways. It may be bolted, riveted or screwed; but should be made with a "running fit" so that while held firmly, it may be turned to permit removal of the Erlenmeyer. The plate F should be rigidly attached to the support at B. In this plate there is turned a hole tapered to receive snugly the rubber stopper which holds

the various attachments. If this hole is left rough from the lathe tool the stopper will be held very firmly after the various glass

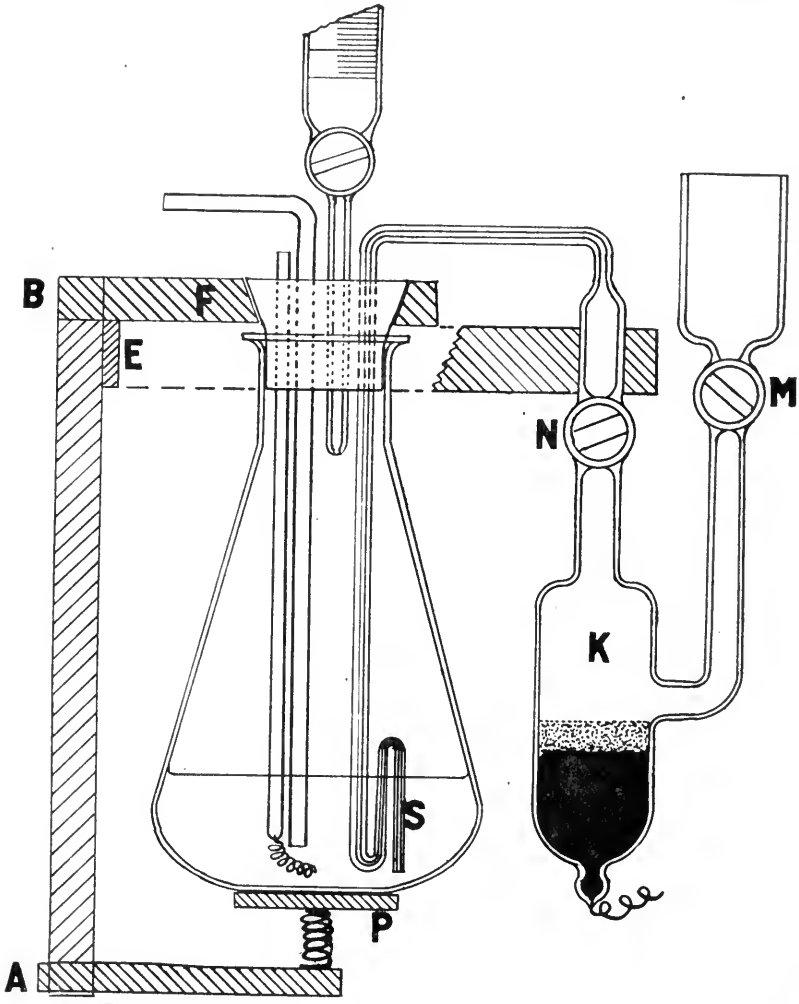

FIG 21. A HYDROGEN ELECTRODE VESSEL SUITABLE FOR TITRATIONS

tubes have been forced into place. The support has been illustrated in the drawing as if it were at the left. As a matter of fact it is behind the vessel, and carries at E a bar which supports the

calomel cell K. The supporting system is illustrated roughly for there are various constructions which may be used. In the author's apparatus A is a screw connection and the junctions at B and E are riveted and soldered.

It is of course essential that the solution be shaken after each step of the titration. If the support is clamped to a somewhat flexible rod the whole system may be shaken. Otherwise the glass vessel should be protected from the metal of the supporting plate by an inset of asbestos wood and then, if the spring is not too stiff, the vessel alone may be swirled. During a titration cock M is kept closed and N is left open. If the system is sufficiently rigid, if care is used in the operation of the cocks and if serious temperature changes are avoided very little of the solution will be drawn into the capillary S and very little of the KCl will run or diffuse into the solution.

A wire form of electrode will withstand shaking and possible scraping better than a foil electrode.

Hydrogen is delivered beneath the surface of the liquid. At the first flushing an abundance of hydrogen is used; later but little is necessary. The hydrogen escapes through a tube not shown and should be run through a trap having a *shallow* layer of water.

The burette tip shown in the figure is lengthened by a piece of capillary tubing.

If the hydrogen be replaced by purified nitrogen and if the platinized electrode be replaced by a gold-plated electrode this vessel does very well for oxidation-reduction titrations. In this case the nitrogen is delivered above the solution and not below the surface.

In some cases a preliminary reduction of a solution may be accomplished by making the solution, in the presence of hydrogen, travel down a long spiral of platinized wire. The spiral is made by winding no. 24 copper wire closely upon a rod, mounting it with a spread of the turns just sufficient to hold together descending drops, plating with gold and then platinizing. Liquid delivered slowly at the top of the spiral will be broken into drops which in the descent of the spiral are thoroughly stirred. The reduced solution is brought into contact with an electrode in a constricted part of the enclosing tube and is then delivered to a continuous-flow liquid junction such as that described by Lamb

and Larson or MacInnes (see page 168). The hydrogen by suitable devices may be given the carbon-dioxid partial pressure of the tested solution. Such a scheme is useful only in dealing with continuous treatment processes where abundance of material is available.

Keller (1922) has described a hydrogen electrode with a replaceable disk of platinum gauze. This is held by a cap to a hard rubber support which contains a portable calomel electrode. The system is rugged and may be used as an immersion chain for determining the pH values of liquids in commercial processes.

In conclusion it may be said that with ordinary care almost any simple combination of electrode and electrode vessel will give fairly good results. On the other hand it is often necessary not only to provide against continuous loss of CO_2 from biological solutions but also to arrange for rapid attainment of equilibrium. Since electrode measurements are often the last resort, since one can easily be misled by pseudo-equilibria and since attention to a few simple details of construction and operation frequently increases very greatly the speed of experimentation, the "simplicity" of certain designs is sometimes more apparent than real.

However it would be invidious to select any particular design for criticism, the more so because none yet published is perfectly adapted to *all* purposes. Those described are therefore to be considered as illustrations from which the reader may select items or suggestions to incorporate in his own design.

SUPPLEMENTARY REFERENCES

Bailey (1920), Baker-Van Slyke (1918), Barendrecht (1915), Bose (1900), Bunker (1920), Dolezalek (1899), Eggert (1914–1915), Ellis (1916), Gooch-Burdick (1912), Clark (1915), Cullen (1922), Hasselbalch (1910–1913), Hastings (1921), Hildebrand (1913), Hudig-Sturm (1919), Konikoff (1913), Lewis-Brighton-Sebastian (1917), Linhart (1919), Long (1916), Loomis-Acree (1911), Maloney (1921), McClendon (1915, 1916, 1918), McClendon-Magoon (1916), Michaelis (1910, 1911, 1914), Michaelis-Rona (1909), Myers-Acree (1913), Peters (1914), Rudnick (1921), Sand-Law (1911), Sørensen (1909), Sturm (1918), Treadwell-Weiss (1920), Walpole (1913, 1914), Westhaver (1905), Wilke (1913).

CHAPTER XIII

Calomel Electrodes

Unless otherwise specified the calomel electrode is an electrode in which mercury and calomel are overlaid with a definite concentration of *potassium chloride*. For particular purposes HCl calomel electrodes or those containing some other chloride are used.

The general type of construction is shown by A, fig. 23. A layer of very pure mercury is covered with a layer of very pure calomel and over all is a solution of a definite concentration of KCl saturated with calomel.

The difference of potential between mercury and solution is determined primarily by the concentration of the mercurous ions supplied from the calomel. But, since there is equilibrium between the calomel, the mercurous ions and the chlorine ions, the concentration of the mercurous ions is determined by the chlorine ion content furnished chiefly by the KCl. One of three concentrations of KCl is usually employed—either 0.1 molecular, 1.0 molecular or saturated KCl. These are ordinarily referred to as the "tenth normal-," "normal-" or "saturated calomel electrodes."

The mercury used in the preparation of these electrodes or "half-cells" should be the purest obtainable. In Chapter XV methods of purification are described. Sufficient mercury should be used to cover the platinum contact deeply enough to prevent solution reaching this contact on accidental shaking.

More portable half-cells are made by amalgamating a platinum wire or foil. This is done by electrolyzing a solution of mercurous nitrate, the wire being the negative pole. Provision s then made for keeping a paste of calomel about this wire.

Some success has been attained with the use of the better grades of calomel supplied on the market but the risk is so great hat it is best to prepare this material in the laboratory. A hemical and an electrolytic method will be described.

The chemical preparation of calomel. Carefully redistill the best obtainable grade of nitric acid. Dilute this slightly and with it lissolve some of the mercury prepared as described in Chapter

XV, always maintaining a large excess of mercury. Throw the solution into a large amount of distilled water making sure that the resulting solution is distinctly acid. Now, having distilled pure hydrochloric acid from a 20 per cent solution and taken the middle portion of the distillate, dilute and add it slowly to the mercurous nitrate solution with constant stirring. When the precipitate has collected, decant and treat with repeated quantities of pure distilled water (preferably conductivity water). The calomel is sometimes washed with suction upon a Buchner funnel but if due regard be taken for the inefficiency of washing by decantation it is preferable to wash repeatedly by decantation since there is thereby obtained a more even-grained calomel. Throughout the process there should be present some free mercury.

Electrolytic preparation of calomel. Doubtless the better preparation of calomel is formed by electrolysis according to the method of Lipscomb and Hulett (1916). This is carried out in the same way that the mercurous sulfate for Weston cells is formed. For the preparation of mercurous sulfate Wolff and Waters (1907) employ the apparatus shown in figure 22. An improvised apparatus may be made of a glass tube with paddles, platinum wire electrode and mercury contact and with two spools for bearing and pulley. In place of the sulfuric acid there is used normal hydrochloric acid. A direct current (from a four-volt storage battery) must be used. The alternating current sometimes used in the preparation of mercurous sulfate does not seem to work in the preparation of calomel according to some preliminary experiments which Mr. McKelvy and Mr. Shoemaker of the Bureau of Standards kindly made for the writer. During the electrolysis the calomel formed at the mercury surface should be scraped off by the paddles c and c (fig. 22). The calomel formed by this process is heavily laden with finely divided mercury.

Calomel formed by either the chemical or the electrolytic process should be shaken with repeated changes of the KCl solution to be used in the half-cell before the calomel is placed in such a cell.

The variations in the potentials of calomel electrodes have been the subject of numerous investigations. Richards (1897) ascribed it partly to the formation of mercuric chloride. Compare Richards and Archibald (1902). Sauer (1904) on the other hand con-

cluded that this had little to do with the inconstancy. Arguing upon the well known fact that the solubility of slightly soluble material is influenced by the size of the grains in the solid phase, Sauer thought to try the effect of varying the grain size of the calomel as well as the effect of the presence of finely divided mercury.

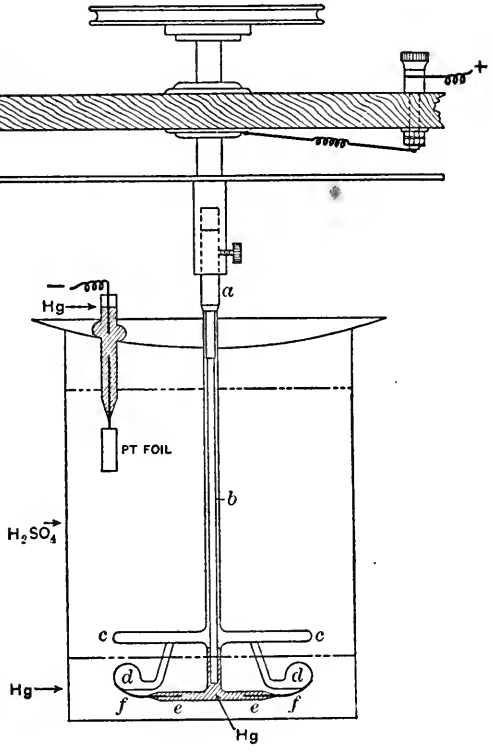

FIG. 22. WOLFF AND WATERS' APPARATUS FOR THE ELECTROLYTIC PREPARATION OF MERCUROUS SULFATE OR OF CALOMEL

With cells made up with various combinations he found the following comparisons:

Hg^- (fine)	calomel (coarse)	against	calomel (fine)	Hg^+ (coarse)	= − 0.00287 volt
Hg^- (fine)	calomel (coarse)	against	calomel (coarse)	Hg^+ (coarse)	= − 0.00037 volt
Hg^- (coarse)	calomel (coarse)	against	calomel (fine)	Hg^+ (coarse)	= − 0.0025 volt

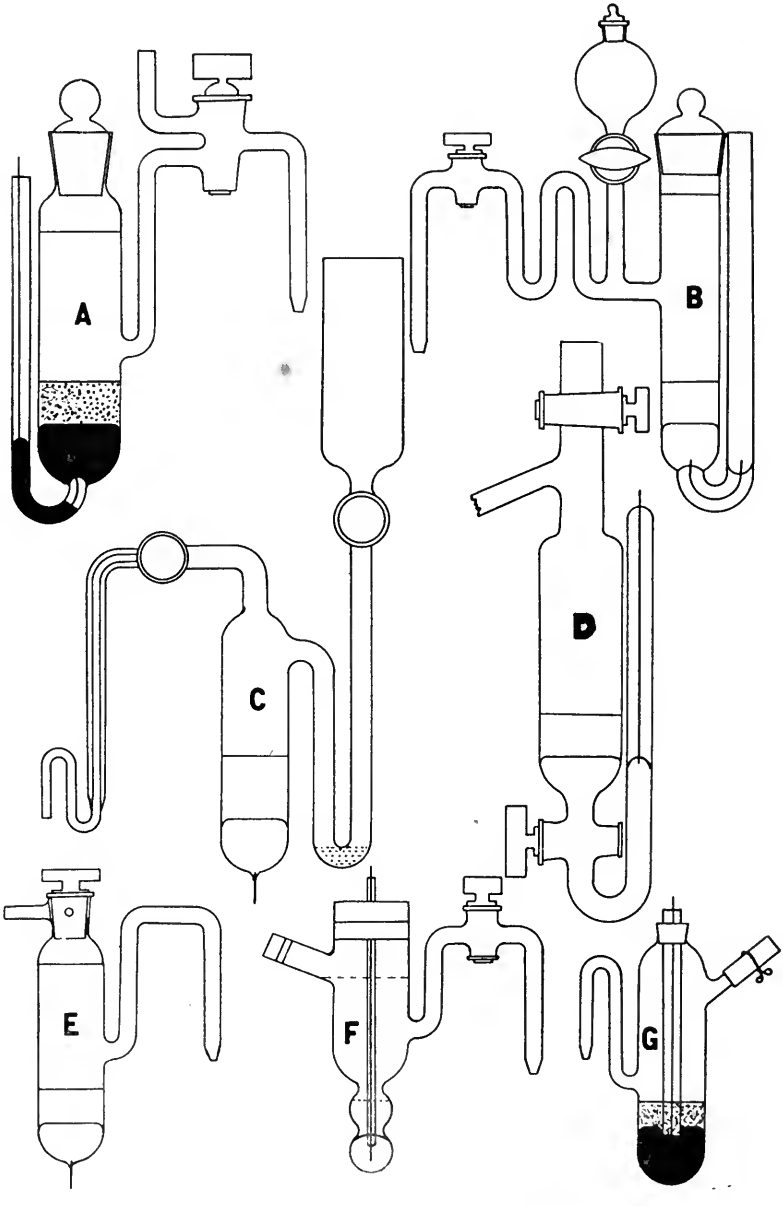

FIG. 23. TYPES OF CALOMEL ELECTRODE VESSELS

Lewis and Sargent (1909) state that they do not confirm Sauer in regard to the effect of the finely divided mercury but that they do confirm him in regard to the state of the calomel. These authors and others recommend that grinding the calomel with mercury to form a paste be avoided as this tends to make an uneven grain. It is better to shake the mercury and the calomel together but this is unnecessary if electrolytic calomel is used.

Here and there in the literature we find various other suggestions such as the elimination of oxygen from the cell; but there seems to be no very substantial agreement in regard to this and several other matters as there is no substantial agreement in the preference of one concentration of KCl over another. By the use of carefully prepared material and the selection of the better agreeing members of a series, calomel electrodes may be reproduced to agree within 0.1 millivolt or better; but it has not yet been established whether or not this represents the order of agreement among electrodes made in different laboratories. Furthermore there still remains the question of the effect of minor disturbances. There is no question that "true" values are not to be expected until all parts of the system are in equilibrium and that a preliminary shaking such as Ellis uses will hasten the attainment of equilibrium. On the other hand a disturbance which will alter the surface structure of the mercury exposed may produce a slight temporary shift in the potential difference. The subject remains for systematic investigation.

The most extensive investigation of unsaturated calomel electrodes was made by Acree and his students (Myers and Acree, Loomis and Acree), but how far the reproducibility which they attained by short circuiting the differences of potential is representative of the general reproducibility of such electrodes is not yet established.

In figure 23 are shown several calomel electrode vessels each with a feature that may be adapted to a particular requirement. Valpole's (1914) vessel, A, is providedwith a contact that leads out of the thermostat liquid and with a three-way cock for flushing away contaminated KCl. A more elaborate provision for the protection of the KCl of the electrode is shown in the vessel of Lewis, Brighton and Sebastian (1917), B. A form useful as a saturated calomel electrode in titrations is shown at C. Fresh KCl

passes through the U-tube to take the temperature of the bath and to become saturated with calomel shown at the bottom of this U-tube. D is Ellis' (1916) vessel, which in the particular form shown was designed to be sealed directly to the remainder of the apparatus used. A valuable feature is the manner of making electrical contact. Instead of the customary sealed-in platinum wire Ellis uses a mercury column. On closing the cocks the vessel may be shaken thoroughly to establish equilibrium. This feature has not been generally practiced. Vessel E is a simple form useful for the occasional comparison electrode. It may be made by sealing the cock of an ordinary absorption tube to a test tube and adding the side arm. F is the vessel of Fales and Vosburgh (1918) with electric contact made as in the familiar Ostwald vessel (G).

In adding new KCl solution to a vessel it must be borne in mind that the solution should be saturated with calomel before equilibrium can be expected. It is well therefore to have in reserve a quantity of carefully prepared solution saturated with calomel.

Lewis, Brighton and Sebastian (1917) state that certain grades of commercial KCl are pure enough to be used in the preparation of KCl solutions for the calomel electrode while other samples "contain an unknown impurity which has a surprisingly large effect upon the E. M. F. and which can only be eliminated by several recrystallizations." It is therefore obvious that the only safe procedure, in lieu of careful testing by actual construction of electrodes from different material, is to put the best available KCl through several recrystallizations.

Acree has called attention to the possible concentration of the KCl solution by the evaporation of water and its condensation on the walls of vessels unequally heated in thermostats.

The values assigned to the potential differences at the several calomel electrodes at different temperatures vary. A judicious selection will wait upon the consideration of several important matters. Some of the more important of these will be presented in Chapter XIX. At this point however we may recount without comment some of the more frequently used values which the reader may choose to use.

Clark and Lubs (1916) give the following compilation of Bjerrum's values and those of Sørensen and Koefoed published by Sørensen (1912):

TABLE 42

AUTHOR	TEMPERATURE	POTENTIAL DIFFERENCE BETWEEN NORMAL HYDROGEN ELECTRODE AND N/10 CALOMEL ELECTRODE WHEN HYDROGEN PRESSURE IS	
		One atmosphere less vapor pressure	One atmosphere
	°C.	*volts*	*volts*
Bjerrum	0	0.3366	0.3367
Sørensen and Koefoed	18	0.3377	0.3380
	20	0.3375	0.3378
Bjerrum	25	0.3367	0.3371
Sørensen and Koefoed	30	0.3364	0.3370
	40	0.3349	0.3359
	50	0.3326	0.3344
	60	0.3290	0.3321
	75	0.3243	0.3315

In the report of the "Potential Commission" of the Bunsen-Gesellschaft (Abegg, Auerbach and Luther, 1911) the normal hydrogen electrode standard of difference of potential was adopted. This of course is only incidental except as temperature coefficients enter. The differences of potential between the normal hydrogen electrode and the tenth-normal and normal KCl calomel electrodes were given as 0.337 and 0.284–0.283 respectively. Auerbach (1912) in a review of this report called attention to the smaller temperature coefficient of the potential difference at the tenth-normal calomel electrode when referred to the normal hydrogen electrode (as having zero potential difference at all temperatures) and suggested that the tenth-normal electrode be taken as the working standard with the value 0.3370 between 20°C. and 30°C.

Loomis and Acree (1911) present a choice of values for the tenth-normal calomel electrode at 25°C. referred to the normal hydrogen electrode. The choice depends upon the ionization ascribed to the hydrochloric acid solutions used in their hydrogen electrodes and upon the values of the contact differences of potential which were involved. Loomis (1915) is inclined to accept the value 0.3360.

In 1914 Lewis and Randall applied "corrected degrees of dissociation" to the hydrochloric acid solutions used in arriving at the difference of potential at 25° between calomel electrodes and the theoretical normal hydrogen electrode. Defining the normal calomel electrode as the combination Hg Hg_2Cl_2, KCl (1M), KCl (0.1M) they reach the value 0.2776. The difference of potential between this electrode and the tenth normal they give as 0.0530. Whence the value for the tenth normal electrode is 0.3306. These values were revised by Lewis, Brighton and Sebastian (1917) to 0.2828 for the difference of potential between the normal calomel and the normal hydrogen electrode, and 0.0529 for the difference between the normal and the tenth normal.

Beattie (1920) using more recent data calculates for the potential difference at the normal calomel electrode 0.2826 and compares this value with 0.2824 which is Lewis, Brighton and Sebastian's result (see above) when corrected by Beattie for the liquid junction potential difference between 0.1 N and 1 N KCl.

It will have been noted that in measurements with the hydrogen electrode there is no concern for the absolute difference of potential between mercury and solution. This is because the calomel half-cell is merely a convenient go-between for measurements in which one hydrogen electrode is compared with another. For this reason it is convenient to retain the "normal hydrogen electrode" standard of reference and it so happens that this is in harmony with a general though not universal custom adopted for all electrode measurements.

Other systems are: first, that in which the difference of potential between the mercury and a normal concentration of KCl in a calomel electrode is taken arbitrarily as zero, and second that in which this difference of potential is given what is considered to be its actual value.

Largely upon the basis of Palmaer's (1903) work the value 0.560 volt has been used as the "absolute" difference of potential between mercury and N/1 KCl saturated with calomel in the presence of solid calomel at 18°C. (The mercury being positive to the solution.) There is some skepticism[1] regarding the re-

[1] Whether this is just or unjust is a question concerning which we are in doubt. No critical review in the light of modern researches is known to the author.

liability of this value, but for the particular purpose with which we are now concerned it makes little difference what the value is if proper *relative* relations are maintained. But the difficulty in maintaining proper *relative* relations when there is no *agreement* on a standard basis of reference is made evident when we consider that the temperature coefficient for the absolute difference of potential between mercury and solution is very different from the temperature coefficient for the difference of potential between calomel electrode and hydrogen electrode when the normal hydrogen electrode is defined as having zero potential difference at all temperatures. Thus, as shown by Fales and Mudge (1920), the absolute temperature coefficient of the saturated calomel half-cell is low and has a positive value. But the temperature coefficient of the values for the saturated calomel half-cell as used in hydrogen electrode work is negative and high. Fales and Mudge seem not to have made any independent measurements which furnish more reliable values for the difference of potential between a saturated calomel half-cell and the "normal hydrogen electrode." These authors have however extended the work of Michaelis and have found evidence that the saturated calomel half-cell is reliable within the temperature interval 5°–60°C.

As a working standard the saturated calomel half-cell is undoubtedly the best as pointed out by Michaelis and Davidsohn (1912). It does not require careful protection from the saturated KCl solution usually employed as a liquid junction and it has a high conductivity permitting full use of the sensitivity of a low-resistance galvanometer. It differs in no way from other calomel half-cells except that the solution is saturated with KCl in the presence of solid KCl at all temperatures used.

There is not very good agreement between the values assigned to the saturated calomel half-cell by different laboratories and it had therefore best be regarded for the time being as a good working-standard to be checked from time to time against carefully made normal or tenth normal calomel electrodes or against a hydrogen electrode in a standard solution. For ordinary measurements however the values given in table A of the Appendix are adequate.

Michaelis (1914) gives the following table of values for the potential differences referred to the normal hydrogen electrode for the tenth normal and the saturated calomel electrodes.

TABLE 43

TEMPERATURE	TENTH NORMAL	SATURATED
15		0.2525
16		0.2517
17		0.2509
18	0.3377	0.2503
19		0.2495
20	0.3375	0.2488
21		0.2482
22		0 2475
23		0 2468
24		0.2463
25		0.2458
30	0.3364	
37		0.2355
38	0.3355	0.2350
40	0.3349	
50	0.3326	
60	0.3290	

SUPPLEMENTARY REFERENCES

Abegg (1902), Abegg-Auerbach-Luther (1909), Auerbach (1912), Bjerrum (1911), Bugarszky (1897), Clarke-Myers-Acree (1916), Coggeshall (1895), Coudres (1892), Ellis (1916), Fales-Vosburgh (1918), Lewis-Brighton-Sebastian (1917), Lewis-Sargent (1909), Lipscomb-Hulett (1916), Loomis (1915), Loomis-Acree (1911), Loomis-Meacham (1916), Michaelis (1914), Michaelis-Davidoff (1912), Myers-Acree (1913), Newberry (1915), Palmaer (1907), Richards (1897), Richards-Archibald (1902), Sauer (1904), Steinwehr (1905). See also Chapter XIX for potential values.

CHAPTER XIV

The Potentiometer and Accessory Equipment

The method usually employed in the measurement of potential differences is the Poggendorf compensation method, the potentiometer method. In principle it consists in balancing the potential difference under measurement against an opposing, known potential difference. When the unknown is so balanced no current can flow from it through a current-indicating instrument such as a galvanometer.

The principle may be illustrated by the arrangement shown in figure 24 which is suitable for very rough measurements.

According to modern theory the conduction of electricity in metals is the flow of electrons, the electron being the unit electrical charge. By an unfortunate chance the two kinds of electricity, which were recognized when a glass rod was rubbed with silk, were given signs (+ for the glass and − for the silk) which now leave us in the predicament of habitually speaking of the flow of positive electricity when the evidence is for the flow of negative charges, the electrons. But so far as the illustration of principles is concerned it makes little difference and we shall depart from custom and shall deal with the negative charges in order to make free use of a helpful analogy. We may imagine the electrons, already free in the metal of our electrical conductors, to be comparable with the molecules of a gas which if left to themselves will distribute themselves uniformly throughout their container (the connected metallic parts of our circuits). We may now imagine the battery S (fig. 24) as a pump maintaining a flow of gas (electrons) through pipes (wires) to R to A to B and back to S. The pipe (wire) AB offers a uniform resistance to the flow so that there is a uniform fall of pressure (potential) from A to B while the pump (battery) S maintains a uniform flow of gas (electrons). If we lead in at C and at D the ends of the pipes (wires) from another pump (battery) X, taking care that the high pressure pipe (wire) from X leads in on the high pressure side of AB, we can move C, D or both C and D until they span

a length of AB such that the difference of pressure (difference of potential) between C and D on AB is equal and opposite to the difference of pressure (difference of potential) exerted between C and D by X. Then no current can flow from X through the current-indicating instrument G and we thereby know that balance is attained.

If we know the fall of electrical potential per unit length along AB the difference of potential exerted by X will be known from the length of wire between C and D. We now come to the manner in which this fall of potential per unit length is determined.

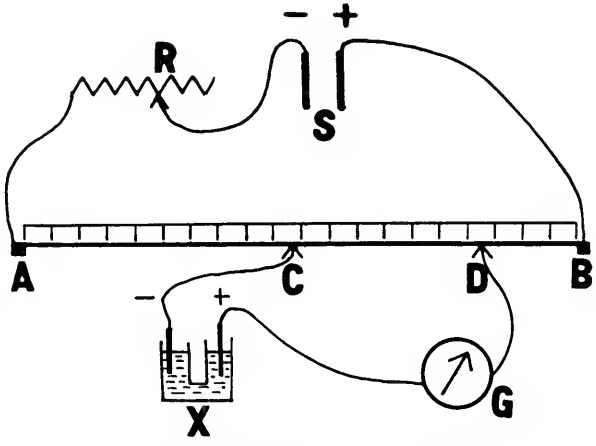

FIG. 24. ELEMENTARY POTENTIOMETER

Choosing for units of electrical difference of potential, electrical resistance and electrical current, the volt, the ohm, and the ampere respectively, we find that they are related by Ohm's law:

$$\text{Current (in amperes)} = \frac{\text{Difference in potential (in volts)}}{\text{Resistance (in ohms)}}$$

or

$$C = \frac{E}{R} \tag{41}$$

With this relation we could establish the fall of potential along AB by measuring the resistance of AB and the current flowing. But this is unnecessary, for we have in the Weston cell a standard

of electromotive force (E. M. F.) which may be directly applied in the following manner. The unknown X (figure 24) is switched out of circuit and in its place is put a Weston cell of known E. M. F. Adjustment of C and D is made until the "null point" is attained, when the potential difference between the new positions of C and D is equal to the E. M. F. of the Weston cell. From such a setting the potential fall per unit length of AB is calculated. It must be especially noted however that for such a procedure to be valid the current in the potentiometer circuit must be kept *constant between the operations of standardization and measurement* for the fundamental relationship upon which reliance is placed is that of Ohm's law $C = \frac{E}{R}$. It will be noted that the establishment of the difference of potential between any two points on AB by the action of S and the resistance of AB is strictly dependent upon the relation given by Ohm's law; but, since we draw no current from X when balance is attained, the resistance of its circuit is of no fundamental importance. It only affects the current which can flow through the indicating instrument G when the potential differences are out of balance. It is therefore concerned only in the sensitivity of G.

The simple potentiometer system described above is susceptible to refinement both in precision and in convenience of operation.

With the inevitable variations in the potentiometer current which occur as the battery runs down it would be necessary to recalculate from moment to moment the difference of potential per unit length of the wire AB if the procedure so far described were used. This trouble is at once eliminated if the contacts of the Weston cell can be thrown in at fixed points and the current be then adjusted by means of the rheostat R so that there is always *the same* uniform current producing, through the resistance between the Weston cell contacts, the potential difference of this standard cell. Having thus arranged for the adjustment of a uniform current at all times and having the resistance of AB already fixed it is now permissible to calibrate the wire AB in terms of volts.

In the Leeds and Northrup potentiometer (fig. 25), the resistance AB of our elementary instrument (fig. 24) is divided into two sections one of which A–D (fig. 25) is made up of a series of

resistance coils between which M makes contact and the other portion of which is a resistance wire along which M′ can slide. When the potentiometer current has been given the proper value, in the manner which will be described, the fall of potential across any one of the coils is 0.1 volt so that as M is shifted from the zero point D the potential difference between M and D is increased 0.1 volt at each step. Likewise, when the current is in adjustment, the shifting of M′ away from D increases by infinitesimal known fractions of a volt the difference of potential between M and M′.

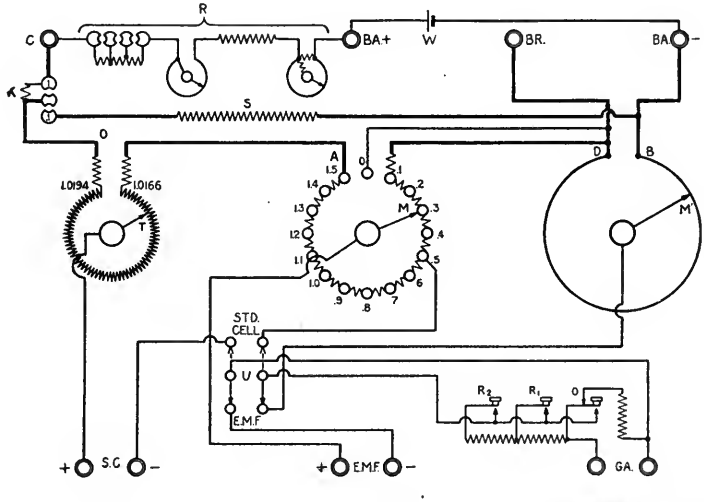

Fig. 25. Wiring of the Leeds and Northrup Potentiometer (Type K)

Now to adjust the potentiometer *current* so that the several resistances in the potentiometer circuit will produce the differences of potential in terms of which the instrument is calibrated, use is made of the Weston cell in the following manner. By means of a switch, U, the unknown is thrown out and the Weston cell is thrown into circuit. One pole of the Weston cell circuit is fixed permanently. The other can be moved along a resistance at T constructed so that the dial indicates the value of the particular Weston cell in use. When so moved to agree with the particular cell in use, this contact at T is left in its position. Now the current flowing from the battery W is adjusted by means of the rheostat R

until the difference of potential between T and 0.5 balances the potential difference of the Weston cell as indicated by the cessation of current in the galvanometer GA. The resistance T to "0.5" is such that the E. M. F. of the battery acting across this resistance will produce the desired potentiometer current. This current now acting across the several resistances furnishes the indicated potentials, i.e., a potential difference of 0.1 volt across each coil.

Another arrangement which employs the ordinary sets of resistances in common use is illustrated in figure 27.

FIG. 26. THE LEEDS AND NORTHRUP POTENTIOMETER

A and B are duplicate sets of resistances placed in series with the battery S. If the current be kept uniform throughout this system the potential difference across the terminals of B can be varied in accordance with Ohm's law by plugging in or out resistance in B. But to keep the current constant while the resistance in B is changed a like resistance is added to the circuit at A when it is removed from B, and removed from A when it is added to B.

As mentioned before, the potential difference could be determined from the resistance in B and a measurement of the current but this is avoided by the direct application of a Weston cell of known potential. Assuming *constant current* a Weston cell replaces X and adjustment to the null point is made by altering the resistance in B with compensation in A. The unknown is then thrown into circuit and adjustment of resistance again

made to the null point. If E_w is the known E. M. F. of the Weston cell, E_x the potential of the measured cell, R_w the resistance in circuit when the Weston cell is in balance and R_c the resistance in circuit when the measured cell is in balance we have

$$C \text{ (constant)} = \frac{E_x}{R_c} = \frac{E_w}{R_w}$$

Whence

$$E_x = \frac{E_w R_c}{R_w} \tag{42}$$

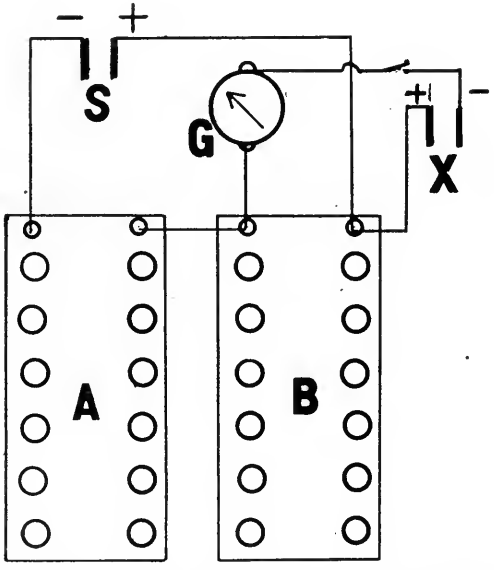

FIG. 27. ELEMENTARY RESISTANCE BOX POTENTIOMETER SYSTEM

The system is improved by providing means of regulating the potentiometer current till constant difference of potential is attained between terminals at which a Weston cell may be thrown into circuit. Then the resistances may be calibrated in volts.

It will be noted that in this arrangement every switch or plug contact is *in the potentiometer circuit.* A bad contact, such as may be produced by failure to seat a plug firmly during the plugging in and out of resistance, or by corrosion of a plug or dial contact, will therefore seriously affect the accuracy of this potentiometer system. It requires constant care.

Lewis, Brighton and Sebastian (1917) used two decade resistance boxes of 9999 ohms each. With an external resistance the current was adjusted to exactly 0.0001 ampere. Thus each ohm indicated by the resistance boxes when balance was attained corresponded to 0.0001 volt. Their standard cell which gave at 25° 1.0181 volts was spanned across B (fig. 27) and 182 ohms of the external resistance.

Another mode of using the simple system illustrated in figure 24 is a device frequently used by physicists, and introduced into hydrogen electrode work by Sand (1911) and again by Hildebrand (1913). Instead of calibrating unit lengths along AD by means of the Weston cell, or otherwise applying the Weston cell directly in the system, the contacts C and D carry the terminals of a voltmeter. When balance is attained this voltmeter shows directly the difference of potential between C and D, and therefore the E. M. F. of X.[1]

A diagram of such an arrangement is shown in figure 28. There is an apparent advantage in the fact that the Weston cell may be dispensed with and resistance values need not be known. There are however serious limitations to the precision of a voltmeter and in two cases which the author knows accuracy within the limited precision of the instruments was attained only after recalibration.

A voltmeter is generally calibrated for potential differences imposed at the terminals of leads supplied with the instrument.

Turning again to figure 24 we recall that when any given fall of potential occurs between A and B, a definite current flows in the circuit SRAB. If the resistance of AB is known a measure of the current flowing permits one to calculate the fall of potential between A and B. Thus a current-measuring instrument (ammeter) placed in series with the fixed resistance AB may be calibrated to indicate differences of potential between A and B.

[1] It is sometimes assumed that because the circuit of the system under measurement is placed in the *position* of a shunt on the potentiometer circuit that its resistance must be high in order that CD (fig. 24) may indicate correctly the potential difference. The fact that no current flows in this branch when balance obtains shows clearly that its resistance can have no effect on the accuracy of the indication. It has also been assumed that if CD is spanned by a voltmeter, the resistance of the voltmeter should be taken into consideration. But a voltmeter is *calibrated* to always indicate the potential difference between its terminals.

To use this system the terminals of the gas chain C and D (fig. 24) are moved to A and B and there permanently fixed. An ammeter is placed between R and S and adjustment of R is made until no current flows in G. The difference of potential between A and B as indicated by the calibrated and renamed reading of the ammeter is then equal to the E. M. F. of the gas chain.

Much the same limitations noted in the voltmeter system apply to the ammeter system.

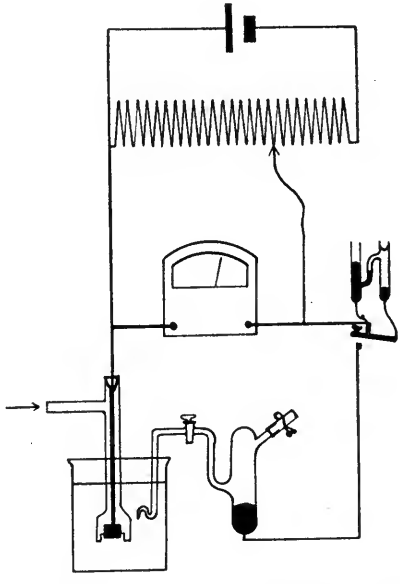

FIG. 28. VOLTMETER POTENTIOMETER SYSTEM

A modification of the system briefly described above is found in the "Pyrovolter." The essential modification is a device of wiring whereby the same indicating instrument is used to measure current (indicated in volts) and to indicate the null point.

In a few instances there has been employed a system of measurement, the principle of which is illustrated in the wiring diagram of figure 29. The E. M. F. of a gas chain is allowed to charge a fixed condenser C. By throwing the discharge key to the right the charge accumulated by the condenser is allowed to discharge through a ballistic galvanometer B, the deflection in which may be made a measure of the accumulated charge and hence of the E. M. F. of the gas chain.

The ballistic galvanometer is one designed to indicate by the angular deflection of its coil the quantity of electricity passing through the coil as a sudden discharge. The quantity of electricity stored in the condenser is a function of its dimensions and material and of the difference of potential imposed at its

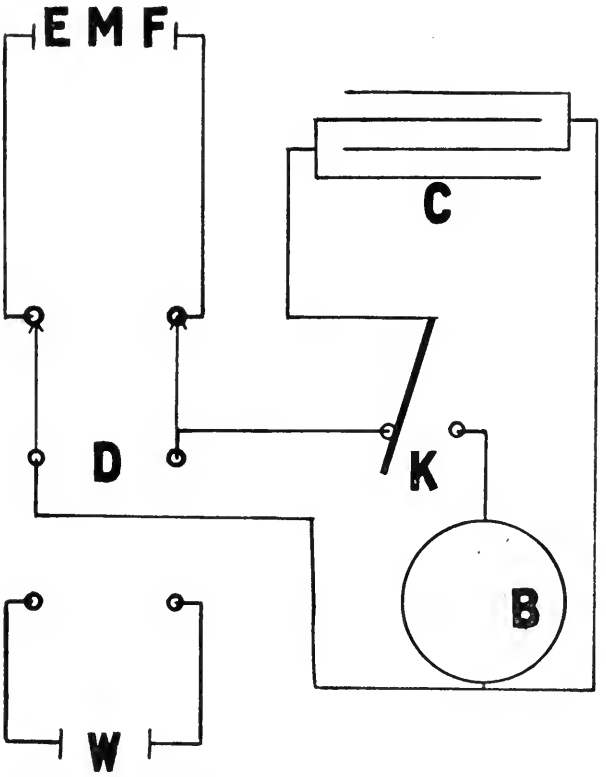

FIG. 29. WIRING DIAGRAM USED IN THE BALLISTIC GALVANOMETER SYSTEM

erminals. The dimensions and material being fixed the charge becomes proportional to the difference of potential. Now a definite difference of potential may be imposed by means of the Weston cell W. The resulting charge in the condenser is discharged through the ballistic galvanometer giving the coil a definite deflection. This serves to calibrate a given set-up if the galvanometer is so designed that the deflection at each section of the

scale is proportional to the quantity of electricity discharged through the coil and if the wiring be such that no serious changes of capacity and inductance occur in manipulation.

The advantage of this condenser method is that the condenser may be conveniently made of such capacity that insignificant current is drawn from the cell under measurement. If then the technique used at the electrodes is refined it should be possible to measure equilibrium potentials which would be easily displaced by current withdrawal. However, until there are published more definite data relating the conditions of electrode measurements to the theory of the condenser method, this system is not to be recommended for ordinary use. In a few instances when the potentiometer had already been adjused to the potential of a gas chain the author has observed what appears to be an excessive E. M. F. unsupported by the equilibrium conditions under measurement. This disappears after an initial throw of the galvanometer and would not be apparent if the measurement were being made by adjusting the potentiometer from an original position sufficiently out of balance to permit a very slight current to flow during successive taps of the key. Will such E. M. F.'s, which are evidently temporary and do not represent the equilibrium conditions under measurement, be recorded in the ballistic galvanometer method?

Goode (1922) has used the 3-electrode vacuum valve in an arrangement for following the electromotive forces of gas chains.

The 3-electrode electron tube is the instrument used as detector and amplifier in radio-communication and is known by various names such as "the audion." A glass bulb (fig. 30) exhausted to a very low gas pressure is supplied with an atmosphere of electrons by their emission from the hot filament F. These electrons produce what may be called a space charge in the bulb. Surrounding the filament is a metallic plate P which can be maintained at a potential about 22 volts more positive than the filament by means of the battery B. Under the influence of this fall of potential electrons migrate from filament to plate, producing the so-called plate-current. But interposed in this electron-drift is a grid, G, of wire or perforated sheet metal through which the electrons must pass in their migration from filament to plate. If this grid is charged positively with relation to the filament it

will tend to neutralize the space charge and so assist the filament-to-plate current. Conversely, if the grid is charged negatively with relation to the filament, it will assist the space charge and so tend to oppose the filament-to-plate current.

Thus the potential difference between filament and grid, a potential difference which may be impressed by a gas chain or other cell, can govern in large measure the filament-to-plate stream of electrons and a measure of this current can be made a measure of the E. M. F. of the cell, C.

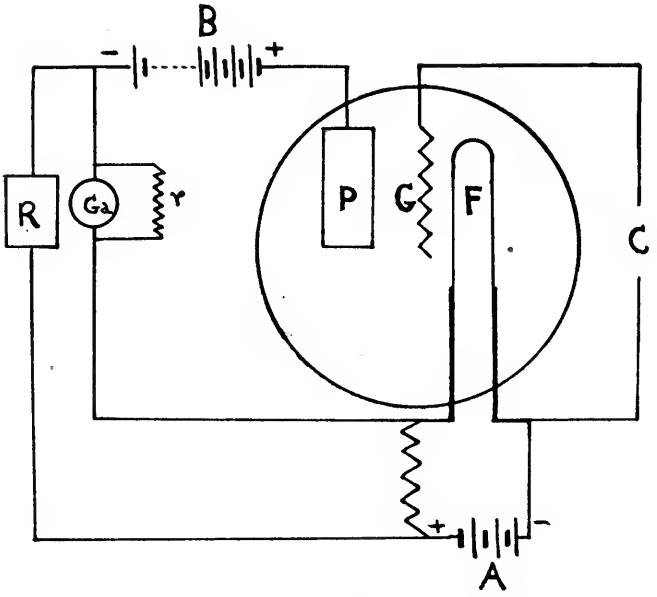

FIG. 30. WIRING OF GOODE'S SYSTEM EMPLOYING THE ELECTRON TUBE

Goode considers the plate current I_p to be made up of a constant current I_o characteristic of a given bulb and set working conditions and a current $I_p - I_o$ which is a function of the potential difference induced by C. To balance I_o Goode found that with the particular bulb he used it was sufficient to place a variable resistance R between the positive terminal of the A battery and the negative terminal of the B battery and to adjust this resistance till the galvanometer Ga was at its zero setting. Under these circumstances the deflection of Ga becomes a function of the

grid potential. Within the range of E. M. F. of the cells under study Goode found that with his particular apparatus the deflection of Ga was a linear function of $I_p - I_o$ when Ga was shunted by a resistance r such that one large scale division corresponded to one unit of pH.

Goode claims that the unique advantage of his system consists in the fact that so little current is drawn from the cell C that continuous readings may be made. This system should, therefore, prove useful in studying those drifts of electrode potential which occur in a variety of cases and which need more thorough investigation.

For the more refined uses to which Goode's system may be put it will be necessary either to know the characteristics of the bulb in use or else to carefully calibrate a given apparatus.

The electron tube, when used as a valve for amplification, should be useful in making hydrogen electrode differences of potential control mechanical devices such as alkali or acid feeds for continuous commercial processes.

NULL POINT INSTRUMENTS

Referring to figure 24 and the accompanying text the reader will see that in the balancing of potential differences by the Poggendorf compensation method there is required a current indicating instrument to determine the null point. Such instruments will be briefly described, and some of their characteristics discussed.

The galvanometer is a current-indicating instrument, which, in the form most useful for the purpose at hand, consists of a coil of wire in the magnetic field of a strong permanent magnet. This coil is connected into the circuit in which the presence or absence of current is to be detected. A current flowing through the turns of the suspended coil produces a magnetic field in its interaction with the field of the permanent magnet and tends to turn the coil so that it will embrace the maximum number of lines of force. The construction of galvanometers need not be discussed since it is a matter for instrument makers, but certain desirable qualities will be treated in a later section, together with the characteristics of other instruments.

Provision should be made for the mounting of a galvanometer

where it will receive the least vibration. If the building is subjected to troublesome vibrations some sort of rubber support may be interposed between the galvanometer mounting and the wall bracket or suspension. Three tennis balls held in place by depressions in a block of wood on which the galvanometer is placed may help. In some instances the more elaborate Julius suspension such as those advertised may be necessary.

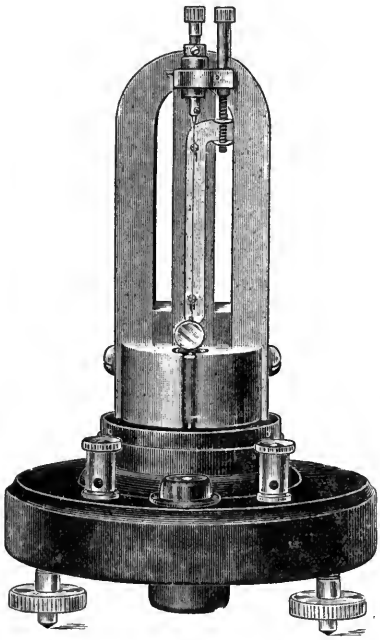

FIG. 31. A GALVANOMETER

The capillary electrometer depends for its action upon the alteration of surface tension between mercury and sulfuric acid with alteration of the potential difference at the interface. A simple form suitable for that degree of precision which does not call for the advantages of a galvanometer is illustrated in figure 32.

Platinum contacts are sealed into two test tubes and the tubes are joined as illustrated by means of a capillary K of about 1 mm. diameter. In making the seals between capillary and tubes the capillary is first blown out at each end and can then be treated as a tube of ordinary dimensions in making a T joint. After a thor-

ough cleaning the instrument is filled as illustrated with clean, distilled mercury, sufficient mercury being poured into the left tube to bring the meniscus in the capillary near a convenient point. In the other tube is now placed a solution of sulfuric acid made by adding 5.8 cc. water to 10 cc. sulfuric acid of 1.84 specific gravity. The air is forced out of the capillary with mercury until a sharp contact between mercury and acid occurs in the capillary. The instrument is now mounted before a microscope using as high power lenses as the radius of the glass capillary will permit. The definition of the mercury meniscus is brought out by cementing to the capillary with Canada balsam a cover glass as illustrated.

An important feature in the use of the capillary electrometer is its short circuiting between measurements. This is done by the key shown in figure 32. Tapping down on the key breaks the short-circuit and brings the terminals of the electrometer into circuit with the E. M. F. to be balanced. If the E. M. F. is out of balance the potential difference at the mercury-acid interface causes the mercury to rise or fall in the capillary. Releasing the key short-circuits the terminals and allows the mercury to return to its normal position. Adjustment of the potentiometer is continued till no movement of the mercury can be detected. To establish a point of reference from which to judge the movement of the mercury meniscus the microscope should contain the familiar micrometer disk at the diaphragm of the eye piece. In lieu of this an extremely fine drawn thread of glass or a spider web may be held at the diaphragm of the eye piece by touches of Canada balsam.

The quadrant electrometer is so little used as a null point instrument that only a brief description will be given. In the form useful for the purpose at hand a very light vane of aluminium is suspended by an extremely fine thread, preferably of quartz, which is metalized on the surface in order to conduct charges to the vane. The vane is surrounded by a flat, cylindrical metal box cut into quadrants. Two opposite quadrants are connected to one terminal and the remaining quadrants to another terminal. If now the vane or needle be charged from one terminal of a high-voltage battery the other terminal of which is grounded, and a difference of potential be established between the two sets

of quadrants, the needle will be deflected by the electrostatic forces imposed and induced. When used as a null point instrument in connection with the potentiometer the two sets of quadrants may be connected as are the terminals of the capillary electrometer and spanned by a short-circuiting key.

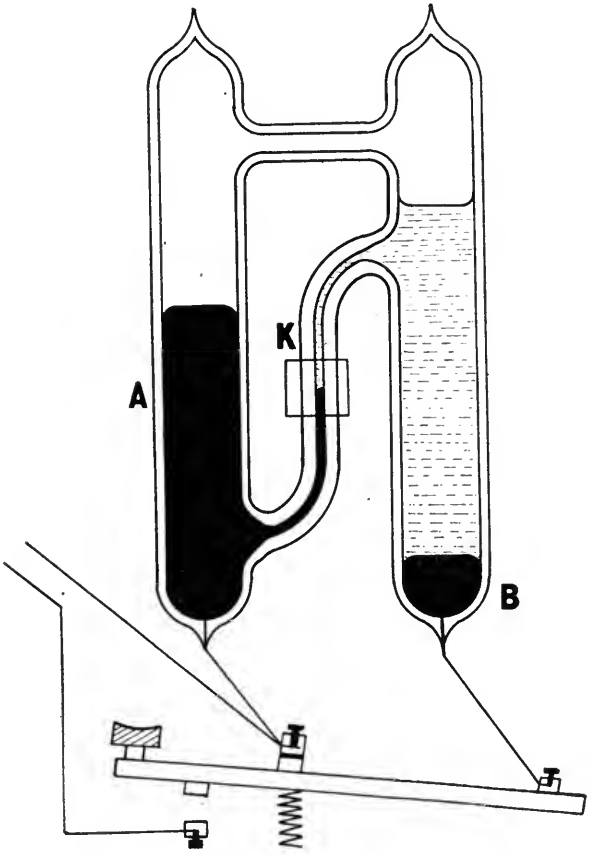

FIG. 32. DIAGRAM OF CAPILLARY ELECTROMETER AND KEY

Since the current drawn for its operation is only the amount necessary to charge a system of very low capacity to the low potential difference when the potentiometer is slightly out of balance with the measured E. M. F. (and to zero potential difference

at balance) the quadrant electrometer might be of special value in the study of easily displaced, electrode equilibria. However, the attainment of the desired sensitivity with some of these instruments is a task requiring great skill and patience. Furthermore the rated sensitivity is sometimes attained by adjusting the so-called electrostatic control to such a value that the zero position of the needle is rendered highly unstable. This combined with the very long period at high sensitivity renders the instrument unsatisfactory for common use. Against these objections are: first, the point mentioned above, and second the advantage that the instrument may ordinarily be left in circuit during the adjustment of the potentiometer as is not the case with the galvanometer.

Telephone receiver. The modern high resistance telephone receiver of the type used in radio reception may serve in an emergency [Kiplinger (1921)]. Lack of balance between potentiometer adjustment and measured E. M. F. is indicated by a click in the receiver when the potentiometer key is tapped; but there is of course nothing but the loudness of the click to show how far from balance the adjustment is, and only the decrement of the sound to indicate that adjustment in the proper direction is being made.

Selection of null point indicators. In the selection of instruments for the measurement of the electromotive force of gas chains it is desirable that there should be a balancing of instrumental characteristics and the selection of those best adapted to the order of accuracy required. A null point instrument of low sensitivity may annul the value of a well-designed, expensive and accurate potentiometer; and a galvanometer of excessive sensitivity may be very disconcerting to use. The potentiometer system and the null point instrument should be adapted one to the other and to their relation to the system to be measured.

The several corrections which have to be found and applied to accurate measurements of hydrogen electrode potentials are matters of a millivolt or two and fractions thereof. Collectively they may amount to a value of the order of 5 millivolts. Whether or not such corrections are to be taken into account is a question the answer to which may be considered to determine whether a rough measuring system or an accurate one is to be used. For all "rough" measurements the capillary electrometer is a good null

point instrument. It has a very high resistance which hinders the displacement of electrode equilibria at unbalance of a crude potentiometer system. It is easily constructed by anyone with a knowledge of the elements of glass blowing, and without particular care may be made sensitive to 0.001 volt.

For "accurate" measurements there is little use in making an elaborate capillary electrometer or in temporizing with poor galvanometers.

The apportionment of galvanometer characteristics is a complicated affair which must be left in the hands of instrument makers, but there are certain relations which should be fulfilled by an instrument to be used for the purpose at hand, and general knowledge of these is quite necessary in selecting instruments from the wide and often confusing variety on the market.

Galvanometer sensitivities are expressed in various ways. Since one's attention is centered upon detecting potential differences the temptation is to ask for the galvanometer sensitivity in terms of microvolt sensitivity. There are two ways of expressing this which lead to different values. One is the deflection caused by a microvolt acting at the terminals of the galvanometer. The more useful value is the deflection caused by a microvolt acting through the external critical damping resistance. But in the last analysis the instrument is to be used for the detection of very small *currents* and these currents when allowed to flow through the galvanometer by the unbalancing of the circuit at a slight potential difference are determined by the total resistance of the galvanometer circuit. The instrument might be such that a microvolt at the terminals would cause a wide deflection, while, if forced to act through a large external resistance, this microvolt would leave the galvanometer "dead." For this reason it is best to know the sensitivity in terms of the *resistance* through which a unit voltage will cause a given deflection. This is the megohm sensitivity and is defined as "the number of megohms (million ohms) of resistance which must be placed in the galvanometer circuit in order that from an impressed E. M. F. of one volt there shall result a deflection of one millimeter" upon a scale one meter from the reflecting mirror (Leeds and Northrup catalogue 21, 1918). The numerical value of this megohm sensitivity also represents the microampere sensitivity if this is defined as the number of millimeters deflection caused by one microampere.

In hydrogen electrode measurements the resistance of the cells varies greatly with design (length and width of liquid conductors) and with the composition of the solutions used (e.g. saturated or M/10 KCl). Constricted, long tubes may raise the resistance of a chain so high as to annul the sensitivity of a galvanometer unless this has a high megohm sensitivity. Dr. Klopsteg (private communication) states that the resistance of the galvanometer coil ideally should be of about the same order of magnitude as that of the cell to be measured if maximum sensitivity is to be gained. Here however we enter complexities, since the arrangements by which high megohm sensitivity is attained affect other galvanometer characteristics. One of these, which is not essential but is desirable, is a short period. A short period facilitates the setting of a potentiometer. If the circuits are out of balance, as they generally are at the beginning of a measurement, the direction for readjustment may be inferred from the direction of galvanometer deflection without bringing the coil back each time to zero setting, but there comes a time when prompt return to zero setting is essential to make sure that slight resettings of the potentiometer are being made in the proper direction.

For a return of the coil to zero without oscillation it is necessary to have some sort of damping. This is generally a shunt across the galvanometer terminals, the so-called critical damping resistance. This shunt permits a flow of current, when the main galvanometer circuit is opened, which is generated by the turning of the coil in the magnetic field. The magnetic field produced in the coil by this current interacting with the field of the permanent magnet tends to oppose the further swing of the coil. When the resistance of the shunt is so adjusted to the galvanometer characteristics that the swing progresses without undue delay to zero setting and there stops without oscillation, the galvanometer is said to be critically damped. Critical damping as applied to deflection on a closed circuit need not be considered when the galvanometer is used as a null point instrument. Since some of the best galvanometers are not supplied with a damping resistance the purchaser of an outfit for hydrogen electrode work should take care to see that he includes the proper unit. Underdamped and overdamped instruments will prove very troublesome or useless.

These very brief considerations are presented merely as an aid in the selection of instruments. The manner in which desirable qualities are combined is a matter of considerable complexity but fortunately makers are coming to appreciate the very simple but important requirements for hydrogen electrode work and are prepared to furnish them. The galvanometer now in use by the author has the following characteristics; coil resistance 530 ohms, critical damping resistance 9,000 ohms, period 6 seconds, sensitivity 2245 megohms. It is not the ideal instrument for the hydrogen electrode system in use but is satisfactory. A shorter period is desirable and a higher coil resistance to correspond better with the average resistance of the order of one to two thousand ohms in some gas chains, would be desirable; but improvement in both of these directions at the same time may increase the expense of the instrument beyond the practical worth. Indeed certain instruments now on the market are satisfactory for almost any type of hydrogen electrode measurements.

In using a galvanometer it is important to remember that while the E. M. F. of a cell is unbalanced its circuit should be left closed only long enough to show the *direction* of the galvanometer deflection. Otherwise current will flow in one direction or the other through the chain and tend to upset the electrode equilibrium. A mere tap on the key which closes the galvanometer circuit is sufficient till balance is obtained.

Of potentiometer characteristics little need be said for the choice in the first instance will lie between instruments sold by reliable makers. In the second instance the choice will lie between instruments of different range and many of the unique instruments may be at once eliminated by a calculation which shows that the reputed accuracy involves too close a scale reading to be reliable. Certain difficulties which enter into the construction of potentiometers for accurate thermo-couple work are hardly significant for the order of accuracy required of hydrogen electrode work. The range from zero to 1.2 volts and the subdivisions 0.0001 volt do for measurements of ordinary accuracy. There should be a variable resistance to accomodate the variations in individual Weston cells of from 1.0175 to 1.0194 volts, and provision for quickly and easily interchanging Weston cell with measured E. M. F.

Several of the features of standard potentiometers may be eliminated without injury to their use for hydrogen electrode measurements and would reduce their cost. Steps in this direction have been taken by at least one manufacturer.

Having described the fundamental principles of the potentiometer it seems hardly worth while to discuss the numerous modifications found among manufactured instruments or used in the construction of home-made designs. With the advent into every town of the numerous and varied parts of radio apparatus certain accessory parts of a potentiometer may be readily purchased and the amateur can concentrate his attention upon the essential resistances. But, unless he is equipped to make these with accuracy and to mount them with care, he may waste the cost of a satisfactory instrument.

With regard to the more special or unique designs found on the market it may simply be said that they were developed for special purposes and that unless these special purposes are to be accomodated, the purchaser will do well to depend only upon an instrument of universal applicability.

When rubber is used as the insulating material of instruments employed as potentiometers the rubber should not be left exposed to the light unduly. The action of the light not only injures the appearance of the rubber but also may cause the formation of conducting surface layers.

If the potentiometer system contains a sliding contact and this contact is *not* involved in the resistance of the primary potentiometer circuit proper, the contact should be kept heavily coated with pure vaseline. If there be any doubt whatever about the quality of this vaseline it should be boiled with several changes of distilled water, skimmed off when cool and then thoroughly dried. If this is done there will seldom be any need to resort to the heroic and dangerous procedure of polishing.

It cannot be too strongly emphasized that while a low order of precision is often adequate for a certain purpose the employment of crude measuring instruments often obscures the data of greatest significance. This statement should not be interpreted as a discouragement to those who are about to undertake measurements with some such system as that illustrated in figure 28 for important data have been obtained with just such instruments.

The statement is intended rather as an encouragement to the beginner who will find the handling of more precise instruments easy and the rewards rich.

THE WESTON CELL

The elementary construction of the Weston cell is shown in figure 33. The mercury in the left arm should be carefully purified (page 239) and the same material should be used for the preparation of the cadmium amalgam. This amalgam consists of 12.5 per cent by weight of electrolytic cadmium. The amalgam is formed by heating mercury over a steam bath and stirring in the cadmium. Any oxid formed may be strained off by pouring the molten amalgam through a test tube drawn out to a long capillary.

Cadmium sulfate may be recrystallized as described by Wolff and Waters (1907). Dissolve in excess of water at 70°C., filter, add excess of basic cadmium sulfate and a few cubic centimeters of hydrogen peroxid to oxidize ferrous iron, and heat several hours. Then filter, acidify slightly and evaporate to a small volume. Filter hot and wash the crystals with cold water. Recrystallize slowly from an initially unsaturated solution. The cadmium sulfate solution of a "normal" Weston cell is a solution saturated at whatever temperature the cell is used, and therefore the cell should contain crystals of the sulfate. The ordinary unsaturated cell has a cadmium sulfate solution that is saturated at about 4°C.

In the study of Weston cells considerable attention has been paid to the quality of the mercurous sulfate. Perhaps the best and at the same time the most conveniently prepared material is that made electrolytically. Where the alternating current is available it is preferable to use it. A good average set of conditions is a sixty cycle alternating current sent through a 25 per cent sulfuric acid solution with a current density at the electrodes of 5 to 10 amperes per square decimeter. With either the alternating or direct current the apparatus described on page 192 is convenient.

In the Weston cell the lead-in wires of platinum should be amalgamated electrolytically by making a wire the cathode in a solution of pure mercurous nitrate in dilute nitric acid.

After filling the cell it may be sealed off in the blast flame or corked and sealed with wax.

Since the preparation of a good Weston cell is a matter of considerable detail, since such cells must be properly and carefully made in order to establish the true potential differences in a potentiometer system, and since reliable cells of certified values may be purchased at a reasonable price, it hardly pays the individual investigator to construct his own. It would, however, be a convenience if the materials could be purchased of the Bureau of Standards as was once proposed.

In some portable Weston cells of commerce the mercury is introduced as amalgamated electrodes and the cadmium sulfate solution, instead of being always in the presence of cadmium sulfate crystals, is often saturated at about 4°C. Since this leaves the solution unsaturated at ordinary temperatures this cell is

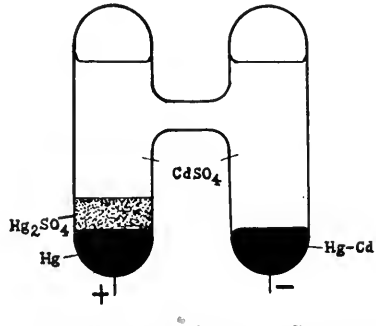

FIG. 33. DIAGRAM OF THE WESTON STANDARD CELL

sometimes called the "unsaturated" type. The result is a cell having a much lower temperature coefficient than that of the "normal" cell. There remain, however, large, if opposite, temperature coefficients for the two arms; and it is therefore necessary to protect the cell from temperature changes which will affect the two arms unequally. Furthermore in all Weston cells there may be observed some degree of hysteresis and in particular cases this may be very marked. It is therefore advisable under all circumstances to protect any Weston cell from temperature fluctuations.

Weston cells are standardized in terms of the international volt the secondary standard for which is the average E. M. F. of

"normal" Weston cells maintained at each national standards laboratory.

As the result of cooperative measurements by the national standards laboratories of England, France, Germany and the United States the value 1.01830 international volts at 20°C. was assigned to the "normal" Weston cell. The United States Bureau of Standards maintains a group of these normal Weston cells whose mean value is taken as 1.0183 international volts and serves for the standardization of the commercial cells. It is important to note that this international agreement came into force January 1, 1911, and that prior to that time the values in force in different countries varied considerably.

TABLE 44

TEMPERATURE	DIFFERENCE
°C.	
0	+0.000,359
5	+0.000,366
10	+0.000,304
15	+0.000,179
20	0.000,000
25	−0.000,226
30	−0.000,492
35	−0.000,791
40	−0.001,114

The temperature coefficient of the "normal" Weston cell is given by Wolff (1908) as:

$$E_t = E_{20} - 0.000,040,75\ (t - 20) - 0.000,000,944\ (t - 20)^2 + 0.000,000,009,8\ (t - 20)^3 \qquad (43)$$

By this formula the differences in volts from the 20° value are as given in table 44.

In other words a *normal* Weston cell should have its certified value corrected by addition of the above corrections when used at temperatures other than 20°C. But an "unsaturated" Weston cell may for all ordinary purposes be considered as having no temperature coefficient and its certified value may therefore be used as given for all moderate variations from 20°C. The change in E. M. F. of the unsaturated type is less than 0.000,01 volt per degree,

provided the precautions regarding temperature fluctuations previously mentioned are observed.

While most commercial cells are of the "unsaturated" type, the purchaser should be informed whether a given cell is of the one type or the other.

STORAGE BATTERIES

The storage battery or accumulator is a convenient and reliable source of current for the potentiometer. Standard potentiometers are generally designed for use with a single cell which gives an E. M. F. of about two volts.

The more familiar cell to which our attention shall be confined consists of two series of lead plates immersed in a sulfuric acid solution of definite specific gravity. The plates of one series are connected to one pole of the cell and the plates of the other series are connected to the other pole. When a current is passed through the cell it will produce lead peroxid upon the plates by which the positive current enters and spongy lead upon the other plates. On charging, therefore, the plates in connection with the positive pole assume the brown color of the oxid while the plates in connection with the negative pole assume the slate color of the spongy metal. The poles should be distinctly marked so that one need not inspect the plates to distinguish the polarity but should the marks become obscured and the cell be a closed cell the polarity should be carefully tested with a voltmeter before attaching the charging current. In lieu of a voltmeter the polarity may be tested with a paper moistened with KI solution. On applying the terminals to the paper a brown stain is produced at the positive pole,—positive reaction at positive pole.

In charging a cell the positive pole of the charging circuit should be connected to the positive terminal of the cell, else the cell will be ruined. If a direct current lighting circuit is available, it may be used to charge a cell, or battery of cells, provided sufficient resistance be placed in series. A 16-candle-power carbon filament on a 110-volt circuit allows about half an ampere to pass. A bank of 6 lamps in parallel will allow three amperes to pass if we neglect the battery resistance. Ordinarily one will do well to charge at a rate lower than that specified by the maker, for the

care of a battery consists chiefly in keeping the deposits even. Low rates of charge and discharge favor this. On charging, the voltage will rise rapidly to 2.35 volts where it will remain during the greater part of the period. When it rises to 2.5 volts the charging should be discontinued. It is when it has reached this voltage that the cell will "gas" vigorously. If a cell should fail to "gas" after a reasonable time it may have an internal short circuit due to warping of the plates or the scaling of conducting material. In searching for such a condition a wooden pry, never a metallic one, should be used. Careful handling and charging will generally prevent such short circuits.

It is more economical to charge from a low voltage circuit but this is seldom available. Indeed there is often available only an alternating current of lighting-circuit voltage. To use the energy of an alternating current it must either be used with a motor generator furnishing a direct current (preferably of low voltage) or else rectified. There are now readily available a variety of rectifiers used in charging the batteries of radio amateurs. Most of these rectifiers when of the mechanical type are designed for charging a six-volt battery. If the operator of a hydrogen electrode has a two-volt cell for his potentiometer and a four-volt battery for operating the relay of the temperature control system he has a combination suited to the common and inexpensive type of rectifier.

In the discharging of a cell the sulfuric acid is converted to sulfate which is deposited. The result is the lowering of the specific gravity of the battery liquid. Thus the specific gravity of the liquid is highest when the battery is fully charged and lowers on discharging. If there be reason to suspect that the proper specific gravity is not being maintained it should be measured with a hydrometer. Fresh sulfuric acid may be added if one follows carefully the specifications given by the manufacturer of the cell. In making fresh solution only sulfuric acid free from metals other than lead, free from arsenic, and free from chloride and nitrate should be used. There will be a continuous loss of water from the battery liquid due to evaporation and gassing. This should be replaced by distilled water *during the recharging of the cell.*

In discharging a cell its voltage should not be allowed to fall below 1.8 volts. When a cell has reached this voltage it should be

recharged immediately. If however the cell has been discharged to a lower voltage it should be recharged at half rate.

In using a storage cell to supply potentiometer current it is essential that the highest stability in the current should be attained since the fundamental principle of the potentiometer involves the maintenance of constant current between the moment at which the Weston cell is balanced and the moment at which the measured E. M. F. is balanced. Steadiness of current is attained first by having a storage cell of sufficient capacity, and second by using it at the most favorable voltage. Capacity is attained by the number and size of the plates. A cell of 60 ampere-hour capacity is sufficient for ordinary work. The current from a storage cell is steadiest when the voltage has fallen to 2 volts. When a potentiometer system of sufficient resistance is used it is good practice to leave the cell in circuit, replacing it or recharging it of course when the voltage has fallen to 1.8 or 1.9 volts, and thus insure the attainment of a steady current when measurements are to be made.

In no case should a cell used for supplying potentiometer current be wired so that a throw of a switch will replace the discharging with the charging circuit. The danger of leakage from the high potential circuit is too great a risk for the slight convenience.

CHAPTER XV

Hydrogen Generators, Wiring, Shielding, Temperature Control, Purification of Mercury

Hydrogen generators. When there is no particular reason for attaining equilibrium rapidly at the electrode a moderate supply of hydrogen will do. When, however, speed is essential, or when there are used those immersion electrodes which are not well guarded against access of atmospheric oxygen an abundant supply of hydrogen is essential. Indeed it may be said that one of the most frequent faults of the cruder equipments is the failure to provide an adequate supply of pure hydrogen or the failure to use generously the available supply.

Hydrogen generated from zinc and sulfuric acid has been used in a number of investigations. If this method be employed, particular care should be taken to eliminate from the generator those dead spaces which are frequently made the more obvious evidence of bad design, to have an abundant capacity with which to sweep out the gas spaces of cumbersome absorption vessels and to properly purify the hydrogen. To purify hydrogen made from zinc and sulphuric acid pass it in succession through KOH solution, $HgCl_2$ solution, P_2O_5, red-hot, platinized asbestos, and a solution of $Na_2S_2O_4$ (See Franzen, Ber., **39**, 906) (Henrich, Ber., **48**, 1915, p. 2006).

A very convenient supply of hydrogen is the commercial, compressed gas in tanks. According to Moser (1920) the industrial preparation varies but the chief methods are the electrolytic and the Linde-Cara-Franck processes. Of these the first yields the better product. Hydrogen by the second process contains among other impurities, iron carbonyl which may be detected by the yellow flame and the deposit of iron oxid formed when the hydrogen flame impinges upon cold porcelain. Moser found that it was impractical to remove this iron carbonyl and he states that hydrogen containing it is unfit for laboratory purposes. On the other hand, electrolytic hydrogen ordinarily contains only traces of air and CO_2 and is free from arsenic and CO. To purify it

pass the gas over KOH and then through a tube of red-hot, platinized asbestos. If it is desired to dry the hydrogen, use soda lime or P_2O_5, but not H_2SO_4 which is reduced. If P_2O_5 is used it should be free from P_2O_3, i.e., distilled in a current of hot dry air.

In purchasing tank hydrogen it is well to be on guard against tanks which have been used for other gases.

For controlling the flow of gas from a high pressure tank the valve on the tank itself is seldom sufficiently delicate. There should be coupled to it a delicate needle valve, if this can be obtained. If not there will be found on the market diaphragm valves for the reduction of the pressure. Even then there should be placed between the tank and the electrode vessel a T tube, one branch of which dips under mercury and forms a safety valve.

Having metal connections to start with, it will be found very satisfactory to lead off with copper tubing, such as is used for automobile connections or specified as soft drawn, seamless copper tubing 4 mm. internal diameter and wall thickness 24 B. S. gauge. This can be soldered in the flame of a blast lamp, using borax for a flux, with a silver solder composed of 6.5 parts copper, 2.0 parts zinc and 11.0 parts silver. This solder is described as fusing at about 983°C. A nickel wire is useful in spreading the flux and solder.

On the whole electrolytic generators are more satisfactory if a direct current such as that of a lighting circuit is available. In figure 34 is shown a generator the body of which is an ordinary museum jar. The glass cover may be perforated by drilling with a brass tube fed with a mixture of carborundum and glycerine. If this mixture is kept in place by a ring paraffined in position, and the brass tube is turned on a drill press with intermittent contact of the drill with the glass, the perforation may be made within a few minutes. The electrolyte used is ten per cent. sodium hydroxid. The electrodes are nickel. To remove the spatter of electrolyte and to protect the material in the heater the hydrogen passes over a layer of concentrated KOH solution; and to remove traces of residual oxygen the hydrogen is passed through a heater. In the design shown the gas passes through a tungsten filament lamp. Lewis, Brighton and Sebastian use a heated platinum wire. More commonly there is used a gas-heated or electrically heated tube containing platinized asbestos. In

the author's design shown in figure 34 the wiring is so arranged that when there is no demand for hydrogen the heater may be turned off at S_2 and a lamp thrown into series with the generating

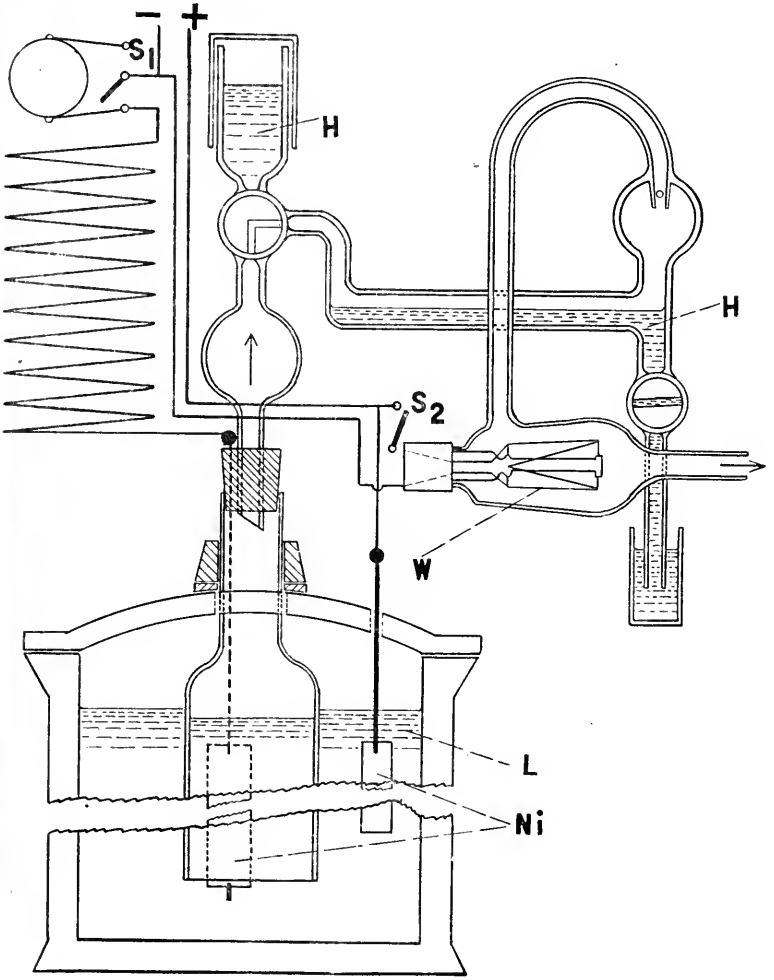

FIG. 34. AN ELECTROLYTIC HYDROGEN GENERATOR

circuit by switch S_1. The generator then continues to operate on a low current and sufficient hydrogen is liberated to keep the system free from air. Such a generator can be run continuously for months at a time. When in use the generator carries about

4.5 amperes. If this current be taken from a high voltage lighting system there must be placed in series a proper resistance which can be either built up by a bank of lamps or constructed from nichrome wire.

Since rubber connections are often used in leading hydrogen it is of interest to note the following relative rates of diffusion of gases through rubber.

Gas	*Rate*
Nitrogen	1.00
Air	1.15
Oxygen	2.56
Hydrogen	5.50
Carbon dioxid	13.57

Wiring. Whenever a set-up is to be made more than an improvisation it pays to make a good job of the wiring. A poor connection may be a source of endless trouble and unsystematized wiring may lead to confusion in the comparison of calomel electrodes and the application of corrections of wrong sign.

Soldered connections or stout binding posts that permit strong pressure without cutting of the wire are preferable to any other form of contact. If for any reason mercury contacts are used they had best be through platinum soldered to the copper lead. Copper wires led into mercury should not take the form of a siphon else some months after installation it may be found that the mercury has been siphoned off.

Thermo-electromotive forces are seldom large enough to affect measurements of the order of accuracy with which we are now concerned if care be taken to make contacts so far as possible between copper and copper at points subject to fluctuations in temperature.

A generous use of copper knife switches, can be made to contribute to the ease and certainty of check measurements. For instance if there be a battery of hydrogen electrodes and a set of calomel electrodes, wires may be led from each to a centre connection of single-pole, double-throw switches as shown in figure 35. All the upper connections of these switches are connected to the + pole of the potentiometer's E. M. F. circuit, and all the lower connections to the — pole. By observing the rule that no two switches shall be closed in the same direction, short-circuiting of

combinations is avoided. The position of a switch shows at once the sign of its electrode in relation to any other that may be put into liquid junction. This is a great convenience in comparing calomel electrodes where one half-cell may be positive to another and negative to a third. Such a bank of single pole switches permits the comparison of any electrode with any other when liquid junction is established; and, if a leak occur in the electrical system the ability to connect one wire at a time with the potentiometer and galvanometer often helps in the tracing of the leak.

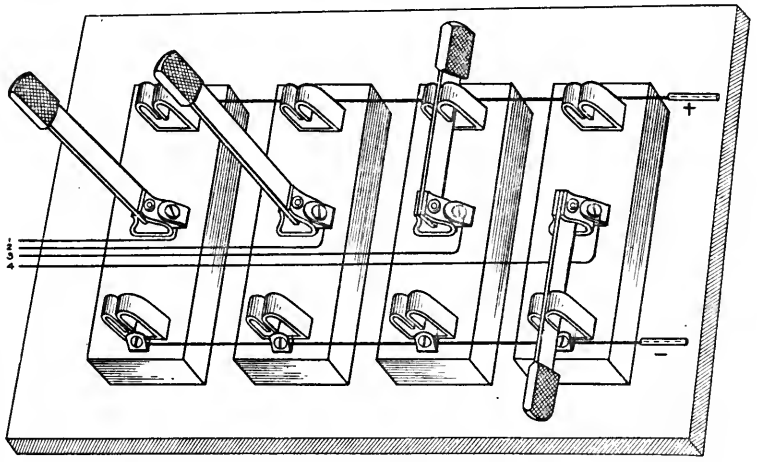

FIG. 35. SWITCHES FOR CONNECTING HALF-CELLS WITH POTENTIOMETER

Shielding. Electrical leaks from surrounding high potential circuits are sometimes strangely absent from the most crude systems and sometimes persistently disconcerting if there is not efficient shielding. The principle of shielding is based on the following considerations. If between two supposedly well-insulated points on a light or heating circuit, or between one point of such a circuit and a grounding such as a water or drain pipe, there is a slight flow of current, the electrical charges will distribute themselves over the surface films of moisture on wood and glass-ware. At two points between which there is a difference of potential the wires of the measured or measuring system may pick up the difference of potential to the detriment of the measurement. If however *all* supports of the measured and measuring systems lie on a good conductor such as a sheet of metal, the electrical leakage from without

will distribute itself over an *equipotential* surface and no differences of potential can be picked up. To shield efficiently, then, it is necessary that *all* parts of the system be mounted upon metal that can be brought into good conducting contact. In many instances the complications of hydrogen electrode apparatus and especially the separation of potentiometer from temperature bath make a simple shielding impracticable. Care must then be taken that all of the separate parts are well connected. Tinfoil winding of wire in contact with unshielded points can be soldered to stout wires for connection to other parts by dropping hot solder on the well-cleaned juncture.

Shielding should not be considered as in any way taking the place of good insulation of the constituent parts of the measured or measuring systems.

For further details in regard to shielding see W. P. White (1914).

Temperature control is a matter where individual preference holds sway. There are almost as many modifications of various types of regulators as there are workers. Even in the case of electrical measurements where orthodoxy interdicts the use of a water bath it has been said (Fales and Vosburgh) that it can be made to give satisfaction.

Yet there are a few who may actually make use of a few words of suggestion regarding temperature control for hydrogen electrode work.

As a rule the water bath is not used because of the difficulty of preventing electrical leakage. Some special grades of kerosene are sold to replace the water of an ordinary liquid bath but for most purposes ordinary kerosene does very well. The free acid sometimes found in ordinary kerosene may injure fine metallic instruments. To avoid this use the grade sold as "acid-free, medium, government oil."

A liquid bath has the advantage that the relatively high specific heat of the liquid facilitates heat exchange and brings material rapidly to the controlled temperature, but compared with an air bath it has the disadvantage that stopcocks must be brought up out of the liquid to prevent the seepage of the oil. The advantage of the high specific heat of a liquid is falsely applied when the constancy of a liquid bath is considered to be a great advantage over the more inconstant air bath. The lower the specific heat of the

fluid the less effect will variation in the temperature of that fluid have upon material which it is desired to keep at constant temperature. For this reason a well-stirred air bath whose temperature may oscillate about a well-controlled mean may actually maintain a steadier temperature in the material under observation than does a liquid bath which itself is more constant. It is the temperature of the material under observation and not the temperature of the bath which is of prime interest.

An air bath can be made to give very good temperature control and since it is more cleanly than an oil bath and permits directness and simplicity in the design of apparatus a brief description of one form used by the writer for some years may be of interest.

A schematic longitudinal section illustrating the main features is shown in figure 36.

The walls of the box are lined with cork board finished off on the interior with "compo board." The front is a hinged door constructed like the rest of the box but provided with a double glass window and three 4-inch hand holes through which apparatus can be reached. On the interior are mounted the two shelves A and B extending from the front to the back wall and providing two flues for the air currents generated by the fan F.

The writer at one time used a no. 0 Sirocco fan manufactured by the American Blower Company, demounted from its casing and mounted in the bearing illustrated. He now uses a four-blade fan taken from a desk-fan and mounted so that it turns in the hole F of the partition and blows toward E. The baffle plates at E, made of strips of tin arranged as in an egg-box, and intended to establish parallel lines of flow when the centrifugal fan was used, are now eliminated.

In the illustration the oil cup is shown as if it delivers into an annular space cut out of the Babbit-metal bearing. In reality this annular space is provided by cutting away a portion of the steel shaft.

The heating of the air is done electrically with the use of bare, nichrome wire of no. 30 B. and S. gauge. When using the centrifugal fan the wire is strung between rings of absestos board (the hard variety known as "transite" or "asbestos wood") which fit over the fan at H. With the blade-fan the partition at F is made of asbestos board and the wire is strung over the opening. The

air is thus heated at its position of highest velocity. The electrical current in this heating coil can be adjusted with the weather so that the time during which the regulator leaves the heat on is about as long as the time during which the regulator leaves the heat off. In other words adjustment is made so that the heating and cooling curves have about the same slope, or so that the heating balances the loss of heat through the walls.

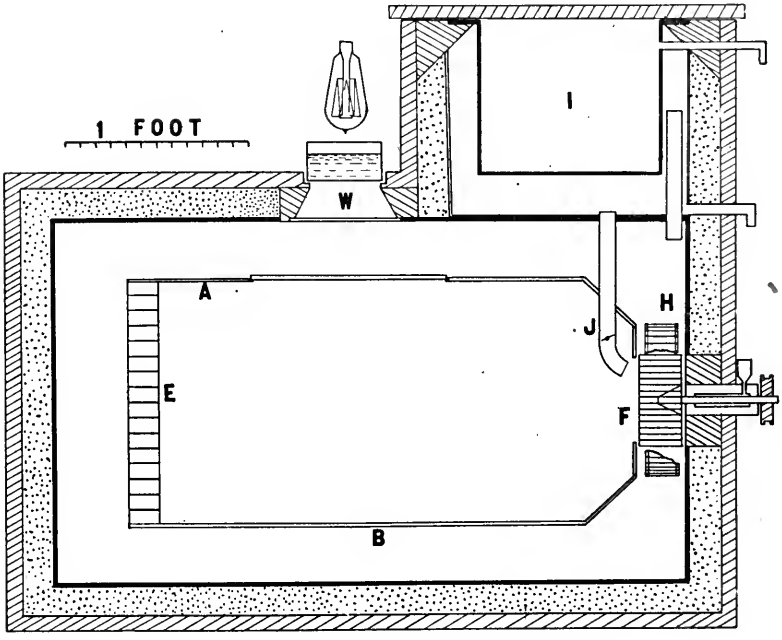

FIG. 36. CROSS SECTION OF AN AIR BATH

When the room temperature is not low enough to provide the necessary cooling the box I is filled with ice water. Surrounding this is an air chamber into which air is forced from the high pressure side of the fan. J should be provided with a damper which can easily be reached and adjusted.

To lessen danger of electrical leakage over damp surfaces the air is kept dry by a pan of calcium chlorid.

A double window at W over which is hung an electric light provides illumination of the interior. A solution of a nickel salt is placed at this window to absorb the heat from the lamp.

The double window in the door (not shown) should be beveled toward the interior to widen the range of vision.

Such a box has been held for a period of eight hours with no change which could be detected by means of a tapped Beckmann thermometer and with momentary fluctuations of 0.003° as determined with a thermo-element. The average operation is a temperature control within ±0.03° with occasional unexplained variations which may reach 0.1°. Because of the slowness with which air brings material to its temperature the air bath is continuously kept in operation, and if a measurement is to be made quickly the solution is preheated.

Given efficient stirring and a considerate regulation of the current used in heating, accurate temperature control reduces to the careful construction of the regulator. For an air bath the ideal regulator should respond instantaneously. This implies rapid heat conduction. Regulators which provide this by having a metal container have been described but glass will ordinarily be used. At all events there are two simple principles of regulator construction the neglect of which may cause trouble or decrease sensitivity and attention to which improves greatly almost any type. The first is the protection of the mercury contact from the corroding effect of oxygen. The second is the elimination of platinum contacts which mercury will sooner or later "wet," and the substitution of an iron, nickel or nichrome wire contact.

After trials of various designs the author has adopted the two forms of regulator head shown in figure 37.

For precise control at an inaccurately adjusted temperature form A is used. The platinum lead-in wire P is fused to the nichrome wire N. After filling the instrument with mercury, dry hydrogen is flushed through the head by way of the side tubes. These are then sealed off and serve as reservoirs for excess mercury. Adjustment is made by slightly overheating the body of the mercury, breaking off the capillary column by a tap of the hand and storing the detached portion in one of the side tubes. Such an adjustment is often troublesome when regulation at a particular temperature is desired; but, once the adjustment is made it is permanent, provided the contact wire is ground down to a fine thread so that it will not fill the capillary enough to cause the mercury thread to part on occasions of overheating.

Form B permits delicate adjustment of the contact by means of the screw S but it requires skill to make such a head properly. The nichrome wire must fit very closely in the capillary R to prevent the wax and mercury seal at W from creeping downward. Such a close fit implies very careful glass blowing to maintain a

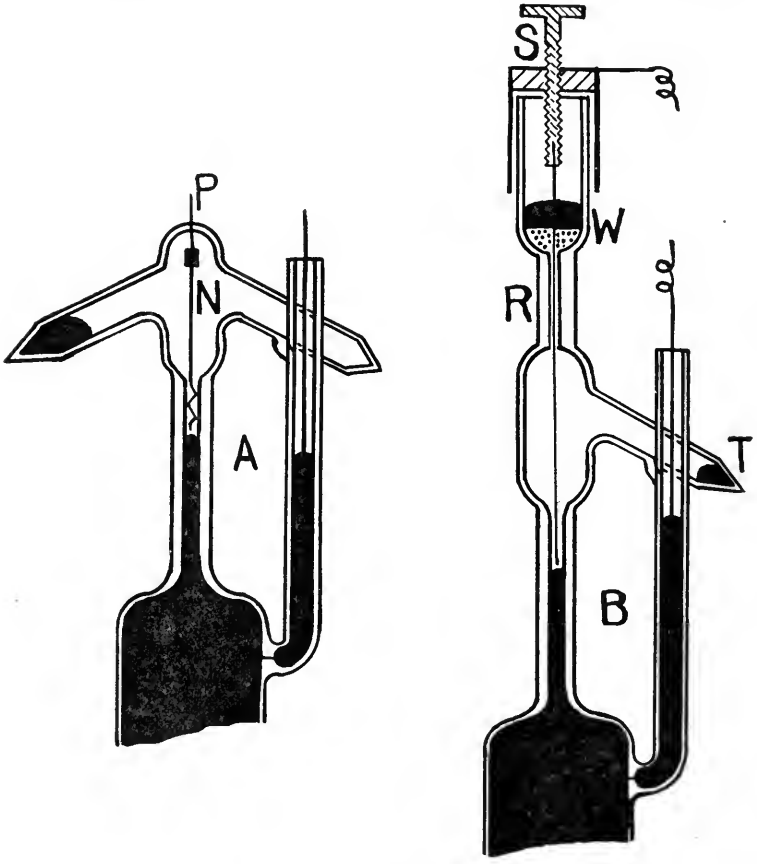

FIG. 37. THERMO-REGULATOR HEADS

straight and unconstricted capillary. With the contact wire in place and the proper amount of mercury in the apparatus hydrogen is run in at T escaping through R. Then a bit of beeswax is melted about W and at the moment it hardens the hydrogen supply is shut off, T is sealed, and then the wax is covered with a *shallow* layer of mercury.

If the wire does not fit R with precision or if overheating occurs the mercury at W may find its way into the regulator head. It is much safer then, although it increases the difficulties of adjustment, to make the seal at W with DeKhotinsky cement.

For an air bath it is best to seal such regulator heads to a grid of tubes.

The permanency of regulators of such design when properly made is a great asset and well repays care in preparation. Regulators of each of these types have been in continuous operation for years without serious trouble. One of type A survived a severe laboratory fire and after readjustment operated well.

Filling such regulators with mercury can be done most easily by first evacuating the vessel under some one of the various high vacuum pumps and then letting the mercury in slowly through one of the side arms drawn to a fine point which is broken under mercury.

A description of methods of purifying mercury will be found on page 239.

For electrical control of temperature the scheme of wiring shown in figure 38 has proved satisfactory.

Lamps which are neat, convenient, replacable forms of resistance, which are obtainable in variety and which indicate whether or not current is flowing are shown in figure 38 by L. R is a resistance formed by a few turns of number 30 nichrome wire on Pyrex glass, porcelain or asbestos board. By shifting the brass contact clamp along this resistance the proper amount of current to operate the relay may be found by trial. Too strong a current is to be avoided. A sharp, positive action of the relay should be provided against the day when the relay contact may become clogged with dust. To reduce sparking at the regulator and at the relay contacts, inductive coils in the wiring should be avoided. Spanning the spark gaps with properly adjusted condensers made of alternate layers of tin foil and paraffine paper may eliminate most of the sparking, if the proper capacity be used. For air regulation it is essential that the heater be of bare wire so that it cools the moment the current is turned off. Furthermore it is essential to adjust the current till the heating rate is close to the cooling rate of the air bath. For such control of the heating current there are inserted in series with the heater

two lamp sockets in parallel permitting the insertion of either a fuse, one lamp or two lamps of various sizes. The other lamp shown in the heating circuit reduces sparking at the relay.

For relay contacts the tungsten contacts used in gas engines are very good.

Although methods of tapping an alternating current for the operation of a relay have been described it is safer to depend upon a battery.

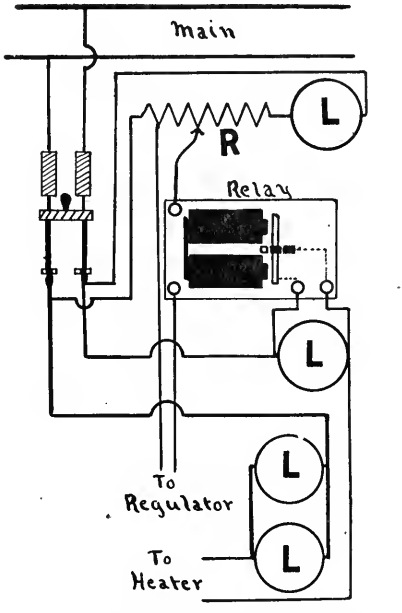

FIG. 38. WIRING FOR TEMPERATURE CONTROL

Purification of mercury. Pure mercury is essential for many purposes in hydrogen electrode work,—for the calomel and the mercury of calomel electrodes, for Weston cells should these be "home made," for thermo-regulators and for the capillary electrometer.

The more commonly practiced methods of purification make use of the wide difference between mercury and its more troublesome impurities in what may be descriptively put as the "electrolytic solution tension." Exposed to any solution which tends to dissolve base metals the mercury will give up its basic impurities

before it goes into solution itself, provided of course the reaction is not too violent for the holding of equilibrium conditions.

The most commonly used solvent for this purpose is slightly diluted nitric acid although a variety of other solutions such as that of ferric iron may be used.

To make such operations efficient it is necessary to expose as large a surface as possible to the solution. Therefore the mercury is sometimes sprayed into a long column of solution which is supported by a narrow U-tube of mercury. The mercury as it collects in this U-tube separates from the solution and runs out into a receiver. To insure good separation the collecting tube should be widened where the mercury collects but this widening should not be so large as to prevent circulation of all the mercury. A piece of very fine-meshed silk tied over the widened tip of a funnel makes a fine spray if the silk be kept under the liquid. This simple device can be made free from dead spaces so that all the mercury will pass through successive treatments. It is more difficult to eliminate these dead spaces in elaborate apparatus; but such apparatus, in which use is made of an air lift for circulating the mercury, makes practicable a large number of treatments. A combination of the air lift with other processes and a review of similar methods has been described by Patten and Mains (1917).

Hulett's (1905, 1911) method for the purification of mercury consists in distilling the mercury under diminished pressure in a current of air, the air oxidizing the base metals. Any of these oxids which are carried over are filtered from the mercury by passing it through a series of perforated filter papers or long fine capillaries. A convenient still for the purpose is made as follows. Fuse to the neck of a Pyrex Kjeldahl flask a tube about 30 cm. long which raises out of the heat of the furnace the stopper that carries the capillary air-feed. Into the neck of the flask fuse by a T-joint seal a 1.5 cm. tube and bend this slightly upward for a length of 15 cm. so that spattered mercury may run back. To the end of this 15 cm. length join the condensing tube, which is simply an air condenser made of a meter length of narrow tubing bent zigzag. Pass the end of this through the stopper of a suction flask and attach suction to the side tube of this flask. The mercury in the Kjeldahl flask may be heated by a gas flame or an electric furnace. For a 220 volt D. C. circuit 12 meters of no. 26

nichrome wire wound around a thin asbestos covering of a tin can makes a good improvised heating unit if well insulated with asbestos or alundum cement. A little of this cement applied between the turns of wire after winding will keep the wire in place after the expansion by the heat.

In the construction of such stills it is best to avoid soft glass because of the danger of collapse on accidental over-heating. Hostetter and Sosman describe a quartz still.

Both the air current, that is delivered under the surface of the mercury by means of a capillary tube, and the heating should be regulated so that distillation takes place smoothly.

Since it is very difficult to remove the last traces of oxid from mercury prepared by Hulett's distillation the author always makes a final distillation in vacuo at low temperature. An old but good form of vacuum still is easily constructed by dropping from the ends of an inclined tube two capillary tubes somewhat over barometric length. One of these is turned up to join a mercury reservoir, the other, the condenser and delivery tube, is turned up about 4 inches to prevent loss of the mercury column with changes in external pressure. The apparatus is filled with mercury by suction while it is inclined to the vertical. Releasing the suction and bringing the still to the vertical leaves the mercury in the still chamber supported by a column of mercury resting on atmospheric pressure and protected by the column in the capillary condenser. The heating unit is wire wound over asbestos. The heat should be regulated by a rheostat till the mercury distills very slowly. By having the mercury condense in a capillary the still becomes self-pumping.

Perhaps few of us who work with mercury have a proper regard for the real sources of danger to health. The vapor pressure of mercury at laboratory temperatures is not to be feared, but emulsification with the dust of the floor may subdivide the mercury until it can float in the air as a distinct menace. Its handling with fingers greasy with stop cock lubricant is also to be avoided on account of possible penetration of the skin but more particularly because of the demonstrated ease with which material on the hands reaches the mouth.

REFERENCES

Potentiometers

Bartell (1917), Bovie (1915), Hildebrand (1913), Leeds and Northrup Catalogue 70, McClendon (1915), Nye (1921), Sand-Law (1911), Slagle-Acree (1921), Wenner-Weibel (1914), White (1914), Will Corporation (1921).

Galvanometers

Leeds and Northrup Company Catalogue 20 (1918), White (1906).

Capillary electrometer

Boley (1902), Le Blanc (1890), Lippmann, G. (1875), Smith (1900) (1903).

Quadrant electrometer

Beattie (1910–12), Compton-Compton (1919), Dolezalek (1906).

Weston standard cell

Bureau Standards Circular 60, Report to International Committee (1912), Cohen-Moesveld (1920), Cohen-Walters (1920), Wolff (1908), Wolff-Waters (1907), Hulett (1906), Melon (1921), Oblata (1920).

International electrical units

Dellinger (1916), Bureau Standards Circulars Nos. 29, 60.

CHAPTER XVI

The Relation of Hydrogen Electrode Potentials to Reduction Potentials

We must remember that we cannot get more out of the mathematical mill than we put into it, though we may get it in a form infinitely more useful for our purpose.—John Hopkinson

As indicated in Chapter X the hydrogen electrode is but a special case of a general relation for the potential difference between a metal and a solution. The hydrogen electrode is constructed of a noble metal laden with hydrogen, and it may be asked what relation it bears to those electrodes which consist of the noble metal alone and which are used to determine the so-called oxidation-reduction potentials of solutions such as mixtures of ferrous and ferric iron.

If a platinum or gold electrode be placed in a mixture of ferrous and ferric sulfate there will almost immediately be assumed a stable potential difference which is determined by the *ratio* of the ferrous to the ferric *ions*. The relation which is found to hold is given by the equation:

$$E_h = E_k - \frac{RT}{nF} \ln \frac{[Ferro]}{[Ferri]} \qquad (44)$$

where E_h is the observed potential difference between the electrode and the standard normal hydrogen electrode, E_k is a constant characteristic of this particular oxidation-reduction equilibrium and equal to E_h when the ratio $\frac{[Ferro]}{[Ferri]}$ is unity, R, T, n and F have their customary significances, and [Ferro] and [Ferri] represent concentrations of the ferrous and the ferric ions respectively. This equation will be referred to later as Peters' equation. Its general form is:

$$E_h = E_k - \frac{RT}{nF} \ln \frac{[RED]}{[OX]} \qquad (45)$$

where [RED] represents the concentration of the reductant and [OX] represents the concentration of the oxidant.

If we plot E_h on one coordinate and the percentage reduction on the other coordinate, we obtain a set of curves identical in form for a given value of n. The position of each curve along the E_h axis is determined by the value of E_k which fixes the middle point. Such a set of curves would present a picture comparable with that shown in figure 2. The picture, however, would be incomplete for reasons which will be given later.

It will be clearly understood that in using the term oxidation or the term oxidant we do not imply that oxygen is concerned. Oxidation is one of those terms established under an old order of thought and carried on into a new order with its meaning broadened. In the transformation of ferrous to ferric iron by chlorine there is every reason to believe that the process is directly one of electron transfer; yet we speak of it as an "oxidation" because it was seen fit at one time to systematize such reactions in terms of the participation of oxygen. The counterpart of oxidation is reduction. This term does not directly indicate any relation to hydrogen, but it is often assumed that hydrogen is concerned in reduction in much the same way that oxygen was thought to be concerned in every "oxidation."

Before coming to a more generalized theory we shall describe the relation between the hydrogen electrode and the oxidation-reduction electrode *in terms of* hydrogen and hydrogen ions.

It is known that certain reducing agents are so active that they evolve hydrogen from aqueous solutions. In such a solution an electrode would become charged with hydrogen and would conduct itself much like a hydrogen electrode. The relations then obtaining can be extended and, if we wish to represent the interaction of the reducing agent with the hydrogen ions, we have:

$H^+ +$ reducing agent $\rightleftarrows H +$ oxidation product.

If equilibrium is established for the above reaction

$$\frac{[H^+]\,[RED]}{[H]\,[OX]} = K$$

or

$$K\frac{[H]}{[H^+]} = \frac{[RED]}{[OX]}$$

Substituting $K\frac{[H]}{[H^+]}$ for the ratio $\frac{[RED]}{[OX]}$ in Peters' equation (45) and placing $n = 1$ for the case at hand we have

$$E_h = E_k - \frac{RT}{F} \ln K\frac{[H]}{[H^+]}$$

Since the atomic hydrogen bears a definite relation to the partial pressure of molecular hydrogen, P, through the equilibrium

$$[H]^2 = K_h P$$

we may substitute, collect constants under another constant K′, bring this under E_k and so obtain:

$$E_h = E'_k - \frac{RT}{F} \ln \frac{\sqrt{P}}{[H^+]} \qquad (46)$$

Compare this with the general relation for the hydrogen electrode

$$E_h = E_H - \frac{RT}{F} \ln \frac{\sqrt{P}}{[H^+]} \qquad (47)$$

E_H in (47) is zero *by definition* when there is used the "normal hydrogen electrode" system of reference. When (46) is placed on the same basis E'_k is also zero, since each of the other terms in (46) is identical with the corresponding term in (47).

In other words we have substituted for the oxidation-reduction equilibrium the corresponding point of equilibrium between hydrogen and hydrogen ions, and have considered the potential difference at the electrode as if it were that of a hydrogen electrode. An inference is that wherever we have an oxidation-reduction equilibrium the components will interact with hydrogen ions (or water) liberating free hydrogen and building up in the electrode a definite pressure of hydrogen. Conversely, if hydrogen is already present in the electrode at a pressure too high for the oxidation-reduction equilibrium in question, hydrogen will be withdrawn until its pressure is in harmony with the oxidation-reduction equilibrium (the position of the latter having been shifted more or less by the reduction). When a constant pressure of hydrogen is maintained at the electrode, as it is in the customary use of the hydrogen electrode, no true equilibrium can be attained

until this hydrogen has so far reduced all the substances in the solution that they can support one atmosphere pressure of hydrogen.

Incidentally it may be mentioned that it is a matter of indifference whether we regard the reductant to interact with the hydrogen ions or the oxidant with the hydroxyl ions or each with water. By use of the equilibrium equations which are involved we reach the same end-result whatever the path. And furthermore by the use of certain theoretical relations between the hydrogen electrode and the oxygen electrode we could *define* potential differences in terms of that of an oxygen electrode.

This method of relating oxidation-reduction to electrode potentials is convenient for showing the condition which must obtain for a true hydrogen electrode potential; but when we attempt to follow some of the logical consequences of this, the customary exposition, we not only meet some serious difficulties but obscure some very important relations.

Let us calculate the hydrogen pressure in equilibrium with an equimolecular mixture of ferrous and ferric chlorid in a solution held at pH 1. A platinum electrode in such a solution will have a potential about 0.75 volt more positive than the "normal hydrogen electrode." Let us consider this to be the difference of potential between a hydrogen electrode at pH 1 and a normal hydrogen electrode. Let us calculate, then, the hydrogen pressure at 25°C. from the equation:

$$0.75 = -0.0599 \log \frac{\sqrt{P}}{0.1}$$

We find the hydrogen pressure to be about 10^{-27} atmospheres. At one atmosphere pressure a gram mol of hydrogen occupies about 22 litres and contains about 6×10^{23} molecules. If the pressure is reduced to 6×10^{-23} atmospheres there would be but one molecule of hydrogen in 22 litres. If reduced to 10^{-27} atmospheres there would be but *one molecule* in about 37,000 litres. To assume any physical significance in such values is, of course, ridiculous.

It is only by courtesy then that an electrode in a mixture of ferrous and ferric iron at pH 1 can be considered as a hydrogen electrode.

This is but an instance of the physically absurd values encountered when *restricted points of view and restricted methods of expressing relations* are applied to electrode potential differences. One or two other instances will be given to illustrate the fact that our present equations are incomplete in that they tell us little or nothing about the *mechanisms* at electrodes (see Langmuir 1916, also Smits and Aten 1916).

Lehfeldt (1899) says of the so-called solution pressures postulated by Nernst and briefly discussed in Chapter X:

> . . . . . we have Zinc..... 9.9×10^{18}
> Nickel..... 1.3×10^{9}
> Palladium..... 1.5×10^{-36}
>
> The first of them is startlingly large. The third is so small as to involve the rejection of the entire molecular theory of fluids.

Lehfeldt then shows that, in order to permit at the electrode the pressure indicated above for palladium, the solution would have to be so dilute as to contain but one or two ions of palladium in a space the size of the earth. No stable equilibrium could be measured under such a circumstance. On the other hand Lehfeldt calculates that to produce the high pressure indicated for zinc "1.27 grams of the metal would have to pass into the ionic form per square centimeter, which is obviously not the case." There is thus very good reason to refrain from attributing a limited and sometimes obviously untrue physical significance to the integration constant in the fundamental equation for electrode potentials (see page 153).

Another aspect of the matter was emphasized in a lively discussion between Haber, Danneel, Bodländer and Abegg in *Zeitschrift für Elektrochemie*, 1904. Haber points out that, if the well established relation between silver ion concentration and the potential difference between a silver electrode and a solution containing silver ions be extrapolated to include the conditions found in a silver cyanide solution, the indicated concentration of the silver ion will be so low as to have no physical significance. Haber mentions the experiment of Bodländer and Eberlein where the potential and the quantity of solution were such that there was present at any moment less than one discrete silver ion. The greater part of the discussion centred upon the resolution of the equilibrium constant into a ratio of rates of reaction, and upon

the conclusion that, if the silver ion in the cyanide solution has a concentration of the order of magnitude calculated, it must react with a speed greater than that of light or else that the known reactions of silver in cyanide solutions must take place partly with the silver complexes and not wholly with the silver ions. However we are now more directly concerned with another aspect of this interesting situation. The potentials observed in silver cyanide solutions are well defined. We may choose to extend to such solutions the relation between the potential of a silver electrode and silver ion concentration. When we do, we find that the silver ion concentration by itself cannot account for the well-defined potential. How then is the stable and reproducible potential supported?

None of these discussions affect in any serious way those relations for concentration chains which are founded upon thermodynamic reasoning provided it be remembered that the thermodynamic reasoning alone does not furnish any conception of the physical mechanisms of a process. The points mentioned do however make it evident that values sometimes used are mere "calculation numbers" employed in a region of extrapolation where the actual physical significance is unknown. The inevitable conclusion is that our equations are insufficiently generalized.

Such "calculation numbers" as those mentioned in the preceding discussion are often of very great usefulness, but lest they continue to obscure phenomena of significance we shall soon have to have equations more intimately related to the mechanisms as Langmuir pointed out in his 1916 paper.

Now it will not remove the fundamental difficulty to use the treatment which follows; but this treatment may aid the student to retain an orderly view of important relations, and it will provide a basis from which to discuss the interrelations of electrodes of different types. From this discussion a generalized point of view will be reached.

It is generally agreed that the fundamental process in oxidation-reduction is an exchange of electrons. A familiar example is:

$$\text{Ferric ion} + \text{electron} \rightleftharpoons \text{ferrous ion}$$
$$Fe^{+++} + e \rightleftharpoons Fe^{++}$$

Since such a reversible reaction is not dependent upon the presence of an electrode (acting as a catalyst) it is probable that

an exchange of electrons is going on continuously. There must then be some condition virtually equivalent to a free-electron pressure. We may imagine a moment in the exchange during which the electron is balanced between the forces of each ion. At this moment the electron may be considered to belong to neither ion and to be a property of the environment. Undoubtedly the situation is not so simple as this picture suggests; and, although the presence of free electrons has been demonstrated in liquid ammonia and methylamine solutions, the experimental evidence is not sufficient to justify our assuming the presence of free electrons in aqueous solutions to be a fact. However, it may be said at once that we are not now concerned with the objective actuality. A pressure of free electrons is merely postulated as the virtual equivalent of a condition not yet clearly formulated; and it is to be used in much the same way that Nernst used "solution tension," destined from the first to be eliminated from those equations which are employed to formulate experimental data.

Assuming then the presence of free electrons as representative of some condition which may be tentatively evaluated in terms of electron pressure, electron concentration, or electron activity, let us consider the electrons to obey the laws of an ideal solution, their concentration thus being amenable to the law of mass action.

Then, for the equilibrium between ferrous and ferric ions we may write

$$\frac{[Fe^{+++}]\,[e]}{[Fe^{++}]} = K_{Fe}$$

Let the symbol [RED] stand for the concentration of a reductant and [OX] for the concentration of the reductant's oxidation product. Then, in general, for the type of reaction represented below where n electrons are concerned we have the equilibrium equation (48)

$$OX + n\,e \rightleftarrows RED$$

$$\frac{[OX]\,[e]^n}{[RED]} = K_1 \tag{48}$$

or

$$[e] = \sqrt[n]{K_1 \frac{[RED]}{[OX]}} \tag{49}$$

For the reaction $2H^+ + 2e \rightleftarrows H_2$ the equilibrium equation is

$$\frac{[H^+]^2\,[e]^2}{[H_2]} = K'_H \tag{50}$$

In (50) $[H_2]$ refers to the concentration of molecular hydrogen in solution. Since we shall deal with the partial pressure of gaseous hydrogen, as is the custom, we introduce $[H_2] = K\,P$ where K is the equilibrium constant and P is the partial pressure of gaseous hydrogen expressed in atmospheres. Collecting constants we have

$$\frac{[H^+]^2\,[e]^2}{P} = K_H$$

or

$$[e] = \sqrt{K_H \frac{P}{[H^+]^2}} \tag{51}$$

By the same procedure similar equations can be developed for any pair of oxidation-reduction products.

We shall now introduce [e] into an equation formulating the difference of potential between an electrode and an aqueous solution with which it is in contact.

We shall assume the presence of free electrons in metals, as is commonly done. We have already postulated free electrons in solution as the virtual equivalent of the ability of the solution to give up electrons to a body brought into the solution. We shall now ascribe to the electrons in the metal phase and to the electrons in the solution phase activities ξ_m and ξ_s respectively, defining activity as Lewis has done (see page 278).

The change in free energy accompanying the isothermal transfer of one Faraday of electrons from one phase to the other is

$$\Delta F = RT \ln \frac{\xi_m}{\xi_s}$$

If E is the difference of potential between metal and solution and F the Faraday, $EF = \Delta F$

Hence:
$$E = \frac{RT}{F} \ln \xi_m - \frac{RT}{F} \ln \xi_s$$

More rigid equations of the same general form have been used by Herzfeld (1915, 1918), Langmuir (1916), Smits and Aten (1916), and Reichinstein (1921) and have been derived by reasoning on kinetic as well as on thermodynamic theory. Certain aspects of the following treatment have been developed more fully by Smits and Aten.

Now in the above equation we have used electron activity. In order to bring the further treatment into harmony with that used consistently throughout this book, we shall have to sacrifice a certain degree of generality and shall imagine that we are dealing with very dilute solutions wherein activity approaches concentration. The like assumption will be made for the activity of the electrons in the metal. Then we may write

$$E = \frac{RT}{F} \ln [e]_m - \frac{RT}{F} \ln [e]_s \qquad (52)$$

where $[e]_m$ is the concentration of electrons in metal and $[e]_s$ the concentration in the solution.

Substitute for $[e]_s$ its equivalent in any one of the equilibrium equations and we have a result such as that given below.

For instance, let two hydrogen electrodes be constructed of the same metal so that when these two electrodes are opposed as in a gas chain the Volta-effect between the electrodes and the copper of the measuring system will be compensated. The total E. M. F. of the gas chain is:

$$\text{E.M.F.} = \frac{RT}{F} \ln [e]_m - \frac{RT}{F} \ln [e]_m + \frac{RT}{F} \ln \frac{\sqrt{K_H \frac{P}{[H^+]^2}}}{\sqrt{K_H \frac{P'}{[H^+]'^2}}}$$

If $P = P'$

$$\text{E.M.F.} = \frac{RT}{F} \ln \frac{[H^+]'}{[H^+]}$$

This is the simplest equation for a hydrogen electrode concentration cell. In a similar way we obtain the equation for a concentration cell of two "reduction potential" electrodes.

It will be noted that in the case mentioned above the terms containing $[e]_m$ certainly cancel out. But will they if for one of

two like electrodes another of a different metal is substituted? Whatever the arguments for and against this may be, we believe that the *electrochemical experimental* data are quite insufficient to decide the question. Lest important phenomena be thus obscured, as Smits believes, the reader should be on his guard; but lest it be supposed that characteristic differences between different metals are thus eliminated it may be said at once that these differences will presently be found to be embodied in a complex of constants. We shall tentatively assume that the concentrations of the electrons in different metals are sufficiently alike to permit differences to be ignored for purposes of approximate treatment and shall regard the term $\frac{RT}{F} \ln [e]_m$ as a constant, E_m.

We then have a general equation for the difference of potential between any electrode and a solution of hypothetical electron concentration $[e]_s$, namely,

$$E = E_m - \frac{RT}{F} \ln [e]_s \qquad (53)$$

To obtain an expression relating the potential difference at an electrode with the equilibria of the ions in solution it is now only necessary to write a given reaction in a form involving electron concentration, to solve for $[e]_s$ and to introduce the equivalent of $[e]_s$ in equation (53). Thus the working equation is obtained by a uniform process, and, whatever the limitations of the development may be, it furnishes at one and the same time an easy method of remembering electrode relations and a viewpoint which helps to clarify the interrelationships of different systems.

Since it will be convenient to refer all electrode potential differences to that of the normal hydrogen electrode as the standard, the nature of the relation will be treated first.

Combine equations (51) and (53) to give

$$E = E_m - \frac{RT}{F} \ln \sqrt{K_H \frac{P}{[H^+]^2}}$$

But

$\frac{RT}{F} \ln \sqrt{K_H}$ is a constant which we may call E_H.

Hence

$$E = E_m - E_H - \frac{RT}{F} \ln \frac{\sqrt{P}}{[H^+]} \tag{54}$$

For an oxidation-reduction electrode we have from equations (49) and (53)

$$E = E_m - \frac{RT}{nF} \ln K_1 \frac{[RED]}{[OX]}$$

or, separating the new constant as we have done above, we have

$$E = E_m - E_1 - \frac{RT}{nF} \ln \frac{[RED]}{[OX]} \tag{55}$$

If now a normal hydrogen electrode and an oxidation-reduction electrode be opposed in a "chain" we have from (54) and (55) the full equation:

$$\text{E.M.F.} = E_m - E_m + E_H - E_1 + \frac{RT}{F} \ln \frac{\sqrt{P}}{[H^+]} - \frac{RT}{nF} \ln \frac{[RED]}{[OX]}$$

By definition E in equation (54) is zero when P and $[H^+]$ are unity. Then $E_m - E_H = 0$. The above equation then (when one of the electrodes is the "normal hydrogen electrode") reduces to

$$\text{E.M.F.} = E_m - E_1 - \frac{RT}{nF} \ln \frac{[RED]}{[OX]} \tag{56}$$

It will be noted that the constant in this equation (algebraic sum of E_m and E_1) is not the simple constant of the oxidation-reduction equilibria, but is a complex. Furthermore the value is dependent upon the standard of reference used—in this case the normal hydrogen electrode. The complex nature of this constant has been discussed by Haber.

It is customary to combine such constants as E_m and E_1 in the last equation. Furthermore it is convenient to maintain the same basis of reference, the normal hydrogen electrode. When this is done it shall be indicated by using for the electrode potential the symbol E_h.

With these understandings we may at once write equations for several types of electrode-solution systems.

For the hydrogen electrode

$$E_h = -\frac{RT}{F} \ln \frac{\sqrt{P}}{[H^+]} \tag{57}$$

For the oxygen electrode

$$E_h = E_{k0} - \frac{RT}{F} \ln \frac{[OH^-]}{\sqrt[4]{P_{O_2}}} \tag{58}$$

For an oxidation-reduction electrode

$$E_h = E_{k1} - \frac{RT}{nF} \ln \frac{[RED]}{[OX]} \tag{59}$$

For a metal electrode in contact with solution containing metal ions of the electrode metal

$$E_h = E''_m - \frac{RT}{nF} \ln \frac{[M]_s}{[M^{n+}]} \tag{60}$$

Here $[M]_s$ is the hypothetical concentration of metal in solution supposedly in equilibrium with the electrode. $[M^{n+}]$ is the concentration of metal ions with n positive charges.

If $[M]_s = K[M]_m$, where $[M]_m$ is the concentration of undissociated metal in the electrode and K is the equilibrium constant, we may substitute and collect constants thereby obtaining:

$$E_h = E' - \frac{RT}{nF} \ln \frac{[M]_m}{[M^{n+}]}$$

If the particular metal is always of the same density and state, and its electron concentration is constant (compare Smits), we can regard $[M]_m$ in the above equation as constant and so obtain equation (61) which is customarily used to relate the potential difference at a given metal electrode to the concentration of the metal ions in the solution.

$$E_h = E_M + \frac{RT}{nF} \ln [M^{n+}] \tag{61}$$

The potentials of amalgam electrodes may be derived in a comparable way.

In correlating all equilibria about the hypothetical electron concentrations of solutions, and connecting each in an electrode potential equation by means of equation (53) there is made evident a definite interrelationship of all reactions involving electron transfer. In the elementary development given, rigidity has been sacrificed for the sake of a simplicity which it is believed represents relations with sufficient truth to indicate the following important matters easily overlooked.

In the first place it is readily perceived that it is a mere matter of choice whether we regard a given electrode to be acting as an "oxidation-reduction electrode" or as a hydrogen electrode; and it only requires extension of the same principle to show that this same electrode can be considered as a metal electrode in equilibrium with a solution of its own ions. As indicated on page 245 a platinum electrode immersed in a solution of ferrous and ferric ions if treated as a hydrogen electrode, furnishes a hydrogen pressure which can be considered only as a "calculation value." By a similar procedure it can be shown that the estimated platinum-ion concentration would be a mere "calculation value" so that we naturally avoid considering the electrode in this case as anything other than a means of picking up electrons in their transfer between Fe^{++} and Fe^{+++}.

Likewise a platinum electrode immersed in a solution may be said to function as an *actual* hydrogen electrode only when a finite concentration or pressure of hydrogen is known or provided. For such a pressure to be definite and stable the solution must be reduced to such an extent that any oxidation-reduction equilibrium in the solution is at a state compatible with the state of the equilibrium of the reaction:

$$2H^+ + 2e \rightleftharpoons H_2$$

which is under measurement. This is another way of stating the principle discussed on page 244.

Another interesting relation is obtained by taking into consideration a certain hypothetical relation between the hydrogen electrode and the oxygen electrode. There are reasons for believing that an oxy-hydrogen gas cell, i.e., a cell composed of a hydrogen and an oxygen electrode, each under one atmosphere of the respective gases should show an E.M.F. of 1.23 volts at all pH values. It is at once evident then that an oxygen electrode should enable one to measure pH values (see equation (58)), or more directly pOH values. As a matter of fact the oxygen electrode does not work well in practice and although Grube and Dulk (1918) believe that they have obtained experimental evidence for the theoretical relation between the oxygen electrode and the hydrogen electrode, the oxygen electrode is by no means a practical instrument. Why this is so has been a matter for

considerable debate. No satisfactory explanation has been offered. If, however, we assume the theoretical relations as a basis for argument, it is evident from what has already been said that we are privileged to express the relations between different electrodes in terms of an oxygen electrode. Likewise it is evident that to obtain an *actual* oxygen electrode potential it would be necessary to oxidize the material in solution to a point compatible with a definite and finite oxygen pressure.

Leaving out all question of the numerical value of the oxy-hydrogen electrode and all question regarding the actuality of a hydrogen or oxygen pressure the genesis of equations (57) and (58) shows that a system can be *defined* in terms of either a hydrogen electrode or an oxygen electrode.

In the second place experimental data obtained with electrode measurements alone do not reveal the components which enter into the *constant* of an electrode potential equation. We shall presently deal with some relations between oxidation-reduction potentials and the pH of the solution, and shall adopt for the sake of convenience the assumption that the reductant is an anion created from the oxidant by the introduction of one or more electrons. But the equations used to formulate the experimental data require only that proper *relative* relations be observed and it would be just as legitimate to consider the relation between oxidant and reductant from either of the following points of view:

$$OX + 2e \rightleftharpoons R\bar{E}\bar{D}$$

$$OX + H_2 \rightleftharpoons \text{hydrogenated reductant.}$$

The same form of electrode equation is obtained in either case and the decision between the two points of view is inextricably bound up in the complex nature of the constants which enter into the working equations.

Thirdly, it is of great practical importance for many studies to note: that in any case where a definite potential difference is to be established at the electrode there must be in the system two species, one of which is the direct or indirect reduction product of the other, and that the ratio of their concentrations or activities must be of finite magnitude. Neglect of this principle is not

infrequent, and is doubtless due to the emphasis which has been placed upon the final, working form of the equation for the difference of potential between a metal and a solution of its ions. In obtaining the final form of this equation certain assumptions have been made and the potential difference at the electrode is made to appear as if it were dependent only upon the concentration of one species, namely the metal ions. Whether this be the explanation or not, there are not infrequently encountered in the literature attempts to measure electrode potential differences with a single oxidant or reductant. It should be plain from a study of figure 39 that, when the oxidant or reductant alone is present, the electrode potential difference becomes asymptotic to the E_h axis. Were it possible to eliminate absolutely every trace of the oxidant, the potential difference obtained with the reductant alone would tend to become infinite. Wherever stable potentials have been reported as having been found with reductant alone it is doubtless due to the presence of the oxidant as an impurity.

From the foregoing discussions it should be evident that the designation of a particular electrode-solution system depends so far as convenience is concerned upon relations which we seek, it being more convenient in some instances to formulate all data in terms of hydrogen electrode potentials and in other instances in terms of reduction potentials. So far as the actual physical maintenance of electrode conditions is concerned the designation of an electrode as of one or the other type will certainly depend upon a finite ratio of two products, one of which is the reduction product of the other; but the discovery of what these species are is often a most difficult problem for the solution of which the electrode equations by themselves are not sufficient.

SOME ELEMENTARY RELATIONS OF HYDROGEN ION CONCENTRATIONS TO OBSERVED "REDUCTION" POTENTIALS

In dealing with an oxidation-reduction equilibrium, as, for instance, that between ferrous and ferric iron, our first concern is with the relation between electrode potential difference and the ratio of the concentrations of the components added, or analytically determined. Now it is found that a given ratio of ferric and ferrous salts does not give the same potential under

all circumstances as it should if we could substitute this fixed ratio in Peters' equation. It is *convenient* to assume that the true ratio to be substituted is the ratio of the ion concentrations and when this ratio can be found its substitution in Peters' equation often yields a good constant. Alteration of the ion concentration from that of the total salt added may be due to incomplete ionization of the salt as added or to the withdrawal of ions by the formation of complexes. Very often the concentration of the active agents is determined by the concentration of the hydrogen ions and it is with this that we are now concerned.

To illustrate the problem let us assume that the active oxidant is neither acidic nor basic so that we can neglect any acidic or basic dissociation and in dilute solution identify the active concentration [OX] with the total oxidant $[S_o]$. Let us next assume that on reduction an electron is introduced into the body to make the reductant virtually acidic. The concentration of active reductant then becomes the concentration of the anion of an acid. [RED] must be identified as $[R\bar{E}D]$, and, when there is sought the relation between observed potentials and total reductant and oxidant, use must be made of the equation for the acid dissociation: $[R\bar{E}D] = \frac{[S_r]\,K_a}{K_a + [H^+]}$ where $[S_r]$ is the total concentration of reductant and K_a is the acid dissociation constant for that particular seat of ionization concerned. Substituting the above in equation (59)

$$E_h = E_{k_1} - \frac{RT}{nF} \ln K_a + \frac{RT}{nF} \ln \left[K_a + [H^+] \right] - \frac{RT}{nF} \ln \frac{[S_r]}{[S_o]}$$

or collecting constants

$$E_h = E_k + \frac{RT}{nF} \ln \left[K_a + [H^+] \right] - \frac{RT}{nF} \ln \frac{[S_r]}{[S_o]} \qquad (62)$$

In order to emphasize the effect of $[H^+]$ let us assume that the ratio $\frac{[S_r]}{[S_o]}$ is to be kept constant while $[H^+]$ is varied. Inspection of (62) shows that while $[H^+]$ is large in relation to K_a, E_h will vary as $\frac{RT}{nF} \ln [H^+]$. When $[H^+]$ approaches and passes K_a, variation of E_h passes over gradually from the relation indicated above to the other extreme where there is no appreciable variation of potential with change in $[H^+]$.

Ordinarily these relations are not perceived because the variation of $[H^+]$ is insufficient, but the principle involved is to be found in the case of ferro-ferricyanide potentials as pointed out by Kolthoff, and they are more clearly to be perceived in the data on the oxidation-reduction potentials of certain dyes briefly reported by Clark (1920) and by Clark and coworkers (1921).

Let us also consider the equilibria of the quinone-hydroquinone system.

$$\text{Quinone} + 2 \text{ electrons} \rightleftarrows \text{anion of hydroquinone}$$

$$OC_6H_4O + 2\,e \rightleftarrows \bar{O}C_6H_4\bar{O}$$

If in equation (59) we identify [OX] as the total concentration of quinone, $[S_q]$, then in the same equation [RED] must be identified as the concentration of the divalent anion of hydroquinone $[\bar{D}]$, and n = 2.

$$E_h = E_{k_1} - \frac{RT}{2F} \ln \frac{[\bar{D}]}{[S_q]} \tag{63}$$

If $[S_d]$ is the total concentration of hydroquinone, $[H_2D]$ the undissociated hydroquinone, $[\bar{H}D]$ the first anion, $[\bar{D}]$ the second anion, K_1 the first acid dissociation constant and K_2 the second acid dissociation constant we have:

$$\frac{[\bar{H}D]\,[H^+]}{[H_2D]} = K_1; \frac{[\bar{D}]\,[H^+]}{[\bar{H}D]} = K_2$$

and

$$[S_d] = [H_2D] + [\bar{H}D] + [\bar{D}]$$

Solving the above equations for $[\bar{D}]$ and substituting in (63) we have:

$$E_h = E_{k_1} - \frac{RT}{2F} \ln K_1K_2 + \frac{RT}{2F} \ln \left[[H^+]^2 + K_1[H^+] + K_1K_2 \right] - \frac{RT}{2F} \ln \frac{[S_d]}{[S_q]} \tag{64}$$

The second term can be combined with E_{k_1} to give E'_k as will be done later.

We shall consider only the order of magnitude of K_1 and K_2 and their combined influence. Scudder's tables give $K_1 = 1 \times 10^{-10}$. Let K_2 be assumed to be of the order 10^{-11}. Neg-

lecting numbers of insignificant orders of magnitude we find that while $[H^+]$ is large in relation to K_1 and K_2 (higher than 10^{-7}) the third term in equation (64) reduces to $+ \frac{RT}{2F} \ln [H^+]^2$.

Then

$$E_h = E_k' - 0.000,198TpH - \frac{0.000,198T}{2} \log \frac{[S_d]}{[S_q]} \qquad (65)$$

Thus, if the ratio of total hydroquinone to total quinone be kept constant, the electrode potential difference, E_h, is a linear function of pH *within the limits of the assumptions made above.* A departure from this relation should begin to appear near pH 9, should become very marked at pH 10, and, if other phenomena could be ruled out, E_h should no longer vary with pH when pH is larger than about 12 provided the magnitude of K_2 has been correctly guessed.

The experimental data to be mentioned in a later chapter indicate that the hydrogen pressure in equilibrium with an equimolecular mixture of quinone and hydroquinone is physically of an entirely negligible magnitude.

As Biilmann has shown (see Chapter XX), a platinum electrode in the presence of a definite mixture of quinone and hydroquinone can be made to measure pH values.

Besides cases of the type given above we have cases such as that of iron where the reaction

$$Fe^{+++} + e \rightleftharpoons Fe^{++}$$

is essentially the destruction by the electron of a point of basic ionization.

It is also conceivable that the addition of two electrons may change an ampholyte to a diacidic compound.

Available data are quite insufficient to show whether or not ionizations at points other than those immediately concerned in the oxidation-reduction process produce a marked effect upon the point actually concerned in the oxidation-reduction process. They probably do for any strain in the electronic forces at one point of a molecule must be felt to some extent at all other points.

There may also be found cases where the electronic fields of force are so altered by the introduction of the electrons concerned

in reduction that the reductant, instead of becoming more acidic or less basic becomes less acidic or more basic. The system hemoglobin-oxyhemoglobin comes to mind; but the available data are altogether too meagre to permit a formulation of actual cases, or even to permit an appraisal of the present method of presentation. We have only to keep in mind the fact that, if this method of treatment proves to be valuable, there may be found a wide variety of cases reducible to a form comparable with that of equation (62). There we find three terms. Of these the middle term is the one which will vary from case to case. It will contain $[H^+]$ and the constants of the oxidation-reduction equilibrium. This term will determine, not only the general form of the curve relating E_h to $[H^+]$, but also deviation or inflexion points when $\frac{[S_r]}{[S_o]}$ and n are kept constant and $[H^+]$ is varied.

Whenever the magnitudes of the equilibrium constants are in such relation to $[H^+]$ that the middle term reduces to $\frac{RT}{nF} \ln [H^+]$, as it may in (64), the electrode potential becomes a linear function of pH. Under these limited circumstances there can be calculated a hypothetical, constant, hydrogen pressure by the method given at the beginning of this chapter,—which pressure may be considered characteristic for the given equilibrium. Since such pressures are often of very small magnitude, and since they vary in magnitude even more than hydrogen ion concentrations, it is sometimes convenient to use a logarithmic system of notation similar to the pH of hydrogen electrode work and to let $\log \frac{1}{P_{H_2}} = rH$, where P_{H_2} is the pressure of molecular hydrogen in atmospheres.

Clark and coworkers have calculated rH values characteristic of various oxidation-reduction indicators. Examples are shown in table 45.

As indicated above such rH values have a limited significance. Even near neutrality the indigo system departs from constant rH and in a manner indicated by a full equation comparable with (64).

The manner in which the three variables—electrode potential, pH and percentage reduction, are related in certain cases is illustrated in figure 39.

When it is desired to express the state of a solution without regard to any particular equilibrium it is best to return to the concept formulated in equation (53) as having the desired generality. But lest terms such as electron concentration, pressure or activity gain an unwarranted appearance of reality through use, and lest numerical values connected with this concept be given meanings too arbitrary, it will be best to retain the use of the electrode potentials themselves and in general to call them *reduction potentials.* These specify with directness the general state of the solution.

TABLE 45

INDOPHENOL-INDOPHENOL WHITE		TETRA SULFONATES OF INDIGO AND OF INDIGO WHITE	
pH	rH	pH	rH
4.36	21.3	3.09	12.2
5.33	21.4	4.51	12.2
6.64	22.0	5.90	12.3
7.00	21.4	6.48	12.5
8.98	20.7		
10.23	20.5	As pH increases rH increases	

Since a given mixture of oxidation and reduction products at a given pH stablizes the "reduction potential" of a solution, we have a condition comparable with the buffer action in the acid-base system. To distinguish stabilization of oxidation-reduction from acid-base buffer action we may use the term *poising action.* Thus a solution may be said to be *poised* at a given reduction potential when the addition or subtraction of oxidants or reductants does not seriously alter the reduction potential.

For example in figure 39, if methylene blue at pH 4.6 is about 75 per cent reduced we know that the reduction potential of the solution should be at about +0.1. If quite appreciable additions of oxidants or reductants do not displace the reduction potential very much from this point it is evident that the solution is "poised" at + 0.1.

This brief outline will have indicated the profound importance of the hydrogen ion concentration of a solution for processes of

oxidation-reduction. A striking demonstration is given in a lecture experiment by Stieglitz (1917, page 292). Formaldehyde in acid solution is comparatively inactive with silver ions. On alkalization of the mixture vigorous reduction of the silver occurs. It may also be shown that a proper mixture of ferro- and ferricyanid is inactive toward indophenol in neutral and alkaline solutions, that up to acidities of pH 4 the potential of the ferro-ferri mixture does not vary with pH while that of indophenol-indophenol white does. At acidities near pH 4 the two systems run into one another and the indophenol is reduced.

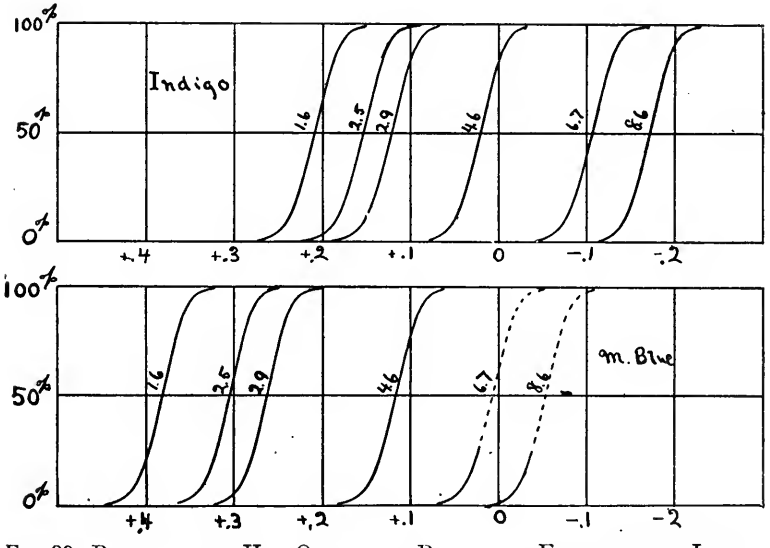

FIG. 39. RELATION OF pH TO OXIDATION-REDUCTION EQUILIBRIA OF INDIGO-INDIGO WHITE AND METHYLENE BLUE-METHYLENE WHITE

Abscissas: reduction potential. Ordinates: percentage reduction. Figures on curves: pH values.

Finally it may be said that all oxidation-reduction equilibria do not lend themselves equally well to potentiometric study. An enormous amount of experimental and theoretical investigation remains to be done.

In passing, it may be mentioned that the instruments and many of the principles which have been here described for the determination of hydrogen ion concentration are applicable in the deter-

mination of oxidation-reduction equilibria and in the titration of oxidizing or reducing substances. The oxidation-reduction electrode with potentiometric measurement has been applied extensively to the determination of the end points of titrations and to the study of oxidation-reduction equilibria.

While the effect of hydrogen ion concentration has been recognized in many of these studies altogether too little use has been made of the methods which have been applied in biochemistry for the control and measurement of pH.

CHAPTER XVII

Sources of Error in Electrometric Measurements of pH

Besides faults in the potentiometric system there are a variety of sources of error which demand special attention. Some of these are specific to hydrogen electrode work; others are not.

Sometimes the most trivial occurrence may cause considerable trouble; such is the bubble of gas that may persistently cling to the bore of a stopcock key which is part of a liquid connection. This is mentioned simply to emphasize the constant watchfulness required of the operator of a hydrogen electrode system. A well-shielded electrical system may be put out of commission in the most unexpected way. Miserly supply of hydrogen with which to sweep out hydrogen electrode vessels is perhaps one of the commonest faults, but the hoarding of solutions which should be used to rinse away the buffer action of solutions previously used in a vessel may also be serious.

Aside from such questions of technique there are certain inherent difficulties in the application of the hydrogen electrode method.

We have already discussed in Chapter XVI the relation between the hydrogen electrode and the "reduction electrode," and have shown that no true hydrogen electrode potential can be attained until the solution is so far reduced that it can support one atmosphere of hydrogen. It is thus made perfectly obvious that a measurement of pH must be preceded by a very thorough reduction of the solution.[1]

When we speak of reduction we mean reduction in its wide sense and include among the oxidizing agents those metal ions which at a given concentration may be reduced by one atmosphere of hydrogen.

The hydrogen electrode if properly treated gives such a precisely defined potential in certain well buffered inorganic solutions, reaches this potential so rapidly, returns when polarized, and

[1] In some instances it is important to remember that reduction of the constituents of a solution may so change the acidic or basic properties of these constituents that serious shifts in pH may occur.

adjusts itself to temperature and pressure changes so well that there is little doubt of its being a reversible, accommodating, relatively quick-acting electrode. It is perhaps because of this that it shows a hydrogen electrode potential in solutions which could be slowly reduced by hydrogen. For instance certain culture media may exhibit upon an electrode of platinum uncharged with hydrogen a potential which is distinctly toward the oxidizing region of oxidation-reduction potential. That they are capable of reduction and that the first reduction potential is not a pseudo potential is shown by the orderly progress of the potential toward that of a hydrogen electrode under the activity of bacteria. Yet such culture media if treated in the first place as in making a hydrogen electrode measurement exhibit a fairly constant and reproducible hydrogen electrode potential the calculated pH value from which checks well with colorimetric measurements. The explanation seems to be that although that complete reduction of material to a point where the oxidation-reduction equilibrium will support an atmosphere of hydrogen is not attained, there is established a virtual hydrogen electrode equilibrium by reason of the rapidity of action between hydrogen and hydrogen ion and the slowness of action between hydrogen and oxidizing agents.

The effect of an intense oxidizing agent will be at once recognized. At the other extreme are the cases where no drift in the E. M. F. in the direction of an oxidizing action at the hydrogen electrode will be detected. Between these extremes lie the subtle uncertainties which make it advisable to check electrometric measurements with indicator measurements and to apply tests of reproducibility, of the effect of polarization, of the effect of time on drift of potential and all other means available to establish the reliability of an electrometric measurement in every doubtful case.

There are effects of unknown cause which are included under the term "poisoned electrodes." An electrode may be "poisoned" by a well defined cause such as those to be mentioned presently; but occasionally an electrode will begin to fail for reasons which cannot be traced. There is hardly any way of putting an observer on his guard against this except to call his attention to the fact that if he is familiar with his galvanometer he will notice a peculiar drift when balancing E. M. F.'s.

Arsenic deposits, adsorption of material by the platinum black

(with such avidity sometimes that redeposition of the black is necessary), the deposit of films of protein, have all been detected as definite causes of electrode "poisoning." Michaelis (1914) places free ammonia and hydrogen sulfid among the poisons. However, there is no special difficulty in obtaining hydrogen electrode potentials agreeing with colorimetric measurements in bacterial cultures containing distinct traces of ammonia or hydrogen sulfid and apparently reliable measurements have been made of the pH values of ammonium-ammonium chloride mixtures. Of the antiseptics used in biological solutions Michaelis (1914) states that neither chloroform nor toluol interfere if dissolved. He does not mention that chloroform may hydrolyze to hydrochloric acid. Drops of toluol however affect the electrode. Phenol is permissible but of course in alkaline solutions participates in the acid-base equilibria.

There is an extensive literature upon the so-called "poisons" which interfere with the catalytic activity of the finely divided noble metals used on the hydrogen electrode. This literature is most suggestive, but there is still need for more direct studies of the conditions surrounding the catalytic activity of the hydrogen electrode.

Simply for the sake of clearness we may distinguish two functions of the electrode. The electrode is first of all a convenient third body by which there is established electrical connection with the system hydrogen-hydrogen ions. That the equilibrium of this system should not be disturbed by the presence of a substance "poisoning" the catalytic activity of the platinum black has been tacitly assumed in the derivation of the thermodynamic equation for electrode potential difference. If the reduction of the solution could be accomplished without dependence upon the catalytic activity of the electrode it should be theoretically possible to attain a true hydrogen electrode potential even in the presence of a "poison." However, in ordinary practice an electrode is used not only as an electrode *per se* but also as a hydrogenation catalyst. As such it is very sensitive to "poisons." "Poisons" are then to be regarded as the cause of sluggish electrodes. Among these we find all degrees. Hydrogenation to a point compatible with a true hydrogen electrode potential may be delayed but slightly and we may say that the electrode is a bit slow in attaining a stable potential without our ever suspecting a "poison;"

or the black metal may be so seriously injured that it becomes entirely impractical to await equilibrium.

And just as "poisons" may render an electrode useless for practical measurements, so the employment of accelerators of catalysis may promote efficiency. With the exception of a brief, unpublished note by Bovie little work has been done in this direction.

From what has already been said the effect of the presence of oxygen is obvious. Indifferent gases such as nitrogen may be considered merely as diluents of the hydrogen and as such must be taken into consideration in accurate estimations of the partial pressure of hydrogen. Gases like carbon dioxid on the other hand act not only as diluents but also become components of any acid-base equilibrium established in their presence.

In very many instances biological fluids contain carbonate and the double effect of the carbon dioxid upon the partial pressure of the hydrogen and upon the hydrogen ion equilibria render accurate measurements difficult unless both effects are taken into consideration and put under control.

At high acidities in the neighborhood of pH 5 carbon dioxide will have relatively little effect upon a solution buffered by other than carbonates. As the pH of solutions increases the participation of CO_2 in the acid-base equilibria becomes of more and more importance. The CO_2 partial pressure in equilibrium with the carbonates of a solution is a function of both the pH and the total carbonate. If, however, we consider for the sake of the argument that the total carbonate remains fairly low and constant, the CO_2 partial pressure becomes less with increase in pH while its effect upon the hydrogen ion equilibria increases with increase in pH. Therefore it may be said that it is of more importance under ordinary conditions to maintain the original CO_2 content of the solution than it is to be concerned about the effect of CO_2 upon the partial pressure of the hydrogen. Furthermore the effect of diminishing the partial pressure of the hydrogen is of *relatively* small importance.

For these reasons the bubbling of hydrogen through the solution is to be avoided unless one cares to determine the partial pressure of CO_2 which must be introduced into the hydrogen to maintain the carbonate equilibria and then provides the proper mixture (Höber). The method usually employed is to use a vessel such as that of Hasselbalch, of McClendon or of Clark in which a

preliminary sample of the solution can be shaken to provide the solution's own partial pressure of CO_2, and in which there is provision for the introduction of a fresh sample with its full CO_2 pressure. The hydrogen supply is then kept at atmospheric pressure and the partial pressure of hydrogen in the electrode vessel is either considered to be unaffected by the CO_2 pressure or corrected from the known CO_2 pressure of the solution under examination.

Of course in cases where the total carbonate in solution rises to considerable concentrations the partial CO_2 pressure may become of very significant magnitude and its effect in lowering the hydrogen pressure must be carefully considered.

In determining the hydrogen ion concentration of the blood by the electrometric method the two outstanding difficulties encountered are the presence of carbonate and oxyhemoglobin. If hydrogen is swept through the fluid it will remove so much of the CO_2 that the hydrogen ion concentration is lowered. If hydrogen is not swept through, the CO_2 will escape into the hydrogen atmosphere about the electrode and reduce the partial pressure of the hydrogen. The oxygen present in the oxyhemoglobin "depolarizes" the hydrogen electrode and makes necessary the employment of the plasma.

Evans (1921) has maintained that in the electrometric measurement of carbonate solutions the carbonate is reduced to formate and that for this reason previous measurements of the pH of blood have been in error. There are various theoretical reasons for doubting the validity of Evans' last conclusion; but since the question is one of fact Cullen and Hastings (1922) have investigated the matter and have failed to confirm Evans.

The criterions of a good hydrogen electrode measurement are difficult to place upon a rigid basis but certain practical tests are easy to apply. Reproducibility of an E. M. F. with different electrodes and different vessels is the foremost test of reliability, but not a final test. Second is the stability of this E. M. F. when attained. It is not always practicable to distinguish between a drift due to alteration in the difference of potential at liquid junctions and a drift at the electrode but in most cases the drift at the liquid junction is less rapid and less extensive than a drift at the electrode when the latter is due to a failure to establish a true hydrogen-hydrogen ion equilibrium. A test which is sometimes applied is to polarize the hydrogen electrode slightly and

then see if the original E. M. F. is reëstablished. This may be done sufficiently well by displacing the E. M. F. balance in the potentiometer system. Where salt and protein errors do not interfere the gross reliability of a hydrogen electrode measurement may be tested colorimetrically. This checking of one system with the other is of inestimable value in some instances as it has proved to be in the study of soil extracts. There the possibilities of various factors interfering with any accurate measurement of hydrogen ion concentration dimmed the courage of investigators until Gillespie (1916) demonstrated substantial agreement between the two methods. Subsequent correlation of various phenomena with soil acidity so determined has now established the usefulness of the methods.

In addition to the tests so far mentioned there remains the test of orderly series. Certain of the general relations of electrolytes are so well established that, if a solution be titrated with acid or alkali and the resulting pH values measured, it will be known from the position and the shape of the "titration curve" whether the pH measurements are reasonable or not. This of course is a poor satisfaction if there is any reason to doubt the measurements in the first place but it is a procedure not be scorned.

In dealing with protein solutions Robertson (1910) found that the electrode was injured by deposits of protein which he ascribed to acid coagulation of the protein by the acid absorbed in the platinum black from previous measurements. Robertson therefore recommends that in a series of measurements with protein solutions the series be treated from the alkaline to the acid solutions. If his explanation be true there are instances where the reverse procedure should be followed. See sections on isoelectric points.

Not infrequently the attempt is made to measure electrometrically the pH value of an unbuffered solution such as that of KCl. It is not entirely the fault of the method but rather of the nature of the solution that this is a task requiring the very highest refinements known to experimental art. If for the sake of the argument we assume that the solution under examination is that of a *perfectly* neutral salt having under *ideal* conditions a hydrogen ion concentration of 0.000,000,1 N, a simple calculation will show what an enormous displacement in pH will be caused by the admittance of the slightest trace of CO_2 from the atmosphere,

of alkali from a glass container, of impurities occluded in the electrode or of impurities carried into the solution with the solvent or solute. Conversely, even if the measurement were such as to give the true value under ideal conditions it would have little practical significance because of the difficulty in holding the conditions ideal.

By the same reasoning it appears probable that it would be difficult to obtain true electrode potentials even with a potentiometric system drawing no current during its adjustment. When no buffer is present there is a negligible reserve of hydrogen ions. But the introduction of the electrode with its enormous surface must displace the equilibrium. How much the displacement will be depends both on relative proportions of electrode and solution and on the technique used.

The effect of temperature variations upon the accuracy of electrometric measurements is a question upon which it is difficult to pass judgment. Of course, if measurements are not intended to be refined one may assume the temperature of the room to be the temperature of the system at the moment of the electrical measurement. It is then a simple matter to select from tables the values and factors applicable at the selected temperature. Since such a procedure introduces errors which are not serious for many purposes the author's insistence upon temperature regulation has been criticized. Those who take this position are doubtless able to escape the psychological effects of uncertainty, but they can hardly escape the inconvenience of having to deal with new values and new factors with every shift in temperature. Temperature control so simplifies rough measurements that much time is saved, and for this reason is recommended even when it is unnecessary. But before the practice of neglecting temperature control can have scientific standing it needs more experimental investigation than it has been accorded. Calculations are quite insufficient for we have little data upon the hysteresis in the adaptation of different systems to temperature variation.

Cullen (1922), finding that the temperature in an electrode vessel is seldom that of the surrounding air in a room subject to temperature variation, has devised a modification of the Clark electrode vessel whereby the temperature of the *solution* can be measured. The same modification can easily be made in a calomel electrode vessel.

CHAPTER XVIII

Standard Solutions for Checking Hydrogen Electrode Measurements

In the routine measurement of hydrogen ion concentrations it is desirable to frequently check the system. To do so in detail is a matter of considerable trouble; but if a measurement be taken upon some solution of well defined pH, and it is found that the potential of the chain agrees with that determined by careful and detailed measurements upon all parts, it is reasonably certain that the several sources of E. M. F. are correct.

Any one of the buffer mixtures whose pH value has been established may be used for this purpose, but there are sometimes good reasons for making a particular choice.

Sørensen (1909) used a mixture of 8 volumes of standard glycocoll solution to 2 volumes of standard hydrochloric acid solution for the details in the preparation of which see page 109. Michaelis (1914) recommends what has come to be known as "standard acetate." This is a solution tenth molecular with respect to both sodium acetate and acetic acid. Its preparation and hydrogen electrode potential at 18°C. have been carefully studied by Walpole (1914). Walpole proposes two methods for its preparation:

(1) From N-sodium hydroxid solution free from carbon dioxid and N-acetic acid adjusted by suitable titration (using phenolphthalein), so as to be exactly equivalent to it.

(2) From N-sodium acetate and N-acetic acid adjusted by titration of a baryta solution, the strength of which is known exactly in terms of the N-hydrochloric acid solution used to standardize electrometrically the normal solution of sodium acetate.

Walpole defines N-sodium acetate as a "solution of pure sodium acetate of such concentration that when 20 cc. are taken, mixed with 20 cc. of N-hydrochloric acid, and diluted to 100 cc. the potential of a hydrogen electrode in equilibrium with it is the same as that of a hydrogen electrode in equilibrium with a solution 0.2 normal with respect to both acetic acid and sodium chloride." By mixing the N-acetate with the N-HCl in accordance with this

definition and then determining the potential of a hydrogen electrode in equilibrium with it Walpole shows that the N-sodium acetate solution may be accurately standardized. In the following table are given Walpole's values showing the relation of

TABLE 46

CUBIC CENTIMETERS OF N/1 HCl TO 20 CUBIC CENTIMETERS N/1 NaAc DILUTED TO 100 CUBIC CENTIMETERS	E. M. F.
19.00	0.5270
19.40	0.5155
19.50	0.5125
19.90	0.4945
20.00	0.4898
20.39	0.4712
20.89	0.4549
21.00	0.4525

the E. M. F. of the chain: Hg | Hg_2Cl_2 KCl (0.1M) | KCl (sat.) | Acetate | H_2Pt at 18°, to the cubic centimeters of N-HCl added to 20 cc. N-sodium acetate and diluted to 100 cc. If, for instance, the potential found is 0.4800 volts, the ratio $\frac{\text{Concentration of HCl}}{\text{Concentration of Na Ac}}$ is $\frac{20.2}{20.0}$. Hence the sodium acetate is 0.9901N.

These values are more convenient to use if plotted as Walpole has done.

TABLE 47

TEMPERATURE	E. M. F.	TEMPERATURE	E. M. F.
15	0.5170	21	0.5180
16	0.5171	22	0.5183
17	0.5172	23	0.5186
18	0.5174	24	0.5190
19	0.5175	25	0.5195
20	0.5178	34–38	0.5200–0.5205

Walpole found that the E. M. F. of the chain: Pt H_2 | "standard acetate" |sat. KCl| 0.1M KCl Hg_2Cl_2| Hg at 18°C. is 0.6046. The contact potential still to be eliminated was estimated by the Bjerrum extrapolation to be 0.0001 volt. Hence the true poten-

tial is 0.6045. This value seems to be the value of the chain corrected to one atmosphere hydrogen plus vapor pressure.

Michaelis (1914) gives the values in table 47 for the difference of potential between the saturated KCl calomel electrode and the hydrogen electrode in his standard acetate.

It will be noted that both Sørensen's standard glycocoll and the standard acetate solutions must be constructed by adjustment of the components. While there is no great difficulty in this there remain the labor and the chance of error that are involved. Clark

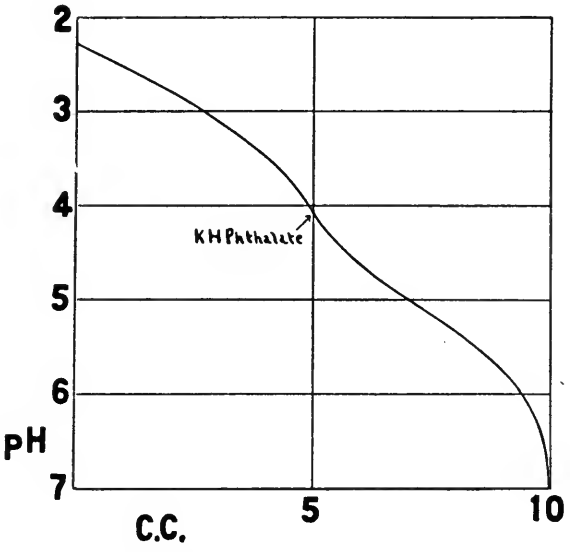

FIG. 40. TITRATION OF PHTHALIC ACID WITH KOH

and Lubs (1916) have shown that acid potassium phthalate possesses a unique combination of qualities desirable for the standard under discussion. The first and second dissociation constants of phthalic acid are so close to one another that the second hydrogen comes into play before the first is completely neutralized (see fig. 40). As a consequence the half-neutralized phthalic acid (KHPhthalate) exhibits a good buffer action. The salt of this composition crystallizes beautifully without water of crystalliza-

tion, and, as was shown by Dodge (1915) and confirmed by Hendrixson (1915) it is an excellent substance for the standardization of alkali solutions. As such it is used to standardize the alkali entering into the buffer mixtures of Clark and Lubs (see page 102). The outstanding feature is that the ratio of acid to base is fixed by the composition of the crystals and not by adjustment as in other standards. The salt may be dried at 105°C. and accurate concentrations constructed. The diffusion potential against saturated KCl is somewhat higher than that of standard acetate as estimated by the Bjerrum extrapolation but not so high as to make good readings difficult.

Clark and Lubs (1916) found for the chain:

$$HgHg_2Cl_2 \mid KCl \text{ (saturated)} \mid M/20 \text{ KHPhthalate} \mid H_2 \text{ Pt}$$

at 20°C. an E. M. F. of 0.4807 corrected to one atmosphere of hydrogen. Their saturated calomel electrode was 0.0882 volt more negative than the average of a set of tenth normal calomel electrodes. Assuming 0.3379 (cf. Chapter XIX) as the value of the tenth normal calomel electrode and 0.0004 volt for the diffusion potential still to be eliminated, the hydrogen electrode potential of M/20 KHPhthalate at 20° is 0.2306.

Unfortunately the temperature relations of such chains are not accurately known. For ordinary work the pH of M/20 KHPhthalate may be considered as 3.97 between 20° and 30°C. Assuming a liquid junction potential difference of 0.0004 volts we can reckon from these data the following total electromotive forces at various temperatures of the chain:

Calomel electrode of KCl conc. X	Sat. KCl	Hydrogen electrode at one atmosphere in KHPhthalate $\left(\frac{m}{20}\right)$

TABLE 48

TEMPERATURE	TOTAL E. M. F.		
	X=0.1M	X=1.0M	X=saturated KCl (approximate)
18	0.5675	0.5158	0.4800
20	0.5689	0.5170	0.4802
22	0.5704	0.5181	0.4806
24	0.5719	0.5192	0.4812
26	0.5733	0.5204	0.4817
28	0.5748	0.5215	0.4822
30	0.5763	0.5227	0.4827

These values are entirely provisional for temperatures other than 20°C. and require experimental verification before they can be used for precise standards. They are given as convenient standards for ordinary check measurements.

CHAPTER XIX

Standardization of pH Measurements

In the development of the theory of electrolytic dissociation the hydrogen electrode came upon the scene comparatively late and after many of the quantitative relations had been established by conductance data. It was therefore natural that these data should have been accepted in the standardization of potentiometric measurements. It now appears that the interpretation of conductance data is more complicated than at first supposed and that certain of the values that have been used in the standardization of potentiometric measurements are in doubt. The resulting confusion demands careful consideration.

Let us review briefly the way in which conductance data enter into the potentiometric system.

The following equation relates the potential difference, E, at a hydrogen electrode to the partial pressure, P, of hydrogen, the concentration of hydrogen ions, C, and the constant K,

$$E = K - \frac{RT}{F} \ln \frac{\sqrt{P}}{C}$$

As shown in a previous chapter we are forced to one or another set of comparisons such as is found in a concentration cell where P and K are constant. In this case we have a measurable electromotive force and the relation

$$\text{E. M. F.} = \frac{RT}{F} \ln \frac{C_1}{C_2}$$

Thus we determine the *ratio* of two hydrogen ion concentrations if the solutions are sufficiently dilute to permit the application of the gas laws from which the above equation was derived. To apply this equation directly to the determination of either concentration C_1 or C_2 the other concentration must be known. Conductance data have been relied upon to furnish the known concentration.

Likewise, when a chain composed of a calomel electrode and a

hydrogen electrode is used, the value assigned to the calomel electrode is such that when it is subtracted from the total E. M. F. of the chain the resulting E. M. F. is as if between a normal hydrogen electrode and the hydrogen electrode under measurement. This implies the experimental determination of the difference of potential between a normal hydrogen electrode and the calomel electrode or else between the calomel electrode and a hydrogen electrode in some solution of *known* hydrogen ion concentration. To determine this known hydrogen ion concentration conductance data upon hydrochloric acid solutions have been relied upon.

Unfortunately hydrochloric acid solutions exhibit the so-called anomalies of strong electrolytes which have already been mentioned. Although it was known from the first that hydrochloric acid solutions do not obey the dilution law, it was supposed that the ratio of the equivalent conductances at dilution v and at infinite dilution (where there is complete dissociation) would give the percentage ionization at dilution v and hence the hydrogen ion concentration at this dilution. However, this conclusion involves the assumption that the mobilities of the ions remain unaltered between dilution v and infinite dilution. Jahn (1900) and Lewis (1912) have questioned this assumption and within recent years the conclusion has become firmly established among many investigators that the mobilities do change or else that the chemical activity of the ions of strong electrolytes is not strictly proportional to their concentration. In other words conductance data alone are not sufficient to define with precision the hydrogen ion concentrations of the hydrochloric acid solutions which have been used to standardize the hydrogen electrode system of concentration chains. In support of this contention there have been brought forward comparisons of the concentration chains themselves. There is evidence that the ratio $\frac{C_1}{C_2}$ in the concentration chain formula is not necessarily determined with accuracy when a measurement of the E. M. F. of such a chain is taken. What is it then that is determined? The way in which this question will be answered will doubtless form another interesting chapter in the philosophy of science. Focused upon this point are two tendencies; the one seeking to find the factors which interfere with the application of the simple gas laws so that the experimental data may be corrected

to apply to the "ideal;" the other seeking to formulate either the empirical data or the thermodynamic relations without special reference to the mechanisms involved.

It was an astute suggestion of Lewis (1907) that the simple thermodynamic relations be assumed to hold, not for concentration pressure relations, but for quantities which, when introduced into the equations embodying the gas laws, will make these laws apply. The two new quantities are *activity* and *fugacity*. In the special case of a "perfect" solution, a very dilute solution, obeying the laws of gases, activity and fugacity are equal to concentration and pressure respectively. But when a solute ceases to conduct itself in accord with the laws of gases, its fugacity and activity remain such that the equations which apply to "perfect" solutions still hold.

Stated in the above manner it may appear to those who insist upon looking for the means of applying concentration relations as if Lewis had made use of a clever dodge. In reality he has simply expressed in a form which he has developed into a self-consistent system that which is the more directly determined experimentally. This is at once evident in the definition of activity by the following postulates.

1. When the activity of a substance is the same in two phases, that substance will not of itself pass from one phase to the other. 2. When the activity of a substance is greater in one phase than in another, the substance will pass from the one phase into the other, when they are brought together.

With these postulates Lewis proceeds to develop a self-consistent system in which it appears that in a "concentration cell" the ratio of activities is related to the E. M. F. by the equation

$$\text{E. M. F.} = \frac{RT}{nF} \ln \frac{\text{activity 1}}{\text{activity 2}}$$

Only at infinite, or very high dilution, when a solution approaches an "ideal" solution, does the more familiar relation of concentration hold true. So long as the limitations were well understood it was permissible to speak of the hydrogen electrode method as a means of determining relative concentrations. If one is willing to use Lewis' terms he would be more precise to speak of the hydro-

gen electrode method as a means of determining relative hydrogen ion activities.

We may note at this point that if we adopt the activity concept and if we refer electrode potential differences to that of the normal hydrogen electrode, confusion is introduced by the use of the term *normal concentration* in the definition of the normal hydrogen electrode. This is clarified if we adopt the definition of Lewis and Randall: "A solution is said to be at (hypothetical) molar concentration with respect to hydrogen ion when the activity of hydrogen ion in this solution is n times as great as in $1/n$ M solution of hydrogen ion, where n is a large number."

The use of the equation given above instead of the equation involving concentrations only shifts our immediate problem to a new position. We are still concerned with a *ratio* and must somehow establish a point of reference. At first sight we have also shifted to a position from which it is difficult to obtain any connection with weights of materials (concentrations).

A formal relation between activity and concentration may be set up by the use of the socalled activity coefficient. Of this Lewis and Randall (1921) state: "The term activity coefficient has been used in two senses, sometimes to mean the ion activity divided by the assumed ion molality, and sometimes to express the ion activity divided by the gross molality of the electrolyte."

Now, if we have a solution of HCl so dilute that we may assume the activity of the hydrogen ion equal to the concentration, and if at the same time the solution is so dilute that we may assume complete ionization, we have a starting point, for then the hydrogen ion activity may be determined from the analytical concentration of the HCl. By the use of the electromotive force equation relating activities we can establish by experiment the relative activity of the hydrogen ion in a more concentrated solution. But there is little assurance that such measurements of relative activity have been made with the highest accuracy because of the experimental and theoretical difficulties of liquid junction potential differences.

By means of conductivity some idea is obtained of ion concentrations and by means of activity coefficients activity and concentration are

related. But since exact treatment of the subject necessitates discussion of assumptions the reader is referred to the original literature.

Using the most probable values for the corrected degree of dissociation of hydrochloric acid solutions, the E. M. F. of the cell: normal calomel electrode-hydrogen electrode in N/10 or N/100 HCl, and the estimated contact potential difference at the liquid juncture, Lewis and Randall obtained the value 0.2776 for the difference of potential between the normal calomel and the normal hydrogen electrodes at 25°. This value was revised to 0.2828 by Lewis, Brighton and Sebastian (1917). Direct comparison with N/10 KCl calomel electrode, as will be noted later, gave 0.3357 as the potential value of this electrode including a slight liquid junction potential difference.

Now let us consider the values hitherto used in biochemical work.

In Sørensen's work, published prior to the adoption of the present standard value of the Weston standard cell, the basis for the particular cell whose value he gave was not stated. If it was the 1.01863 used in Germany prior to 1911 the correction of Sørensen's data to the present international volt will not be significant. Doubtless the international standard was used in Denmark when Sørensen (1912) published the summary of the data of Sørensen and Koefoed. Their values involve two assumptions; first that liquid junction potential differences were eliminated by the Bjerrum extrapolation; second, that in the calculation of the theoretical difference of potential between the normal hydrogen electrode and the hydrogen electrode in the hydrochloric acid solutions used, the correct hydrogen ion concentration was given by conductance data. As already stated there is serious doubt of the validity of the last assumption. Even so we ought, by using the same degree of dissociation for hydrochloric acid solutions, to reconcile Sørensen's value with that of Lewis, Brighton and Sebastian. Sørensen assumed 91.7 per cent dissociation of 0.1M HCl at 18°C. Employing the same value at 25°, as an approximation, we would find that the hydrogen electrode in 0.1M HCl should be 0.0614 volts more negative than a "normal" hydrogen electrode. If however we take "the corrected concentration of H^+ in 0.1M HCl as 0.0816" (Lewis, Brighton and Sebastian) then the difference would be 0.0643. The correction 0.0029 should

bring Sørensen's value into harmony with that of Lewis, Brighton and Sebastian. However, they are:

Lewis, Brighton and Sebastian	0.3357
Sørensen (corr.)	0.3347

The discrepancy of 0.0010 volt remains to be explained. That it may be ascribed partly to an involved potential difference between N/10 KCl and N/1 KCl which has not been noted in the discussion and partly to an excess correction for diffusion potential through the use of the Bjerrum extrapolation seems probable from the treatment accorded this subject by Fales and Vosburgh; but if we attempt to correct Sørensen's data by the use of the curves given by Fales and Vosburgh the discrepancy noted above widens. It is of no particular importance to attempt further to reconcile the two values because Sørensen's original data (1909) show wide variations in the E. M. F.s. of the chains in which hydrochloric acid was used. One might therefore jump to the conclusion that Sørensen's value is unworthy of further consideration now that we have a more probable value. It must be emphasized however that we are not so much concerned with the reliability of Sørensen's original data as we are with the fact that the value thereby assigned to the tenth normal calomel electrode has been widely used in the study of hydrogen electrodes in solutions which exhibit comparatively low diffusion potentials against KCl and which furnish hydrogen electrode potentials reproducible with a considerable degree of precision. Because of this, because of the fact that the Sørensen value and other comparable values have standardized an enormous amount of biochemical data we regard it as important to consider the old value further.

When Sørensen's value has not been used directly it has been used indirectly in the taking over of pH values assigned to standard solutions such as standard acetate. In Walpole's study of acetate mixtures he appears to have been consistent in using the value assigned by Sørensen to the tenth normal calomel electrode referred to the normal hydrogen electrode under one atmosphere of hydrogen plus vapor pressure. He obtained a value for the hydrogen electrode potential in standard acetate agreeing with that found by Sørensen and by Michaelis. In Clark and Lubs' study of phthalate, phosphate and borate buffer mixtures they applied the Bjer-

rum extrapolation, and, with the qualifications stated in their paper reached a value[1] for their tenth normal calomel electrode in substantial agreement with Sørensen's.

Palitzsch doubtless used the Sørensen value, which he originally aided in determining, in his study of borate buffer mixtures.

A variety of similar channels might be followed to show that in the biochemical literature there is substantial agreement so far as the assumed difference between the tenth normal calomel and the normal hydrogen electrodes is concerned. Since the liquid junction potential differences between saturated KCl and the buffer solutions and physiological fluids dealt with in biochemistry are of a low order of magnitude it seems fair to assume that the more precise biochemical data are fairly well *standardized*, though not necessarily accurate. The agreement was furthermore encouraged in other lines of investigation by the recommendation of Auerbach (1912) when, in his summary of the work of the "Potential Commission," he recommended the use of the tenth normal calomel as a working standard because of its low temperature coefficient, and assigned the value 0.337 for use between 20° and 30°.

On the one hand, then, we have what may be regarded as a tacitly accepted and not yet precisely formulated standardization of the tenth normal calomel electrode; and on the other hand a distinctly different value for the tenth normal calomel electrode that is doubtless more nearly correct, though the details by which the value was reached are not presented. The biochemist is thus placed in an embarrassing position. Before making a choice he may consider the present situation in our knowledge of the temperature coefficients of calomel electrodes.

In dealing with the temperature coefficients it will be distinctly understood that we are not concerned with the temperature coefficient of the absolute difference of potential between mercury and solution but rather with the temperature coefficient of the calomel electrode in the cell: calomel electrode-normal hydrogen

[1] Clark and Lubs give their E. M. F.'s reduced to refer to the normal hydrogen electrode under a standard hydrogen *concentration* rather than the standard pressure usually used. Since the calomel values were also referred to the same basis the pH values given by these authors remain as if the customary procedure had been followed.

electrode, when the potential difference at the normal hydrogen electrode is *defined* to be zero *at all temperatures*. Unfortunately we have little data upon this temperature coefficient which are *both* accurate and extensive. Therefore one who chooses to take over the better value for the tenth normal or the normal calomel electrode will still be left in the predicament of not knowing the precise value to use at temperatures other than 25°C.

We can only reach approximate values in the following manner and compare the results with comparatively old experimental data.

Lewis and Randall (1914) have derived a provisional temperature coefficient for the normal calomel electrode which indicates that the values are not a linear function of the temperature. The derivation of these authors as applied to the tenth normal electrode will be followed, but some new values obtained since the writing of their paper will be introduced.

For the cell

$$\mathrm{Pt\ H_2} \left| \begin{matrix} \mathrm{HCl} \\ \mathrm{0.1\ M} \end{matrix} \right| \mathrm{Hg_2Cl_2\ Hg}$$

Lewis and Randall give the empirical equation

$$E = 0.0964 + 0.001881T - 0.000{,}00290T^2$$

whence

$$dE/dT = 0.001881 - 0.00000580T$$

For present purposes this conforms closely enough with Ellis' (1916) data.

It is now assumed that the temperature coefficient of the cell

$$\mathrm{Pt\ H_2} \left| \begin{matrix} \mathrm{HCl} \\ \mathrm{0.1\ M} \end{matrix} \right| \mathrm{Hg_2Cl_2\ Hg}$$

will apply to

$$\mathrm{Pt\ H_2} \left| \begin{matrix} \mathrm{HCl} \\ \mathrm{0.1\ M} \end{matrix} \right\| \left. \begin{matrix} \mathrm{KCl} \\ \mathrm{0.1\ M} \end{matrix} \right| \mathrm{Hg_2Cl_2\ Hg}$$

if the tenth molar hydrochloric acid calomel cell has the same potential as the tenth molar KCl calomel cell. Compare however Lewis, Brighton and Sebastian (1917) who give 0.0012, and MacInnes (1919) who gives 0.0.

For the cell

$$\mathrm{Pt\ H_2} \left| \begin{matrix} \mathrm{HCl} \\ 0.1\ \mathrm{M} \end{matrix} \right\| \begin{matrix} \mathrm{H^+} \\ \mathrm{M} \end{matrix} \Big| \mathrm{Pt\ H_2}$$

Lewis, Brighton and Sebastian give 0.0644. Assuming that in this cell the E. M. F. is proportional to the absolute temperature, $\frac{dE}{dT} = 0.00022$. Hence for the tenth molar KCl calomel electrode against the normal hydrogen electrode

$$\frac{dE}{dT} = 0.00166 - 0.00000580T.$$

For the cell

$$\mathrm{Hg\ Hg_2Cl_2} \left| \begin{matrix} \mathrm{KCl} \\ 0.1\ \mathrm{M} \end{matrix} \right| \begin{matrix} \mathrm{KCl} \\ 1.0\ \mathrm{M} \end{matrix} \Big| \mathrm{Hg_2Cl_2\ Hg}$$

the author finds at 20° 0.0519, and at 30° 0.0536. Interpolation between these values on the assumption that the E. M. F. is a linear function of the temperature gives an E. M. F. at 25° which is within 0.15 millivolts of that found by Lewis, Brighton and Sebastian, and a linear temperature coefficient of 0.000,17. Sauer's value at 18° is 0.0514 and that of Fales and Vosburgh at 25° is 0.0524. Neither of these values falls in with those mentioned above but when taken by themselves and with the 15° value, 0.0509, given in the footnote of the paper by Fales and Vosburgh (1918) they furnish a temperature coefficient of the same order.

With these data we can start from the value 0.2828 as that of the normal calomel electrode (Lewis, Brighton and Sebastian, 1917) at 25°; or with Sørensen's (1912) value, 0.3380, for the tenth normal calomel electrode at 18° and treating each set separately we reach the comparisons shown in table 49.

Bjerrum's values at 0°, 25° and 75° do not fit in with the calculations given above.

The values given above are admittedly uncertain and are to be regarded as provisional in lieu of the experimental data that is needed. It may be emphasized however that there is good reason to believe that the temperature coefficient for the tenth normal electrode is much lower than that of the normal calomel electrode.

Since we can as yet only make a good guess of the temperature relations it seems wise to choose as a standard the calomel electrode with the smaller temperature coefficient and thus lower one chance of error. This fortunately has been, for the most part, the practice in biochemical work although it runs counter to preferences which will not be discussed.

TABLE 49

	LEWIS		TENTH AGAINST NORMAL CALOMEL	SØRENSEN		
t	1.0 N	0.1 N		1.0 N	0.1 N	0.1 N Found
18	0.2844	←0.3360	0.0516	0.2864	←**0.3380**	0.3380
20	0.2840	←0.3359	0.0519	0.2860	←0.3379	0.3378
25	**0.2828**	→0.3356	0.0528	0.2848	←0.3376	
30	0.2817	←0.3353	0.0536	0.2836	←0.3372	0.3370
37.5					0.3364	
40					0.3360	0.3359
50					0.3341	0.3344
60					0.3317	0.3321

Approximate temperature coefficient of normal calomel electrode −0.000,23.

Approximate temperature coefficient of tenth normal calomel electrode −0.000,06.

Let us then assume that this half cell, the tenth normal calomel electrode, is to be the standard to which all working electrodes are to be referred and let us consider finally the choice of values to be assigned.

At 25°C. the difference between the values for the tenth normal calomel electrode given in table 49 is 2 millivolts. A change of this amount would shift the values in the pH scale 0.03 unit pH. This is quite insignificant or within the experimental error in many biochemical studies. For certain purposes it is not insignificant. When carried into mass action relations it might be serious *but* in such relations there are generally involved data taken over from conductance measurements. In such a situation therefore there

are involved complexities which are by no means covered by the mere selection of a more probable value for the standard electrode.

We have already mentioned the fact that even if the value of Lewis, Brighton and Sebastian be absolutely correct at 25° we cannot assign accurately known values at temperatures other than 25°, and we have noted the more or less tacit assumption of standard values for various temperatures in the course of the development of biochemical applications.

In addition to the difficulties mentioned above there is a fundamental question which runs throughout all present-day calculations. As we have reiterated, all hydrogen electrode measurements are referred by one route or another to some experimental standard and the hydrogen ion concentration or hydrogen ion activity, as the case may be, is estimated for this experimental standard by the use of theory which at present is in a state of flux. One's inclination is to accept the latest value advocated by the most advanced thought and yet it is an open question whether the inherent relativity of the whole subject will not force us ultimately to adopt an arbitrary standard. While certain investigators are accepting the value for the normal calomel electrode given by Lewis, Brighton, and Sebastian, Bjerrum is applying the theory of complete dissociation of salts and reaching a very different value. In the author's opinion it will be wise during the present transition period to adopt a provisional standard and in lieu of agreement reached in convention to let that standard be in harmony with that tacitly implied in the greater body of data. The author therefore suggests that the values in column 6 of table 49 be used as provisional standards wherever there is no definite reason to *require* any other value.

We can thus preserve uniformity in pH data and not introduce ill-considered changes which may need subsequent frequent revision before the present theoretical difficulties are removed or before the action of an international committee fixes a standard value.

It may be objected that under such a procedure of standardization the symbol pH loses the precise significance which has been attached to it. It has always been defined as $\log \frac{1}{[H^+]}$. If the

"concentration chain" does not determine with precision the ratio of two hydrogen ion concentrations but rather the ratio of two hydrogen ion activities, and if, in addition, we adopt a standard of reference in the current use of the hydrogen electrode which is not strictly true, then pH is no longer expressive of the true value of $\log \frac{1}{[H^+]}$. We need not be concerned with the casuistry of this situation. We need only remember that the more precise uses to which hydrogen electrode measurements may be put involve theoretical difficulties which we are not yet prepared in every case to deal with accurately,[2] that in the more common uses the uncertainty is not of a serious magnitude and that it is preferable to maintain uniformity in the manner of stating *experimental values*. If we take care to put a definite and unequivocal meaning to experimental data, relieving them as far as possible from ill-defined presumptions, we may be pardoned for continuing to use in descriptive text and in approximate calculations "hydrogen ion concentrations." When we come to exact statements they will be found embodied in pH values of uniform experimental derivation.

In summary then it is suggested that:

1. The following values shall be taken as the *standard* differences of potential, liquid junction potential differences being eliminated, between a tenth normal KCl calomel electrode and a hypothetical hydrogen electrode immersed in a solution normal with respect to the hydrogen ions, under one atmosphere partial pressure of hydrogen, and considered to have zero difference of potential between electrode and solution at all temperatures.

	TEMPERATURE					
	18°	20°	25°	30°	37.5°	40°
Potential difference..	0.3380	0.3379	0.3376	0.3373	0.3364	0 3360

2. The standard experimental meaning of pH shall be the corrected difference of potential between the hypothetical normal

[2] In very many instances constants determined by conductivity methods are employed with precise electrode measurements without any critical examination whatever of their applicability.

hydrogen electrode and the hydrogen electrode under measurement (when this difference is derived by the use of the above values), divided by the numerical quantity 0.000,198,37 T.

3. In every case it shall be specified whether the Bjerrum extrapolation with the use of 1.75N and 3.5N KCl was used to eliminate liquid junction potentials or whether saturated KCl was used and considered to eliminate liquid junction potentials.

There are those who will prefer to use the saturated KCl calomel electrode as a working standard. Its use eliminates the protective devices required to guard the tenth normal calomel electrode against the saturated KCl used as a liquid bridge. Michaelis (1914) has also noted that its temperature coefficient is such that it tends to balance the effect of *fluctuations* in the temperature of a calomel electrode-hydrogen electrode chain. Though there are involved in Michaelis' reasoning some factors which are yet uncertain this advantage may be granted. A practical system which embodies the merits of the saturated calomel electrode and which meets the requirements of the standardization suggested above is illustrated on page 183. In this system the saturated calomel electrode is the working standard whose value is given by careful comparison at known temperatures with a set of tenth normal calomel electrodes.

If any ultimate experimental standard other than the tenth normal calomel electrode be used it is suggested that for the present it be brought into harmony with the above system, which is the system that has practically governed past measurements, and that fundamental revision of any standard *await concerted action based upon thorough investigation of both experimental and theoretical data.*

These suggestions simply put into definite form the current procedure with the recognition on the one hand that the precise use of electrode data involve many theoretical difficulties and on the other hand that the use of such data for the approximate calculation of hydrogen ion concentrations had best be standardized for the sake of uniformity in the records to be handed on to the future.

CHAPTER XX

Supplementary Methods

When any process has been found to be controlled by the concentration of the hydrogen or hydroxyl ions, when the quantitative relations have been established and contributory factors are controllable, there is established a possible means of estimating the concentration of the hydroxyl or hydrogen ions. Many such instances are known. From among them a few may be chosen for their convenience. They are spoken of here as supplementary methods because they are superseded in general practice by indicators and the hydrogen electrode. Several have historical value because they were used in establishing the laws of electrolytic dissociation. Others have value because they are available either for checking the customary procedures or for determinations in cases where there is reason to doubt the reliability of indicator or hydrogen electrode measurements.

An instance of the procedure outlined above is the following. Clibbens and Francis (1912) found that the decomposition of nitrosotriacetonamine into nitrogen and phorone is a function of the catalytic activity of hydroxyl ions. Francis and Geake (1913) then applied the relation to the determination of hydroxyl ion concentrations, Francis, Geake and Roche (1915) improved the technique, and then McBain and Bolam (1918) used the method to check their electrometric measurements of the hydrolysis of soap solutions.

It is just in such checking that the value of these so-called supplementary methods will be appreciated. But, since they will find only occasional use and under circumstances which will require a detailed consideration of their particular applicability, there seems to be no reason to do more than indicate a few of the methods in brief outline.

THE QUINHYDRONE ELECTRODE

We have already seen in Chapter XVI that, when pH is less than about 7, a platinum electrode in the presence of hydroqui-

none and quinone should show a potential difference, which, when referred to the normal hydrogen electrode as a standard may be expressed by the equation

$$E_h = E_k + \frac{RT}{F} \ln [H^+] - \frac{RT}{2F} \ln \frac{[S_d]}{[S_q]} \qquad (66)$$

where E_h is the observed single electrode potential difference, E_k is a constant and $[S_d]$ and $[S_q]$ are the total concentrations of hydroquinone and quinone respectively. We have also previously noted that, under the limitations specified, E_h becomes a linear function of pH when the ratio $\frac{[S_d]}{[S_q]}$ is kept constant and the temperature is constant. (At 30°C, for instance, $\frac{RT}{F} \ln [H^+]$ is -0.06 pH.).

Now quinone and hydroquinone combine in equimolecular proportions to form quinhydrone. (To distinguish this product from similar compounds such as that formed from toluenequinone and toluenehydroquinone it may be called benzoquinhydrone.) In aqueous solutions the reaction is reversible,

$$\text{quinone} + \text{hydroquinone} \rightleftharpoons \text{quinhydrone}$$

and since the solubilities are low, the addition of solid quinhydrone is a convenient way of providing a solution with a mixture of quinone and hydroquinone. We must be careful, however, not to assume that the two are necessarily present in equimolecular concentrations. We may assume that the solid quinhydrone maintains a constant concentration of undissociated quinhydrone in solution. This dissociates and we have the equilibrium condition where D represents hydroquinone, Q quinone and QD quinhydrone:

$$\frac{[Q][D]}{[QD]} = K_1, \text{ and since } [QD] \text{ is constant,}$$

$[Q][D] = K_s$, where K_s is the so-called solubility product. From this it is evident that only the product $[Q][D]$ is kept constant. Ionization of D (hydroquinone) is certainly of fundamental importance as outlined in Chapter XVI and we therefore cannot neglect to consider its effect in the above equation. But

we have already brought the electrode potential equation into such a form and simplified it with the assumption that it is to be used in the region of inappreciable dissociation of D so that we are able at once to say that the very slight ionization of the hydroquinone (D) will not appreciably alter the *ratio* $\frac{[S_d]}{[S_q]}$ from unity. Thus *in acid solutions* the presence of solid quinhydrone maintains a practically constant, unit ratio of its dissociation products. The last term in equation (66) becomes zero, and we have

$$E_h = E_k - 0.000.198\ T\ pH \qquad (67)$$

When E_k has been established a measurement of E_h enables one to calculate pH.

Biilmann (1920) and Biilmann and Lund (1921) have developed the "quinhydrone electrode" for practical use and employ the above equation, derived, however, in another way (assuming the electrode to function as an actual hydrogen electrode. See Chapter XVI).

For the preparation of quinhydrone Biilmann (1921) employed the method of Valeur. Later Biilman and Lund (1921) found it practicable to prepare the quinhydrone as follows:

One hundred grams of ferric ammonium alum in 300 cc. water at 65°C. is turned into a warm solution of hydroquinone in 300 cc. water. The quinhydrone precipitates as fine needles. Cool the mixture in ice and then filter with suction washing the needles three or four times with cold distilled water. Yield, 15 to 16 grams. It is stated that the trace of iron remaining after this process is without serious effect.

To form a "quinhydrone electrode" Biilmann employs a vessel similar to those used for calomel electrodes but with a fairly large platinum electrode (blank platinum). A little quinhydrone is mixed with the *acid* solution under examination, placed in the vessel with the platinum electrode and connected with a saturated or other calomel electrode.

Biilmann determined E_k in equation (67) by simply fixing the pH at a known value with definite buffer solutions and measuring the difference of potential between a quinhydrone electrode in

this solution and a hydrogen electrode in the same buffer without quinhydrone. He gives:

Temperature	E_k
18	0.704
25	0.699

Besides the benzoquinhydrone electrode Biilmann also describes electrodes formed with the xylene and toluene homologues.

Biilmann and Lund describe capillary vessels for use with such electrodes.

Sørensen, Sørensen and Linderstrøm-Lang (1921) discovered that there is a "salt error" with the quinhydrone electrode which becomes very appreciable at salt concentrations of the order of M/5. This they ascribe to an altering ratio of activities for the quinone and hydroquinone with change in salt content.

By methods for the detail of which the reader is referred to the original papers it is predicted that the ratio of the activities of hydroquinone and quinone is defined when the solution is saturated with quinhydrone and one of the components, hydroquinone or quinone; and that under these circumstances there should be less "salt error." There may then be formed what Biilmann and Lund call the hydro-quinhydrone electrode and the quino-quinhydrone electrode.

The hydro-quinhydrone electrode is similar to the quinhydrone electrode described above except that there is present besides solid quinhydrone, solid hydroquinone. At 18°C. the E_k value of this electrode is given by Biilmann and Lund as 0.618.

In the quino-quinhydrone electrode there is present besides solid quinhydrone, solid quinone. At 18°C. the E_k value of this electrode is 0.756. In each case the platinum of these electrodes is positive to the platinum of the hydrogen electrode by the given values.

There are a number of details in the use of these electrodes which require further study and the reader is referred to the original literature for those which have already received attention.

Aside from the great interest of the subject as an example of the general relations pointed out in Chapter XVI the electrodes developed by the Danish investigators should be useful in those cases where the hydrogen of the hydrogen electrode is seriously attacked by the components of a solution. But by the same token

the quinhydrone electrode cannot be used when the reduction potential of a solution is such as to seriously alter the ratio of the hydroquinone and quinone. In either case, however, there remains the possibility of taking advantage of the slowness with which some oxidation-reduction reactions come to equilibrium and experience alone will indicate the limitations of usefulness.

Independently of the Danish investigations Granger and Nelson (1921) worked out some of the relations involved in the quinhydrone electrode.

CONDUCTIVITY

The conductivity of a solution is dependent upon the concentrations of all the ions and upon the mobilities of each. It is therefore obvious that a somewhat detailed knowledge of the constituents of a solution and of the properties of the constituents is necessary before conductivity measurements can reveal any accurate information of the hydrogen or hydroxyl ion concentration. Even when the constituents are known it is a matter of considerable difficulty to resolve the part played by the hydrogen ions if the solution is *complex*. However, the mobilities of the hydrogen and hydroxyl ions are so much greater than those of other ions (see page 163) that methods of approximation may be based thereon. If, for instance, a solution can be neutralized without too great a change in its composition it may happen that with the disappearance of the greater part of the hydrogen ions there will appear a great lowering in conductance. Then, with the appearance of greater hydroxyl ion concentration, the conductance will rise. The minimum or a kink in the curve is a rough indication of neutrality. Thus the conductivity method is sometimes useful in titrations. See Kolthoff for details and references.

The elementary principles of conductivity measurements will be found in any standard text of physical chemistry but the more refined theoretical and instrumental aspects are only to be found by following the more recent journal literature.

CATALYTIC DECOMPOSITION OF NITROSOTRIACETONAMINE

The reaction taking place is represented in outline by the following equation:

$$CO\langle \begin{matrix} CH_2 \cdot C(CH_3)_2 \\ CH_2 \cdot C(CH_3)_2 \end{matrix} \rangle N \cdot NO \rightarrow CO\langle \begin{matrix} CH : C(CH_3)_2 \\ CH : C(CH_3)_2 \end{matrix} + N_2 + H_2O$$

The original quantity of nitrosotriacetonamine is known and the extent of the decomposition at the end of measured intervals of time is measured by the volume of nitrogen evolved.

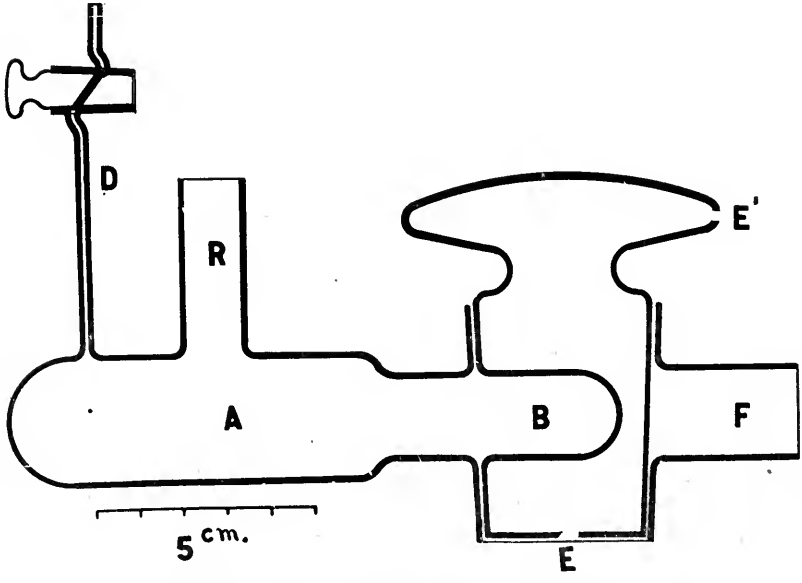

FIG. 41. VESSEL FOR THE CATALYTIC DECOMPOSITION OF NITROSOTRIACETONAMINE

Francis, Geake and Roche (1915) use the vessel shown in figure 41. The tap of the reaction vessel contains a cup B of 7 to 10 cc. capacity into which the alkali or the nitrosoamine can be introduced through F. The solution is then shut in by turning the key through a right angle. The cup becomes a part of the reaction chamber A on turning the key as shown in the figure. The vessel is immersed in a thermostat and shaken during the whole experiment. The holes at E and E′ permit the cup B to be bathed

by the thermostat liquid and so reach thermal equilibrium at the same time as the chamber A. The tube R connects with a constant volume burette where the evolved nitrogen is collected and its pressure read. The tube D is used for washing out the vessel and for filling it with nitrogen when the reaction has to be conducted in an atmosphere free from oxygen.

The unimolecular equation, using the pressure method is

$$k = \frac{2.303}{t} \log \frac{P_\infty - P_o}{P_\infty - P_t}$$

where P_o is the pressure at the time taken as zero, P_t the pressure taken at the time t and $P\infty$ the so-called infinity reading at the end of the experiment. The unit of time taken is the second. At 30°, $\frac{k}{[OH^-]} = 1.92$.

It was found that the constants obtained with nitrosotriacetonamine commence to drift when the ion concentration reaches 0.05N while at 0.35N the drift ceases and the method is again applicable. To bridge the gap it was found that nitroso-vinyl- and isobutyl-diacetonamines could be used.

For temperature coefficients and for the influence of neutral salts etc. the original paper may be consulted.

CATALYTIC DECOMPOSITION OF DIAZOACETIC ESTER

Brodig and Fraenkel (1905) have described the following reaction as applicable to the determination of hydrogen ion concentrations.

$$N_2CH.CO_2\ C_2H_5 + H_2O = N_2 + (OH)CH_2C.CO_2C_2H_5$$

The nitrogen evolved from time to time is measured and the values used in the equation

$$k = \frac{1}{0.4343\ t} \log \frac{a}{a - x}$$

where a is the total gas at the end of the reaction, x the gas after time t minutes and k the reaction constant. At 25°C., $\frac{k}{H^+} = 32.5$.

The method was applied with only partial success by Höber (1900) to blood. Van Dam (1908) used it in the examination of rennet coagulation of milk.

THE INVERSION OF CANE SUGAR

This has been a favorite subject of study by those interested in the catalytic activity of the hydrogen ion. It has been used in a number of instances for the determination of the hydrogen ion concentration of biochemical solutions, but, like all catalytic processes, its close study has revealed a number of complicating factors which necessitate the greatest caution in the interpretation of results.

So numerous are the papers dealing with sugar hydrolysis by acid that the reader is referred to the very thorough review by Woker for the older work. For the more recent investigations see, for example, Jones and Lewis, 1920.

CATALYSES IN GENERAL

Pending further development of the theory of strong electrolytes and of the "salt effect", the investigator, using one or another of the above catalysis methods merely as a check, can place his data upon a reproducible basis by using the following system of comparison. Determine the pH values of a series of buffer solutions lying within the pH range expected of the unknown, and having total salt concentrations comparable to that of the solution to be tested. Under parallel conditions determine the catalytic activity of knowns and unknown. Assume that the result with the buffer agreeing closest to that of the unknown indicates that this buffer and the unknown are at the same pH and check by various modifications of buffer.

MISCELLANEOUS METHODS

Were it worth while there could be compiled under this heading a wide variety of phenomena which have actually been used to determine approximately the hydrogen ion concentration of a solution. We may instance the precipitation of casein from milk by the acid fermentation of bacteria. This has not been clearly distinguished in all cases from coagulation produced by rennet-like enzymes; but, when it has been, the precipitation or non-precipitation of casein from milk cultures has served a useful purpose in the *rough* classification of different degrees of acid fermentation.

In like manner the precipitation of uric acid or of xanthine has been used (Wood, 1903).

The alteration of the surface tension of solutions (Windish and Dietrich, 1919–1921), the distillation of ammonia (Vely 1905), distribution ratios between different solvents, and various other methods have been used to furnish data for the estimation of hydrogen or hydroxyl ion concentrations.

CHAPTER XXI

Applications

Finally, acidity and alkalinity surpass all other conditions, even temperature and concentration of reacting substances, in the influence which they exert upon many chemical processes.—L. J. Henderson.

It is because of the great variety of applications in research, routine and industry that the theories and devices outlined in the previous chapters have been developed. The physical chemist sees in them the instruments of approximation or of precision with which there have been discovered orderly relations of inestimable service to the analyst and with which there have been established quantitative values for affinity or free energy. The biochemist might almost claim some of these methods as his own, not only because necessity has driven him to take a leading part in their development, but also because their application has become part of his daily routine in very many instances.

As mentioned in the preface to the first edition the applications have become so numerous and in many cases so detailed that the time has come for a redispersion among the several sciences of the material that has from time to time been grouped about the activity of the hydrogen ion. This chapter therefore is written only as a cursory review with the hope that it may be of service to the student by revealing the interdependence of specialized lines of research, by suggesting how mistakes still current have been eliminated by those who realize the importance of the subject and by furnishing a rough index to our incomplete bibliography of a voluminous literature.

In the compilation of the bibliography, of which this chapter constitutes an index, no attempt has been made to include all of the very numerous instances in which the activity of the hydrogen or the hydroxyl ions has been found to influence the course of specific chemical reactions, such as the hydrolysis of polysaccharides, special oxidations and condensations, or the nature and accuracy of the numerous color tests used for the qualitative recognition of special chemical groupings. The reader will find in Woker's ex-

tensive monograph, *Die Katalyse*, not only a very complete review of the older, widely scattered literature upon these aspects of hydrogen and hydroxyl ion activity but also an abundance of material which still remains to be reworked with the more modern methods.

In the classification of the bibliography no attempt has been made to place the references in strictly logical catagories, nor has it been practical to make a minute subdivision by subjects with numerous cross references. The grouping is by subjects which are of particular current or historical interest or which fall within the provinces of special branches of science.

General Reviews. Excellent general reviews of biochemical applications are Sørensen's article in *Ergebnisse der Physiologie*, 1912, and Michaelis' monograph *Die Wasserstoffionenkonzentration*, 1914. As we go to press there comes to hand the first part of the 1922 revised edition of this excellent monograph. This first part covers in extended form the theoretical foundations briefly treated in the first edition and deals in more or less detail with many subjects briefly touched upon in the following pages. Prideaux has compiled a great deal of valuable data in *The Theory and Use of Indicators*, London, 1917. In this English work will be found the more important matter which Bjerrum (1914) embodied in his monograph on the theory of titration and which Noyes had previously summarized in his paper "Quantitative application of the theory of indicators to volumetric analysis," (1910). The analyst will find a wealth of helpful suggestions in Stieglitz' *Qualitative Analysis*. A review of the indicator method which is of some general interest, although written specially for the bacteriologist, will be found in *The Journal of Bacteriology*, **2**, nos. 1, 2 and 3 (Clark and Lubs, 1917).

Those who desire to review the theory of electrolytic dissociation with special reference to its bearing on electrode measurements will find useful LeBlanc's *Text Book of Electrochemistry* (1907).

Among several papers which may be called classics in biochemistry there will be recognized the preëminence of Sørensen's *Études enzymatiques*, II, from the Carlsberg Laboratory in Copenhagen and *Das Gleichgewicht zwischen Basen und Säuren im tierischen Organismus* by Henderson of Harvard.

The Theory of Titration is so closely allied with the more

general applications of indicators and the hydrogen electrode that it may well be taken from the alphabetic arrangement to be followed and treated before taking up some general considerations.

The stress which has come to be laid upon that factor of "acidity" with which we have been dealing should not detract from the true importance of the estimation of total acidity or alkalinity by titration.

But the theory of titration is only a special form of the theory with which we have been concerned up to this point; so that we are prepared to sketch in outline those salient features of the well-ordered theory which has displaced the loose empiricism of other days.

In figure 42 are shown the titration curves of hydrochloric, acetic and boric acids, determined as outlined in Chapter II. The ordinates of figure 42 are pH values and the abscissas cubic centimeters of N/10 NaOH added to 10 cc. N/10 acid. At the side of the main part of the figure are representations of the color transformations of two indicators (see Chapter IV).

Although the indicator curves are drawn at one side of the figure the reader will readily see from the theory described in Chapter IV that they could have been placed in the main figure parallel to the titration curves if the abscissas had been made percentage neutralization.

A more complete picture of the conditions of titration would be shown had the curves been extended to indicate what happens when the "end-points" are overstepped. The reader may picture this for himself by imagining that the curve for boric acid continues with the slope shown at 11 and then flattens out between 12 and 13, and that the other curves, after passing pH 10, sweep to the right to join the extended boric curve.

When all but a very small part of the hydrochloric acid has been neutralized there comes a sharp break in the titration curve. On the addition of the last trace of alkali required for complete neutralization the pH of the solution plunges to the alkaline region. In this precipitous change the pH passes the range of methyl red, and, with an amount of alkali that will be detected only by careful observation, it passes into that range of pH where phenolphthalein shows its various degrees of color. Therefore, with the exclusion of carbon dioxid, either indicator may be used to indicate the "end

point" of this titration. The case is very different in the titration of acetic acid. Here we have an acid whose dissociation constant (see Chapter I) is so low that the flat portion of the titration curve lies in that region of pH where methyl red shows its various de-

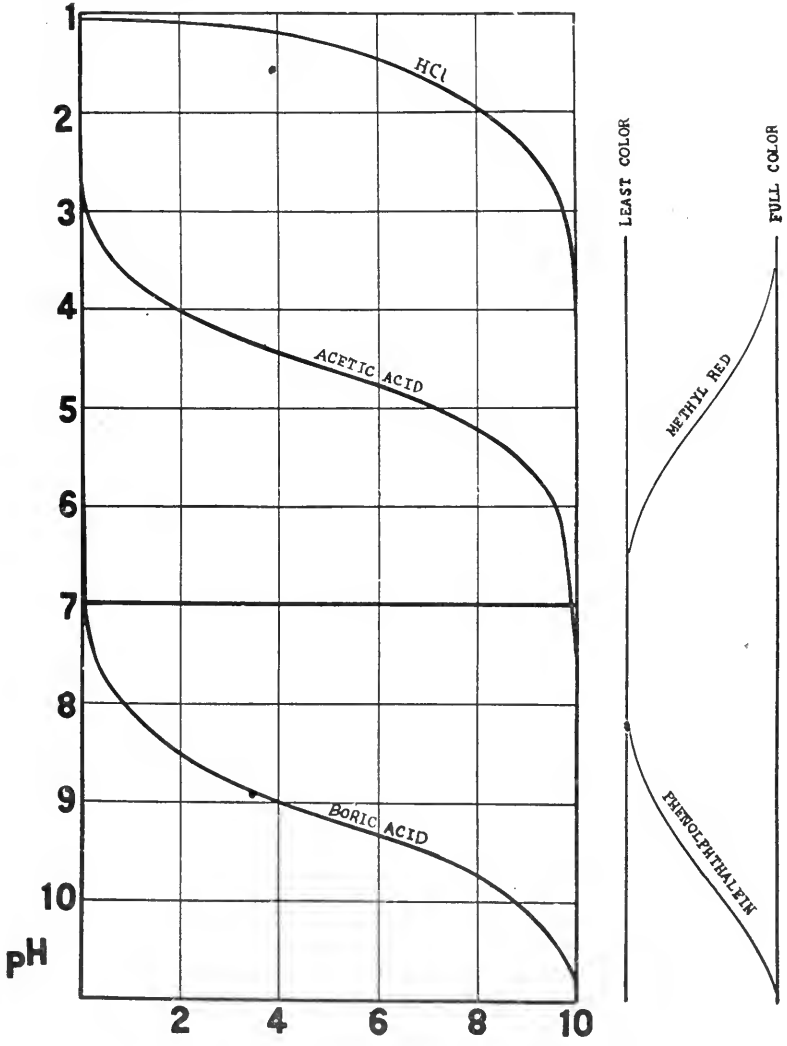

FIG. 42. TITRATION CURVES OF 10 CC. N/10 ACIDS WITH N/10 NaOH

grees of color. In other words the apparent dissociation constant of methyl red is not far from that of acetic acid. Therefore, as the titration of acetic acid proceeds, and long before the neutralization of the acetic acid is complete, methyl red has been partially transformed and at last is so extensively transformed that no marked change of color is observed when the pH of the solution abruptly changes with complete neutralization of the acetic acid. It is at once evident why an indicator with the properties of phenolphthalein must be used in such a case. In the titration of a still weaker acid, such as boric acid, phenolphthalein becomes comparable to methyl red in the latter's conduct in acetate solutions. To titrate boric acid it must be combined with glycerine or mannitol to form a stronger acid. See Liempt (1920).

The titration curve of boric acid is representative of the conduct of many of the weak acidic groups found in the substances of biochemical interest.

Sometimes by a judicious selection of indicators it is possible to titrate in succession a mixture of two acids. For instance A. B. Clark and Lubs (1918) have called attention to the advantages of the two color transformations of thymol blue. The color transformation of thymol blue in the acid range is such that it may be used to indicate the approximate end point of hydrochloric acid in the presence of acetic acid; and the second color change occurs in a region of pH such that it will indicate the end point in the titration of the acetic acid. A. B. Clark and Lubs (1918) and Lubs (1920) have examined other similar uses of this indicator.

The principles thus briefly outlined apply to the titration of bases with strong acids, but, of course, with the direction of pH change reversed and with the end points tending to lie on the acid side of pH 7.0. A hydrogen ion concentration of 10^{-7}N or pH 7.0 is called the neutral point because it is the concentration of both the hydrogen and the hydroxyl ions in pure water; but it is evidently seldom the practical or even the theoretical point of neu trality for titrations.

As phenolphthalein is the more generally useful indicator for the titration of acids with strong bases so is methyl red the more generally useful indicator in the titration of bases with strong acids. Each fails, however, when the acid or base is very weak, and each may be replaced by a more suitable indicator in special cases.

For the treatment of these cases the reader should consult the detailed description of the theory of titration in one of the papers mentioned above.

Where high color or turbidity interferes with the use of indicators in titration the hydrogen electrode is often useful. See Böttger (1897), Hildebrand (1913) Michaelis (1917). Since it may be necessary only to detect the "break" in the titration curve, the hydrogen electrode system and potentiometer system used for this purpose may be very simple. The hydrogen electrode has the advantage that it may often be used where colorimetric tests are impracticable and that it may be linked electrically with automatic regulating and recording instruments such as Leeds and Northrup Company have devised for industrial use.

Pinkhof (1919) has suggested special half-cells with single potentials equal to those of the end-points of titrations, thereby eliminating the necessity of a potentiometer. A galvanometer or electrometer indicates equalization of potentials and hence the attainment of the "end-point."

In like manner one may use two hydrogen electrodes as described in Chapter IX. If one electrode is immersed in a solution having the pH of the desired end-point, the attainment of this end-point in the other solution is indicated by the point of reversal of current in the galvanometer (Klopsteg, 1921).

Since titrimetric determination of total acidity or basicity involves one or another method of estimating pH, the understanding of the principles involved is essential to an intelligent interpretation of the values obtained in the titration of complex mixtures. In a great many instances there have been carried over to the titration of complex mixtures the rule-of-thumb method and the special interpretation first worked out by the analyst for the titration of strong acids and bases. Now it not infrequently occurs, especially among extracts of natural products, that there are present a variety of weak acids and bases; and no precipitous drop in the titration curve can be observed in the pH zones covered by the indicators very generally employed in such titrations. The situation is comparable with an attempt to titrate boric acid with phenolphthalein as indicator. No sharp "end-point" is observable. But there will always remain the distinctive value of a titration and wherever this cannot be precisely analyzed it should be stated in simple straightforward terms.

In the majority of cases the titration of such solutions reduces to a mere revelation of differences in total buffer action furnishing but one definite point on the titration curve. The procedure often followed is comparable with the practice of the ancient Romans who, according to Trillat (1916), (cf. Stephanides 1916) titrated natural waters with drops of red wine. While modern standards of concentration are more exact than the wine standard of the Romans their significance is largely lost by a choice of indicators as accidental as the Roman choice of the coloring matter of red wine. The frank admission that the content of acids in some complex solutions cannot be determined by titration need not destroy the value of the information gained by a titration if this information be correctly used. But too often the matter is carried to an extreme. In the routine methods for titrating milk a perfectly simple test has been so elaborated that it not only has become confusing to the chemist but so misleading to the creamery man that it is causing large economic losses. Often the initial pH of a solution is of greater significance than is the titration value obtained after juggling the solution with acid or alkali. Illustrations of this are to be found in the author's treatment of bacteriological culture media (Clark, 1915).

Having followed some of the salient features of titration and found this procedure linked with the more general aspects of hydrogen ion determinations the reader is reminded of those relations among acids and bases outlined in Chapter I which point to certain general considerations.

General Considerations. As a comprehensive generalization it may be said that the hydrogen ion concentration of a solution influences in some degree every substance with acidic or basic properties. When we have said this we have said that the hydrogen ion concentration influences the great majority of compounds, especially those of biochemical interest. Such a generalization, however, would be misleading if not tempered by a proper appreciation of proportion. Rarely is it necessary to consider the ionization of the sugars since their dissociation constants are of the order of 10^{-13} and their ionization may be generally neglected in the pH region usually encountered in physiological studies. Likewise there are zones of pH within which any given acidic or basic group will be found in dilute solution to be in a practicaliy undissociated or fully dissociated state. Perhaps there is no more vivid way of illustrating this than by a contemplation of the conduct of indicators. Above a certain zone of hydrogen ion concentration phenolphthalein solutions are colorless. Below this zone (until

intense alkalinity is reached) only the colored form exists. Within the zone the *virage* of a phenolphthalein solution is intimately related to the hydrogen ion concentration. The conduct of phenolphthalein, which happens to be visible because of tautomeric changes which accompany dissociation, is a prototype of the conduct of all acids. Just as we may suppress the dissociation of phenolphthalein by raising the hydrogen ion concentration of the solution so may we suppress the dissociation of any acid if we can find a more intensely ionizing acid with which to increase the hydrogen ion concentration of the solution. Similar relations hold for bases, and, if we regard methyl red as a base, we may illustrate with it the conduct of a base as we illustrated the conduct of an acid by means of phenolphthalein.

Such illustrations may serve to emphasize the reason underlying the following conclusion. Whenever, in the study of a physiological process, of a step in analysis requiring pH adjustments or of any case involving equilibria comparable with those mentioned above, there is sought the effect of the pH of the solution, it may be expected that no particularly profound effect will be observed beyond a certain zone of pH. Within or at the borders of such a zone the larger effects will be observed. From this we may conclude that the methods of determining hydrogen ion concentrations should meet two classes of requirements. In the first place, when the phenomenon under investigation or control involves an equilibrium which is seriously affected by the pH of the solution, the method of determining pH values should be the most accurate available. In the second place, when the equilibrium is held practically constant over a wide range of pH, an approximate determination of pH is sufficient and refinement may be only a waste of time.

Neglecting certain considerations which often have to enter into a choice of methods it may be said that the electrometric method had best be applied in the first case and the indicator method in the second. When the nature of the process is not known, and it therefore becomes impossible to tell *a priori* which method is to be chosen, the colorimetric method becomes a means of exploration and the electrometric method a means of confirmation.

Exception will be taken to this statement as comprehensive for there are cases where one or the other method has to be

discarded because of the nature of the solution under examination. Nevertheless, in general, the utility of the colorimetric method lies in its availability where approximations are needed and exact determinations are useless and also in its value for reconnaissance; while the value of the electrometric method lies in its relative precision.

In some instances the qualitative and quantitative relations of a phenomenon to pH should be carefully distinguished. Note, for instance, the significance of an optimum or characterizing point. Consider the conduct of phenol red and of cresol red. These two indicators appear to a casual observer to be very much alike in color and each exhibits a similar *virage* in buffer solutions of pH 7.6, 7.8, etc. Careful study, however, shows that each point on the dissociation curve of phenol red lies at a lower pH than the corresponding point on the dissociation curve of cresol red. If the half transformation point be taken as characteristic it may be used *to identify* these two indicators. Likewise it is the *dissociation constant* of an acid or a base, the *isoelectric point* of a protein, the *optimum pH* for acid agglutination of bacteria, or an optimum for a process such as enzyme activity that furnishes *characteristic* data.

When there is observed a correlation between pH and some effect, the mere determination of pH alone will of course throw but little light upon the real nature of the phenomenon except in rare instances. Determination of the hydrogen ion concentration will not even distinguish whether a given effect is influenced by the hydrogen or the hydroxyl ions, nor will it always reveal whether the influence observed is direct or indirect. It is true, however, that, even when the hydrogen ion concentration is effective through remote channels, it may be very important. Therefore advantage should be taken of the comparative ease with which the concentration of hydrogen ions may be determined or controlled and its influence known or made a constant during the study of any other factor which may influence a process. From this point of view methods of determining hydrogen ion concentration take their place beside thermometers. and buffer mixtures beside thermostats.

Indeed it may be said that the failure to take advantage of buffers is still a prolific source of error in the experimental work

of every branch of science having to do with solutions. In one case the neglect is gross; in another case it may be a perfectly excusable misjudgment. A complete understanding of the effects of the hydrogen or hydroxyl ion is very far from attainment and those who faithfully control the pH of their solutions are often rewarded by the most surprising results. To emphasize this aspect we may call attention to the fact that while the dissociation of glucose is quite negligible in the region of pH 7 so far as any appreciable effect upon the displacement of other acid-base equilibria is concerned, the converse proposition is decidedly not negligible. A shift in pH from 7.0 to 7.4 has a very marked influence upon the conduct of glucose in heated solutions as every media maker knows. Nor may it be forgotten that there are many compounds only the main dissociation constants of which have been determined; until we know the values of secondary acidic or basic dissociations, we have not a complete description upon which to base judgment of the conduct of such compounds in relation to pH.

It is the opinion of the author that altogether too much emphasis has been placed upon the so-called "neutral point." The relation $[H^+]\,[O\overline{H}] = K_w$ holds all along the scale. The equality $[H^+] = [O\overline{H}]$ or pH = pOH occurs at pH 7. This is a convenient reference point and has been seized upon as the point of division in our habitual ideas of "acidity" and "alkalinity." But pH 7 is not used as the end point in titrations, it is not the neutral point in the conduct of ampholytes or selectively adsorbing material, and seldom is anything unique seen to happen when in a series of experiments a solution "crosses the line."

Living cells are dependent upon the maintenance of a strictly limited hydrogen ion concentration in their environment. The recognition of this as a fact, independently of any theory whatever regarding the channels of influence, has brought hydrogen ion methods into the culture laboratory and into the garden. Accustomed as we are to dealing with ponderable quantities of material we are sometimes startled by the fact that a cell is dependent upon the maintenance of an environment varying between the limits 0.000,001 and 0.000,000,01 gram hydrogen ions per liter. Sometimes the permissible limits are even closer but the order of magnitude remains the same. Such values, however, do not

represent entities separable from the other material present in solution. They represent only a position of balance among relatively *large quantities* of material containing a reserve of potential hydrogen ions.

Now that N, the number of molecules of solute present per litre in a molar solution, is accurately known (Millikan), it is certain that even in a solution having a hydrogen ion normality as low as 10^{-13} there are about 10^{10} hydrogen ions per litre. This estimate, when taken in conjunction with the electrical charge associated with each ion, may indicate how it is that a normality of 10^{-13} H^+ may be detected.

But there still remains the fact that this normality is very low in comparison with the other material present even in distilled water. In solutions heavily buffered at pH 13 we find the hydrogen electrode or an acid indicator *rigidly stabilized* in its conduct and it is questioned whether this can be brought about by such extreme relative dilutions of the hydrogen ions alone. Keller (1921) has expressed doubt of another sort. He calls attention to the diminutive size of the hydrogen ion (allowing for hydration) compared with a giant protein molecule, and, picturesquely proportioning the one to the other as a bacterium to a Mont Blanc, he questions the influence upon the protein which is attributed to the hydrogen ion.

All these are "sharp-toothed questions" which, were they "baited with more skill, needs must catch the answer." In many of the answers given, however, there lies an easily detected fallacy. Our present convenient modes of formulating relations are regarded as complete pictures of the physical facts and as such are followed to the bitter end with disastrous results. In a previous chapter we have attempted to broaden the outlook just a little, and have suggested that in many cases a more complete formulation of relations would show that as the physical effectiveness of one ion fades out at extreme dilution other components of the solution maintain the continuity. From this point of view even the more extreme "calculation values" retain a definite significance.

In like manner an extreme hydrogen ion concentration may be significant as an *index* of the state of an equilibrium with which the hydrogen ion itself has little actual physical significance. Its introduction as a component of the equilibrium is a *convenient*

and at the same time a stoichiometrically true and mathematically correct mode of expression containing no implications regarding the actual physical effectiveness of a small hydrogen ion concentration as an individual quantity separable from the other components of a solution. At higher concentrations there can be little doubt of the physical effectiveness of the hydrogen ions whatever their size, or energy relative to other bodies. The energy placed on the grid of an electron tube may be small, but the potential of the grid may determine a large flow of energy between filament and plate. The hydrogen ions in a solution may be small in relative size or relative numbers, but they may control the mobilization of a large reserve. If one seeks to go further, perhaps to formulate a more fundamental basis, he still has to conform to the experimental data at hand.

These data are too extensive, too detailed and altogether too complete to admit any doubt of the pragmatic value of those measurements we now customarily express in terms of hydrogen ion concentration or activity. Such values do indicate definite positions of equilibrium among important components of a solution and they have oriented relations hitherto unsuspected. But it is by no means certain that we have attained the ultimate conception of what our measurements represent in terms of mechanisms. Better descriptions of these we eagerly await. Scientific thought pauses where it is convenient and leaps forward when necessity demands; but experimental measurements remain with whatever force skill, scope and instrumental precision give them—requiring only reinterpretation with the enlargement of vision.

In a crude way we have attempted in a previous chapter to give a generalized picture of oxidation-reduction relations. Here we encounter definite experimental facts which it is sometimes convenient to express in terms of "calculation values." It may now fairly be asked whether these are not significant as indices of equilibria of as much importance to the delicate adjustments of life processes as are hydrogen ion concentrations. If the studies so far made are prophetic there will be found not only a profound interrelationship between hydrogen ion concentrations and oxidation-reduction equilibria but also direct control of certain biological processes by the reduction potential of the medium. See Gillespie (1920), Clark (1920) and Clark and Cohen (1922)

for some applications in bacteriology. See also Chapters XVI and XX.

Adsorption. Hydrogen and hydroxyl ions are particularly subject to adsorption upon surfaces. Since the relative activities of these ions are especially easy to measure, methods of determining pH are of great value for adsorption studies. For a review of recent work see Michaelis (1922).

References. Lachs-Michaelis (1911), Löffler-Spiro (1919), Michaelis (1922), Michaelis-Rona (1910, 1919, 1920), Rona-Michaelis (1919, 1920), Tanner (1922).

Analyses. The empiricism that characterized the development of analytical methods in the hands of Fresenius and others left specifications for the use of mixtures of acids, such as acetic, and their alkaline salts in many separations. These we now know control the hydrogen ion concentration. Here and there in the special literature are to be found the calculated hydrogen ion concentrations in such cases and in other cases directions which are somewhat more precise than the customary "slightly acid" or "slightly alkaline." More recently there has been undertaken direct experimentation with hydrogen electrode or indicator methods. The need of further development was voiced some years ago by Dr. Hillebrand of the Bureau of Standards when he indicated to the Washington Chemical Society the need of a systematic investigation of all analytical methods. One type of information urgently needed may be learned from the papers of Blum, of Fales and Ware and of Hildebrand. Colorimetric pH measurements on carbonate equilibria are furnishing valuable information in several simple analytical methods. Kolthoff is working on the relation of pH to certain oxidation-reduction titrations. Many qualitative color reactions remain to be studied.

References. Anger (1921), Behrend (1893), Bishop-Kittredge-Hildebrand (1922), Bogue (1922), Böttger (1897), Brønsted (1911), Blum (1913, 1914, 1916), Eastman-Hildebrand (1914), Fales-Ware (1919), Garard-Sherman (1918), Haas (1916), Hanzlik (1920), Haskins-Osgood (1920), Hildebrand (1913), Hildebrand-Bowers (1916), Hildebrand-Harned (1912), Hopkins (1921), Kober-Haw (1916), Kober-Sugiura (1913), Kolthoff (1919–1921), Kolthoff-Volgelenzang (1921), Koritschoner-Morgenstern (1919), Kramer-Green (1921), Kramer-Tisdale (1921), Liempt (1920),

Lizius (1921), Marriott (1916), Mattick-Williams (1921), Menten (1920), Oettingen (1900), Osterhout (1918), Robinson (1919, 1922) Robinson-Bandemer (1922), Shohl (1922), Sollmann (1920), Swanson-Tague (1919), Tague (1920), Tillmans-Bohrmann (1921), Tizard-Whiston (1920), Zoller (1920).

Autolysis of tissue is governed by the activity of enzymes which are sensitive to the concentration of hydrogen ions. As the resultant of the activity of two types of enzymes (Dernby) autolysis is controlled by the pH which brings into play the activity of each.

References. Bradley (1916), Bradley-Felsher (1920), Bradley-Taylor (1916), Dernby (1917–1918), Gibson-Umbreit-Bradley (1921), Koehler-Severinghaus-Bradley (1922), Morse, M. (1916–1917).

Bacteriology. A review of the applications in bacteriology up to 1917 is given by Clark and Lubs (1917).

Adjustment of the reaction of media by the old titrimetric procedure was criticised by Clark (1915), and, on the introduction of suitable indicators and the evidence for the advantage of adjusting on the pH basis, the titrimetric method has been abandoned for more significant and easier modern methods. Studies on growth optima (which see below) have shown that for the cultivation of most saprophytes approximate indicator control without the use of standards is sufficient (see Chapter VIII). For special purposes and especially for the study of certain important pathogens it is well to adjust with the precision attained with standards. Seldom is electrometric control necessary.

References. Adam (1921), Baldwin (1919), Barthel (1918-20), (1920), Bovie (1915), Clark (1915), Clark-Lubs (1916), Conn (1919), Cox-Wood (1920), Davis (1920), Dernby (1919), Fennel-Fisher (1919), Foster-Randall (1921), Grace-Highberger (1920), Henderson-Webster (1907), Hurwitz-Meyer-Ostenberg (1915–1916), Jones (1919), Kligler (1917-1918), Kligler-Defandorf (1918), Küster (1921), McIntosh-Smart (1920), Massink (1921), Medical Research Committee (1919), Michaelis (1921), Norton (1919), Ponselle (1920), Reitstötter (1920), Stickdorn (1922), Wolf-Shunk (1921).

The optimal zones and the limits of growth and general metabolism have naturally been the chief interest in the first surveys of the

influence of hydrogen ion concentration upon bacterial activity. It is now clear that in the future more exact studies will have to differentiate between optimal initial pH, optimal zones of growth, optimal zones for general or special metabolism, optimal zones for preservation, etc. The self limitation of acid fermentation, first clearly defined by Michaelis and Marcora (1912), has been applied to certain practical tests; for example see Clark (1915), Avery and Cullen (1919). pH limits for special organisms which have commercial significance are exemplified by control of "rope" in bread (Cohn-Walbach-Henderson-Cathcart) and "scab" on potatoes (Gillespie-Hurst).

References. Adam (1921), Allen (1919), Avery-Cullen (1919), Ayers (1916), Ayers-Johnson-Davis (1918), Barthel (1918), Barthel-Sandberg (1919), Beckwith (1920), Bengtson (1922), Boas (1920), Boas-Leberle (1918), Brown-Orcutt (1920), Bunker (1919), Chambers (1920), Cheplin-Rettger (1920), Clark (1915–18) Clark-Lubs (1915–1917), Cohen-Clark (1919), Cohn-Walbach-Henderson-Cathcart (1918), Cole-Onslow (1916), Cole-Lloyd (1917), Colebrook (1920), Cullen-Chesney (1918), v. Dam (1918), Dernby, (1921), Dernby-Avery (1918), Dernby-Blanc (1921), De Kruif (1922), Duggar-Severy-Schmitz (1917), Erickson-Albert (1922), Euler-Emberg (1919), Euler-Heintze (1919), Evans (1918), Foster (1920–1921), Freear-Venn (1920), Fred-Davenport (1918), Frothingham (1917–1918), Gainey (1918), Gates (1919), Gillespie (1918), Gillespie-Hurst (1918), Grace-Highberger (1920), Hägglund (1915), Hall-Fraser (1921–1922), Henderson (1918), Holm-Sherman (1921–1922), Huddleson (1921), Itano (1916), Itano-Neill (1919), Itano-Neill-Garvey (1920), Johannessohn (1912), Johansen (1920), Jones (1920), Kiesel (1913), Kligler (1918), Kligler-Robertson (1922), Kohman (1919), Kniep (1906), Lazarus (1908), Levine (1920), Lord (1919), Lord-Nye (1919), Lloyd (1916), Lüers (1914), Meacham (1918), Mellon (1921), Meyerhof (1916-1917), Michaelis-Marcora (1912), Scheer (1921), Schoenholz-Meyer (1919–1921), Shaw-Mackenzie (1918), Sherman (1921), Shohl-Janney (1917), Somogyi (1921), Steinberg (1919), Svanberg (1918–21), Swartz (1920) Swartz-Shohl-Davis (1921), Waksman (1918), Waksman-Joffe (1920–21), Williams-Povitzky (1921), Winslow-Kligler-Rothberg (1919) Wolf (1918), Wolf-Foster (1921) Wolf-Harris (1917), Wolf-

Shunk (1921), Wolf-Telfer (1917), Wright (1917), Zeller-Schmitz (1919).

The influence of pH upon bacterial metabolism. The reaction of the medium even within the zone of optimal bacterial growth is found to influence either the rate, or the relative rate of specific types of metabolism. Not only the activity but also the production of enzymes is influenced and the production of special products such as toxins is partially controlled by the pH of the medium.

References. Arzberger-Peterson-Fred (1920), Avery-Cullen (1920), Atkin (1911), Barthel (1921), Barthel-Bengtsson (1920), Barthel-Sandberg (1920), Blanc-Pozerski (1920), Boas (1919), Bronfenbrenner-Schlesinger (1918), Brooks (1921), Bunker (1919), Charpentier (1921), Clark (1920), Cook-Mix-Culvyhouse (1921), Davis (1918, 1920), Dernby-Aleander (1921), Dernby-Blanc (1921), Dernby-David (1921), Euler-Blix (1919), Euler-Emberg (1919), Euler-Hammarsten (1916), Euler-Svanberg (1918, 1919), Fred-Peterson (1920), Gaarder-Hagem (1920–1921), Green (1918), Gröer (1912) Gustafson (1920), Itano (1916), Jacoby (1918), Jones (1920), Lord-Nye (1919), Meyerhof (1917), Neuberg-Hirsch (1919), Northrop-Ash-Senior (1919), Patty (1921), Peterson-Fred-Verhulst (1921), Robinson-Meader (1920), Sasaki (1917), Stevens-Koser (1920), Venn (1920), Waksman-Joffe (1921), Wolf (1920), Wyeth (1919).

Disinfectant action of acids and bases is certainly in large measure a function of hydrogen or hydroxyl ion concentration; but specific effects of certain acids and bases, which were suspected before, have now been more clearly demonstrated by the use of hydrogen ion methods. With the conductivity method Winslow and Lochridge were able to show the effect of the hydrogen ion in simple solutions and predicted relations which more powerful methods have extended to complex media.

Cohen (1922) has reviewed certain relations between pH and viability of bacteria under sub-lethal conditions. Time, temperature, and pH are now linked as controlling factors in canning.

The more direct action of hydrogen ion concentration upon cells must be distinguished from its control upon the effective state of a toxic compound. Knowledge of pH effects is therefore essential to the assay of disinfectants and to the advancement of chemotherapy.

References. Aubel (1920), Bettinger-Delaval (1920), Bial (1902), Bigelow (1921), Bigelow-Cathcart (1921), Bigelow-Esty (1920), Browning-Gulbransen (1921), Browning-Gulbransen-Kennaway (1919), Clark, J. F. (1899), Clark-Lubs (1917), Cohen (1922), Cohen-Clark (1919), Donk (1920), Friedenthal (1919), Krönig-Paul (1897), McClelland-Waas (1922), Müller (1921), Neilson-Meyer (1921), Norton-Hsu (1916), Paul-Birstein-Reuss (1910), Paul-Krönig (1896), Rideal-Evans (1921), Shohl-Deming (1921), Tawara (1921), Traube-Somogyi (1921), Vermast (1921), Waterman (1915), Weiss (1921), Winslow-Lochridge (1906), Wolf-Foster (1921), Wright (1917). See also "Pharmacology."

Acid agglutination of bacteria, first definitely recognized by Michaelis (1911) in its relation to hydrogen ion concentration, has been found to be of some diagnostic use. The discovery by Arkwright of separately agglutinable constituents opened up some investigations of possibly wide bearing. Buchanan has indicated some of the possible relations to serum agglutination.

References. Arkwright (1914), Bach (1920), Barendrecht (1901), Bechhold (1904), Beintker (1912), Beniasch (1912), Bergey (1912), Bondorf (1917), Buchanan (1919), De Kruif (1922), Eisenberg (1919), (contains review and bibliography), Field-Teague (1907), Georgi (1919) Gieszczykiewicz (1916), Gillespie (1914), Grote (1913–1914), Heimann (1913), Jaffé (1912), Kemper (1916), Krumwiede-Pratt (1913), Tiess (1919), Markl (1915), Michaelis (1911, 1915, 1917), Michaelis-Adler (1914), Murray (1918), Poppe (1912), Radsma (1919), Schidorsky-Reim (1912), Sears (1913), Sgalitzer (1913), Tulloch (1914).

d'Herelle phenomenon. Gratia (1921).

Cell interior. Angerer (1920).

Testing fermentation. See various references under other headings and especially Baker (1922), Chesney (1922), Clark (1915–17), Clark-Lubs (1917), Laybourn (1920), Nichols-Wood (1922).

Balloelectricity.

Reference. Christiansen-Christiansen (1919).

Beer. As originally outlined by Pasteur the "reaction" of wort has much to do with the brewing of beer. The control of "disease" and of the protein material held in solution is to some extent dependent upon pH as are the activities of the enzymes concerned at each stage.

References. Adler (1915, 1916), Emslander (1914–1919), Leberle-Lüers (1914), Lüers (1914), Lüers-Adler (1915), Schjerning (1913). See also "Bacteriology," "Enzymes" and "Proteins."

Blood. The hydrogen ion concentration of the blood, while varying slightly among normal individuals, is regulated with remarkable constancy in any one individual in a normal environment. It never varies far from pH 7.4. Van Slyke (1921), places the normal variation between about 7.3 and 7.5 and the limits compatible with life at approximately 7.0 and 7.8. Since the bicarbonate-carbonic acid equilibrium is one of the most important in the regulation of the blood's reaction it is convenient to define the system in terms of this equilibrium. See "carbonate equilibrium" for the derivation of the relation

$$pH = pK_1 + \log \frac{[H\overline{C}O_3]}{[\text{free } CO_2]}$$

Inspection of the relations involving the carbonate ion $\overline{C}\overline{O}_3$ (see page 320) will show that at pH 7.4 $[\overline{C}\overline{O}_3]$ may be neglected and the fixed carbon dioxid may be regarded as entirely bicarbonate. The extent of the bicarbonate dissociation is in doubt but if we substitute $[BHCO_3]$, for $[H\overline{C}O_3]$ where B represents any monovalent base, and modify pK_1 to accord with the experimental conditions, we have

$$pH = 6.1 + \log \frac{[BHCO_3]}{[\text{free } CO_2]}$$

The ratio $\frac{[BHCO_3]}{[\text{free } CO_2]}$ determines pH. Normally it is about $\frac{20}{1}$.

From one point of view the blood may be regarded as a scavenger, burning the waste products in the tissues it perfuses, and carrying off the final products of combustion of which CO_2 is one of the most important for the acid-base equilibria under consideration. With a given content of buffer in the blood the hydrogen ion concentration would be maintained constant under this inflow of CO_2 by the maintenance of a constant CO_2 pressure in the lungs; but with varying buffer content the hydrogen ion concentration could only be maintained constant by a mechanism directly responsive to hydrogen ion concentration and capable of altering the CO_2 pressure. It seems that the respiratory

centre is thus directly responsive to the hydrogen ion concentration and by its regulation of the breathing maintains in the alveolar air that level of CO_2 pressure which is in harmony with the equilibria centered about constant pH under varying conditions. Of this Haldane says: "The respiratory centre is enormously more delicate as an index of change in hydrogen ion concentration of the blood than any existing physical or chemical method." Clinical methods based on the measurement of the alveolar CO_2 tension are now extensively used (see Van Slyke). On the other hand, the CO_2 tension is but one item of a complicated set of equilibria. It often becomes of importance to know the relative proportions of the other constituents of the acid-base equilibria. In pathological conditions the oxidative processes may be at fault and the carbonate equilibria must be adjusted to accommodate the products of incomplete combustion in the effort of the body to maintain constant hydrogen ion concentration in the blood. Therefore it becomes important to learn the relation of the CO_2 content to the alkaline reserve. When this is done by gas chain or indicator titrations the hydrogen electrode and indicator methods again enter the subject from which they were to some extent displaced when it was found that there was no particular object in studying a constant maintained physiologically with a degree of precision often beyond the precision of experimental measurement.

Although it is convenient to express the acid-base equilibria of the blood *in terms of* the bicarbonate system other equilibria are of equal importance to a complete description of the mechanisms. In the plasma are other substances beside the carbonic acid and bicarbonate which participate in the acid-base equilibrium; but the most interesting relations are found in the Donnan equilibrium (see page 328) between the solutes of the plasma and the material trapped within the membranes of the blood cells. Of this material the blood pigment is the most important. When oxidized (as oxyhemoglobin) it is more strongly acidic than when reduced (as hemoglobin). The direct consequence is this: when the blood pigment gives up oxygen to the tissues the blood assumes more basic properties as a whole and is thus able to take up more CO_2 for a given displacement of pH. The converse change occurs on oxidation in the lungs, and tends to

displace CO_2. In this sense the blood pigment is a carrier of CO_2 as well as a carrier of oxygen.

Intimately connected with the regulation of the hydrogen ion concentration of the blood are the functions of the kidneys (see Cushny). By their action there are eliminated the non-volatile products of metabolism, several of which are of great importance for the acid-base equilibria of the blood. The colorimetric determination of the pH of the urine is a comparatively simple procedure which furnishes valuable data when properly connected with other data. (See for instance Blatherwick, and the works of Henderson, of Palmer and of Van Slyke.)

While the greatest interest has centered in the subjects briefly mentioned above, there remain innumerable other problems of importance. Of these there may be mentioned the relation of the pH of the blood to the calcium-carrying power, to the activity of various enzymes, to the permeabilities of tissue membranes, to the activity of leucocytes, and to various reactions used in the serum diagnosis of disease.

The student, if bewildered by the array of references given below, will find it profitable to read the classic work of Henderson, *Das Gleichgewicht zwischen Basen und Säuren im tierischen Organismus*. By following the papers of Van Slyke and his coworkers the student will find reviews of various aspects of the subject. The respiration phase so far as the older work is concerned will be found in Barcroft's monograph. The later work which includes the effects of pH is reviewed by Bayliss, Henderson, Parsons and others. Van Slyke's *The Carbon Dioxide Carriers of the Blood* (1921) reviews the acid-base equilibria of the carbonate in its relation to the acid-base equilibria of the hemoglobin, phosphate, etc.

References on acid-base equilibria of blood and related mechanisms. See also "Urine."

1898—Bugarszky-Tangl, Spiro-Pemsel.

1900—Höber.

1901—Rhorer.

1902—Friedenthal, Höber.

1903—Auerbach-Friedenthal, Farkas, Farkas-Scipiades, Fraenckel, Friedenthal, Höber, Höber-Jankowsky.

1904—Friedenthal.

1905—Foà, Pfaundler.

1906—Abel-Fürth, Benedict, Szili.

1907—Aggazzotti.

1908—Henderson, Henderson-Spiro, Spiro-Henderson.

1909—Hendnrson, Michaelis-Rona, Ringer, Robertson, Szili.

1910—Höber, Kreibich, Robertson.

1911—Adler-Blake, Bottazzi, Hasselbalch-Lindhard, Löb, Polànyi, Schwartz-Lemberger, Skramlik, Winterstein.

1912—Hasselbalch, Hasselblach-Lundsgaard, Lundsgaard, Michaelis-Davidoff, Quagliariello-Agostino, Quagliariello, Rolly, Salge, Sellards.

1913—Elias-Kolb, Henderson-Palmer, Konikoff, Masel, Newburgh-Palmer-Henderson, Palmer-Henderson, Rona-György, Rona-Takahashi, Salge, Snapper.

1914—Barcroft, Blatherwick, Michaelis, Peabody, Peters, Rolly.

1915—Begun-Herrmann-Münzer, Hasselbalch-Gammeltoft, Henderson-Palmer, Levy-Rowntree-Marriott, Ma. de Corral, Menten-Crile, Milroy, Momose, Palmer-Henderson, Poulton, Wilson-Stearns-Thurlow, Winterstein.

1916—Gettler-Baker, Haldane, Hasselbalch-Lindhard, Howland-Marriott, Hurwitz-Lucas, Levy-Rowntree, Marriott, McClendon, McClendon-Magoon, Macleod, Reemlin-Isaacs, Rona-Ylppö, Scott, Ylppö.

1917—Bienstock-Czàki, Cullen, Fitz-Van Slyke, Hasselbalch, Henderson, Höber, Hooker-Wilson-Connet, Isaacs, McClendon-Shedlov-Thomson, Milroy, Palmer-Van Slyke, Parsons, Peters, Scott, Stillman-Van Slyke-Cullen-Fitz, Van Slyke, Van Slyke-Cullen, Van Slyke-Stillman-Cullen.

1918—Bayliss, Goto, · Hasselbalch-Warburg, Henderson-Haggard, Macleod, Macleod-Knapp, Sonne-Jarlöv, Straub-Meier, Zunz.

1919—Debenham-Poulton, Donegan-Parsons, Haggard-Henderson, Haskins, Irwin, Macleod, Parsons, Schloss-Harrington, Van Slyke-Stillman-Cullen, Van Slyke-Austin-Cullen.

1920—Anon, Bayliss, Bisgaard-Nørvig, Blatherwick, Campbell-Poulton, Collip, Collip-Backus, Coulter, Dale-Evans, Davies-Haldane-Kennaway, Dragstedt, Forbes-Halverson-Schulz, Fredericia, Grant, Goldman, Parsons, Haggard-Henderson, Hartridge, Haskins-

Osgood, Henderson, L., Henderson, Y., Henderson-Haggard-Coburn, Hills, Joffe-Poulton, v. Kapff, MacNider, Mellanby-Thomas, Menten, Michaelis, Moore, Parsons, Parsons-Parsons, Parsons-Parsons-Barcroft, Parsons-Shearer, Prentice-Lund-Harbo, Priestley, Raymund, Reimann, Rieger, Suitsu, Van Slyke-Palmer.

1921—Barr-Peters, Bazett-Haldane, Busa, Chistoni, Collip, Doisy-Eaton, Evans, C. L., Fleisch, Gauss, Haggard-Henderson, Haldane, Hastings-Murray-Murray, Henderson, Hill, Jarloev, Ma. de Corral, Means-Bock-Woodwell, Meier-Krönig, Parsons-Parsons, Peters-Barr, Peters-Barr-Rule, Reimann-Reimann, Reimann-Sauter, Roaf, Smith-Means-Woodwell, Trevan-Boock, Van Slyke, Van Slyke-Stadie, Winterstein.

1922—Barach-Means-Woodwell, Barkan-Broemser-Hahn, Cullen, Coulter, Doisy-Briggs-Chouke, Henderson, Hirsch-Peters, Hirsch-Williams, Macleod, Parsons-Parsons, Williams-Swett.

Bread. In the baking of bread it is essential that the proteins, such as glutin, which are responsible for the holding of the gas, shall be conditioned by the proper pH. The pH may also control the growth of the "rope" organism. The activity of yeast and the evolution of CO_2 from baking powders have relations to the pH of the dough.

References. Bailey-Peterson (1921), Cohn-Cathcart-Henderson (1918), Cohn-Henderson (1918), Cohn-Walbach-Henderson-Cathcart (1918), Freear-Venn (1920) Henderson (1918), Henderson-Cohn-Cathcart-Wachman-Fenn (1919), Henderson-Fenn-Cohn (1919), Jessen-Hansen (1911), Landenberger-Morse (1918) (1919), Lüers (1920), Patten (1920), Sharp-Gartner (1922), Wahl (1916).

Breeding. Control of spermatozoan activity. See "Comparative and General Physiology," and C. G. L. Wolf (1921).

Body Fluids (other than blood, urine, digestive juices, cerebrospinal fluid).

References. Aggazzotti (1921), Bloomfield-Huck (1920), Collip (1920), Farkas-Scipiades (1903), Foà, (1905, 1906), Fraenckel (1905), Gies (1916), Goldberger (1917), Hertel (1921), Huddelson (1921), Löb-Higuchi (1910), Loeb-Atchley-Palmer (1922), Long-Fenger (1915, 1916), Marshall (1915), Michaelis-Kramsztyk (1914), Okada (1915), Quagliariello (1916–1921), Schade-Neukirch-Halpert (1922), Shepard-Gies (1916), Uyeno (1919).

CANNING. The National Canners' Laboratory has so related time, temperature and pH that economy and certainty in the commercial sterilization of canned foods can be assured.

References. Bigelow (1921), Bigelow-Cathcart (1922), Kohman (1922), Rogers-Deysher-Evans (1921).

CARBONATE EQUILIBRIA. When carbon dioxid dissolves in water without any base to form carbonate there are presumably present in the water both anhydrous CO_2 and the hydrate, H_2CO_3, carbonic acid. Analytical methods do not ordinarily distinguish these two forms, and, since the sum of the two is generally the more important quantity, we may write the equilibrium equation for the relation between a partial pressure, P, of gaseous carbon dioxid and the dissolved carbon dioxid as follows:

$$[CO_2] + [H_2CO_3] = [\text{free } CO_2] = K_oP$$

In the presence of bases we still have the above relation holding tbetween the partial pressure and that portion of the total CO_2 cwhich remains uncombined. However, variation in the composition of the solution will vary the magnitude of K_o. We probably make no significant error if we regard [free CO_2] in carbonate solutions to be influenced by the total salt (carbonate) just as it is influenced by the total salt concentration in a solution containing no base. On this basis Johnston (1915) uses Bohr's data for the absorption coefficients of carbon dioxid in sodium chlorid solutions of different concentration, and calculates therefrom the values of K_o in terms of molar concentration.

Johnston's table of K_o.

TEMPERATURE	IN WATER	IN 1.17 M SALT	IN 3.44 M SALT
3.5	0.0672	0.0484	0.0270
4.2	0.0500	0.0367	0.0213
16.0	0.0441	0.0328	0.0193
25.0	0.0338	0.0260	0.0159
30.0	0.0297	0.0232	0.0142
40.0	0.0236	0.0185	0.0117

From these values Johnston interpolates the following values of K_o for the indicated concentrations of total base or salt at 25°C. Included below are the values of $pK_o = \log \frac{1}{K_o}$,

TOTAL BASE OR SALT	0.0	0.1	0.2	0.3	0.5	1.0
K_o.....	0.0338	0.0329	0.0321	0.0314	0.0300	0.0270
pK_o	1.471	1.483	1.493	1.503	1.523	1.569

Dissolved CO_2 reacts with water and since $[H_2O]$ may be regarded as constant we have the equilibrium equation

$$\frac{[CO_2]}{[H_2CO_3]} = K' \text{ or } \frac{[CO_2] + [H_2CO_3]}{[H_2CO_3]} = K' + 1 \qquad (68)$$

The H_2CO_3 dissociates in steps and for the first step the equilibrium condition is

$$\frac{[H^+]\,[H\bar{C}O_3]}{[H_2CO_3]} = K'' \qquad 69)$$

Combining equations (68) and (69) and collecting constants we have

$$\frac{[H^+]\,[H\bar{C}O_3]}{[CO_2] + [H_2CO_3]} = K_1$$

or using the convention mentioned above

$$\frac{[H^+]\,[H\bar{C}O_3]}{[\text{free } CO_2]} = K_1 \qquad (70)$$

The constant K_1 is sometimes called the first dissociation constant of carbonic acid. It is not strictly so but is rather of the nature of an "apparent dissociation constant." K_1 is more useful than the true dissociation constant but is probably much smaller.

For the second stage of dissociation the equilibrium condition is:

$$\frac{[H^+]\,[\bar{C}\bar{O}_3]}{[H\bar{C}O_3]} = K_2 \qquad (71)$$

In addition to these equations there is the useful relation of electrical neutrality,

$$[B^+] + [H^+] = [H\bar{C}O_3] + 2\,[\bar{C}\bar{O}_3] + [\bar{O}H] \qquad (72)$$

where $[B^+]$ represents the total concentration of cations other than $H^+]$ and all species are represented in equivalent concentrations.

One of the chief experimental difficulties in handling carbonate solutions is the control or the evaluation of P. But while this is susceptible to management the correct evaluation of K_1 and K_2 is a matter of great complexity for the following reasons. If salts such as Na_2CO_3 and $NaHCO_3$ are used as experimental material to establish various proportions of carbonate and bicarbonate ions it becomes necessary to know the degree of their dissociation at known concentrations of the salts, or if complete dissociation occurs it becomes necessary to know the effect of different concentrations upon activities. This involves the whole unsettled question of the conduct of "strong electrolytes." Hitherto there have been carried over to pH studies the constants derived by the use of conductivity data which are not strictly applicable.

If y_1 represents the degree of dissociation of $NaHCO_3$ and y_2 degree of dissociation of Na_2CO_3 we have the following relations according to Seyler and Lloyd (1917).

[Na]	0.05	0.1	0.2	0.3	0.5	1.0
y_1	0.82	0.78	0.73	0.69	0.64	0.52
y_2	0.56	0.66	0.37	0.31	0.24	0.14

Space does not permit a detailed discussion of the above values and numerous other quantities which enter into the data of carbonate equilibria. We shall proceed with the more general relations indicated by the pure equilibrium equations and shall give without comment Johnston's values for the more important constants.

Putting the equations into logarithmic form, and using for terms such as $\log \frac{1}{K}$ the expression pK, we have the following useful relations:

$$pH = pK_1 + \log [H\overline{C}O_3] - \log [\text{free } CO_2] \tag{73}$$

$$pH = pK_1 + pK_o + \log [H\overline{C}O_3] - \log P \tag{74}$$

$$pH = pK_2 + \log [\overline{C}\overline{O}_3] - \log [H\overline{C}O_3] \tag{75}$$

$$pH = \tfrac{1}{2} pK_o + \tfrac{1}{2} pK_1 + \tfrac{1}{2} pK_2 - \tfrac{1}{2} \log P + \tfrac{1}{2} \log [\overline{C}\overline{O}_3] \tag{76}$$

$$[B^+] = \frac{2K_oK_1K_2P + K_oK_1P[H^+] + K_w[H^+] - [H^+]^3}{[H^+]^2} \tag{77}$$

For the values of pK_o see page 320. From Johnston's selected

values for the first and second acid dissociation constants at 25°C. we have $pK_1 = 6.47$ and $pK_2 = 10.32$. For other values see references.

Inspection of the combined equations will show that pH is defined by any two of the variables or conversely that pH and one variable determine the state of a carbonate equilibrium. By the use of equation (77) the total base can be brought into consideration and it can be shown that the total base and one variable such as pH or P will define the position of a carbonate equilibrium. See (77). Thus a carbonate solution exposed to the atmosphere with its more or less constant partial pressure of CO_2 at 0.0003 atmosphere will tend to reach a definite pH value which is determined by the total base. This may be as low as pH 5.0 for solutions containing very little base or as high as pH 10 in a solution about normal with respect to $[B^+]$. Based upon such relations are analytical methods for determining CO_2 partial pressures from pH and known concentrations of total base.

Equations (73) and (74) are of importance in the study of blood the pH of which may be defined in terms of the ratio of bicarbonate to free CO_2 or in terms of bicarbonate and P. See section on blood. Direct experimental data for which equation (75) expresses the fundamental relations are given as follows by Auerbach and Pick (1912):

pH values for mixtures of sodium carbonate and bicarbonate at 18°C. after Auerbach and Pick

MOLS PER LITRE		pH	MOLS PER LITRE		pH
$NaHCO_3$	Na_2CO_3		$NaHCO_3$	Na_2CO_3	
0.20	0.00	8.35	0.10	0.000	8.35
0.19	0.01	8.90	0.09	0.005	8.98
0.18	0.02	9.15	0.08	0.010	9.30
0.16	0.04	9.45	0.07	0.015	9.50
0.14	0.06	9.65	0.06	0.020	9.60
0.12	0.08	9.96	0.05	0.025	9.87
0.10	0.10	10.10	0.04	0.030	10.05
0.08	0.12	10.35	0.03	0.035	10.23
0.06	0.14	10.45	0.02	0.040	10.35
0.04	0.16	10.65	0.01	0.045	10.7
0.02	0.18	11.0–11.8	0.00	0.050	11.4
0.00	0.20	11.59			

Equation (76) is of importance when it is desired to know the relations between partial pressure of CO_2 and the state of some carbonate equilibrium such as that of calcium carbonate. In this case we have another set of relations. Calcium carbonate is but slightly soluble *per se*. In the equilibrium equation

$$\frac{[Ca^{++}]\,[\overline{CO}_3]}{[CaCO_3]} = K$$

we often have to deal with a constant value of $CaCO_3$ maintained by the presence of solid $CaCO_3$. Under such circumstances we may combine this constant with the dissociation constant giving

$$[Ca^{++}]\,[\overline{CO}_3] = K_s \qquad (78)$$

where K_s is the "solubility product."

By combining (78) with (76) it is seen how Ca^{++} can be governed by P, a relation of geological importance.

K_s varies with the nature of the solid phase, (Calcite, Aragonite or precipitated calcium carbonate of different states of fineness). It is of the order of 1×10^{-8}.

The equations of carbonate equilibria have been left in their more general form to show the more general relations. Modifications for special purposes are very numerous and beyond the scope of this sketch. For detailed treatment see references under "Analyses," "Blood," "Water," "Equilibria," etc. A treatment of the general biological importance of the carbonate equilibria is given in *The Fitness of the Environment* by Henderson.

References. Auerbach-Pick (1912), Bjerrum-Gjaldbaek (1919), Frary-Nietz (1915), Henderson (1913), Henderson-Black (1908), Johnston (1915, 1916), Johnston-Williamson (1916), McClendon (1917), McClendon-Shedlov-Thomson (1917), Michaelis-Rona (1914), Prideaux (1915), Seyler-Lloyd (1917), Thiel-Stroheker (1914), Tillmans (1921), Van Slyke (1917, 1922), Wagner-Enslow (1922), Walker-Cormack (1900), Wilke (1921), Windish-Dietrich (1920).

Catalysis. The catalytic activity of the hydrogen and the hydroxyl ions in such transformations as the hydrolysis of cane sugar has taken a prominent place in the development of the theory of electrolytic dissociation. Under limited conditions one or an-

other of these catalytic processes is proportional to the concentration of the hydrogen or the hydroxyl ions; but there may enter the action of neutral salts. The theory of their influence is now being recast in accord with the concept of "activity." The older literature on hydrogen and hydroxyl ion catalyses is reviewed in the monograph by Woker (1910, 1915). A few recent references are: Abel (1920), Åkerlöf (1921), Jones-Lewis (1920), Kailan (1920), Karlson (1921), Northrop (1921). See Enzymes, Salt Action and Chapter XX.

Cerebrospinal Fluid.

References. Bisgaard (1913), Botazzi-Craifaleanu (1916), Collip (1920), Felton-Hussey-Bayne-Jones (1917), Hertel (1921), Hurwitz-Tranter (1916), Levinson (1917, 1919), Meier (1921), Shearer-Parsons (1921), Weston (1916).

Cheese.

References. Allemann (1912), Barthel-Sandberg (1919), Okuda-Zoller (1921), van Dam (1910).

Colloids. That the dispersion of colloids may be influenced by the "reaction" of the medium has long been known. So widely scattered is the literature on this particular phase of colloid chemistry that the author has made no attempt to assemble it. It is through the study of protein solutions that the most distinctive advances have been made. Beginning with Hardy the study of proteins as amphoteric electrolytes has been carried forward by Pauli, Michaelis, Robertson, Sørensen, Henderson, Loeb and others until there has developed a distinct protest against the separation of *certain* of the phenomena of colloids from the application of the simpler relations of crystalloids. How far the matter may be pushed in its application to other types of material taking the "colloidal state" remains to be determined.

A very good discussion of the relation of the developments in protein chemistry to colloid chemistry is given by Sørensen (1917). Compare Loeb, 1922.)

References. Abderhalden-Fodor (1920), Adolf-Pauli (1921), Arrhenius (1922), Bethe (1920), Clowes (1913), Ellis (1911) Labes (1921), Lachs-Michaelis (1911), Lillie (1909), McBain-Salmon (1920), McDougal-Spoehr (1919), McGuire-Falk (1922), Meier-Krönig (1921), Michaelis (1920, 1921, 1922), Michaelis-Rona (1919-1920), Ostwald (1912), Perrin (1904), Procter (1921),

Rona-Michaelis (1919), Schoucroum (1920), Smith (1920), Spiro (1916), Stiegler (1921), Varga (1919), Walpole (1914), Williams (1920). See also "Proteins," "Adsorption," "Donnan Equilibrium," "Electrophoresis."

Comparative and General Physiology.

References. Aggazzotti (1913), Andrus (1919), Arrhenius (1921), Atkins (1922), Barkan-Broemser-Hahn (1922), Barratt (1905), Bernstein (1913), Bethe (1909), Brenner (1921), Broderick (1921), Burgh-Clark (1921), Burridge (1920, 21), Carr (1921), Clowes-Smith (1922), Cohn (1917), Collett (1919, 1921), Collip (1920–1921), Coulter (1920), Cremer (1906), Crozier (1915–19), Dale (1913), Dale-Thacker (1914), Fletcher-Hopkins (1907), Galeotti (1906, 1920), Garrey (1920), Girard (1909), Goldberger (1917), Gray (1920), Hampshire (1921), E. N. Harvey (1920), R. B. Harvey (1920), Hastings-Murray (1921), Herbst (1904), Hirsch (1921), Hiruma (1917), Höber (1910), Hopkins (1921), Hurwitz (1910), Ivy-Oyama (1921), Jacobs (1920–22), Jameson-Atkins (1921), Jewell (1920), Kahlenberg (1900), Kastle (1898), Kopaczewski (1914), Křižencký (1916), Langefeldt (1921), J. Loeb, (1898, 1903, 1904, 1906), Loeb-Wasteneys (1911), R. Loeb (1920), Lloyd (1916), MacArthur (1920), McClendon (1916, 1920), McClendon-Mitchell (1912), MacDougall (1921), Meyerhof (1918), Mines (1912), Moore (1919, 1920), Moore-Roaf-Whitley (1905), Moore-Whitley-Webster (1921), Morse-Goldberg (1922), Neilson-Meyer (1921), Neugarten (1919), Odén (1916), Ostwald-Kuhn (1921), Parnas-Wagner (1914), Pechstein (1915), Philippson-Hannevart (1920), Plotho (1920), Popielski (1919), Porcelli-Titone (1914), Powers (1921–22), Prentice-Lund-Harbo (1920), Reichel (1922), Resch (1917), Richards (1898), Ritchie (1922), Roaf (1912–1922), Rohde (1920), Rona-Wilenko (1914), Roncati-Quagliariello (1921), Roth (1917), Saunders (1920), Schwyzer (1914), Shelford-Powers (1915), Shohl (1914), Straub-Meier (1919), Traube (1920), Warburg (1910), Wells (1915), Whitley (1905), Wolf (1921).

Crystallography. Wherry (private communication) states that there is reason to believe that the pH of a medium may sometimes control crystal form.

Culture of organisms other than bacteria, plants and tissue.

References. Bodine, (1921), Young-VanSant (1922). See also

numerous notes in references under "Comparative and General Physiology," "Bacteriology," and "Tissue culture."

Dakin's Solution.

Reference. Cullen-Austin (1918).

Digestive System. The digestive tract is primarily the channel for the intense activity of hydrolytic enzymes and as such is provided with mechanisms for the establishment of hydrogen ion concentrations favorable to these enzymes. Hydrogen electrode methods have correlated the regional activity of particular enzymes with the reactions there found, have clarified some of the differences between the digestive processes of infancy and adult life, aided in the explanation of the acid and alkali formation, and have been of service in the improvement of clinical methods for the assay of pepsin activity and the diagnosis of abnormal secretion of hydrochloric acid in the stomach. The control of specific physiological functions such as secretion of conditioning agents (see Bayliss, 1918), permeabilities, and activities of the varied musculature, as well as investigations upon the condition in the digestive tract of substances such as calcium and phosphate which form insoluble precipitates are subjects which present promising material for the application of modern methods. Shohl and King (1920) have recently reviewed and improved methods of studying gastric acidity.

References. Allaria (1908), Ambard-Foà (1905), Auerbach-Pick (1912, 1913), Cannon (1907), Christiansen (1911, 1912, 1921), Davidsohn (1911, 1912, 1913, 1921), Foà (1905, 1906), Fowler-Bergeim-Hawk (1915), Fraenckee (1905), Graham (1911), Hahn (1914), Hainiss (1921), Hammett (1922), Hess (1915), Hess-Scheer (1921), Howe-Hawk (1912), Huenekens (1914), Krummacher (1914), Lanz (1921), Long-Fenger (1917), McClendon (1915, 1920), McClendon-Bissell-Lowe-Meyer (1920), McClendon Myers-Culligan-Gydesen (1919), McClendon-Shedlov-Thomson (1917), McClendon-Shedlov-Karpman (1918), McWhorter (1918), Menten (1915), Michaelis (1917, 1918, 1920), Michaelis-Davidsohn (1910), Myers-McClendon (1920), Nelson-Williams (1916), Okada-Arai (1922), Popielski (1919), Rolph (1915), Rona-Neukirch (1912), Salge (1912), Scheer (1921), Schryver-Singer (1913), Shohl (1920), Shohl-King (1920), Tangl (1906), Traube (1920), Ylppö (1916).

Dissociation Constants as determined with the hydrogen electrode or indicator methods. Compare Chapter I.

References. Agostino-Quagliariello (1912), Dernby (1916), Eckweiler-Noyes-Falk (1920), Eijdman (1906), Kastle (1905), Kolthoff (1918, 1920), Michaelis (1911, 1913, 1914), Michaelis-Garbendia (1914), Michaelis-Rona (1913, 1914, Prideaux (1911), Salm (1906, 1908), Scudder (1914), Tizard (1910), Weisse-Meyer Levy (1916). See Indicator constants.

Donnan Equilibrium. Imagine a solution of a simple electrolyte and a membrane permeable to the electrolyte. Upon one side of the membrane let there be a solution of a substance which cannot penetrate the membrane but which can enter into the equilibrium of the simple electrolyte. A simple case is the following. Let the initial state be illustrated by the following scheme where there is placed upon one side of the membrane M a dilute solution of HCl and upon the other side the acid HR neither the anion nor the undissociated residue of which can penetrate the membrane.

$$\begin{matrix}[HR]_1 \\ [H^+]_1 \\ [\bar{R}]_1\end{matrix} \Big| M \Big| \begin{matrix}[HCl]_2 \\ [H^+]_2 \\ [\bar{Cl}]_2\end{matrix}$$

Chlorine ions (or HCl) will diffuse from right to left until, when equilibrium is attained, there will be the following state

$$\begin{matrix}[HCl]_3 \\ [HR]_3 \\ [H^+]_3 \\ [\bar{R}]_3 \\ [\bar{Cl}]_3\end{matrix} \Big| M \Big| \begin{matrix}[HCl]_4 \\ [H^+]_4 \\ [Cl]_4\end{matrix}$$

If now we place hydrogen electrodes on the two sides of the membrane the E. M. F. of this gas chain will be determined in part by the relative concentrations of the hydrogen ions and in part by a potential difference across the membrane. This membrane potential difference we shall call E_d

$$E.M.F. = \frac{RT}{nF} \ln \frac{[H^+]_3}{[H^+]_4} + E_d$$

We may also place on the two sides electrodes, the potential differences at which are determined by the relative concentrations

of the chlorine ions (e. g. Pt–Cl electrodes or calomel electrodes). For such a chain we would have

$$\text{E.M.F.} = \frac{RT}{nF} \ln \frac{[\bar{Cl}]_4}{[\bar{Cl}]_3} + E_d$$

We have already specified however that the system is at equilibrium. Therefore no energy could be obtained from either one of the chains described above. The E. M. F. in each case is then zero and since E_d is the same in each case

$$\frac{[H^+]_3}{[H^+]_4} = \frac{[\bar{Cl}]_4}{[\bar{Cl}]_3}$$

The rule of electrical neutrality indicates that on the right side of the membrane $[H^+]_4 = [Cl^-]_4$. Combining this relation with the other we then have

$$[H^+]_4^2 = [H^+]_3 [\bar{Cl}]_3$$

There are various directions in which we may now proceed. As one example let us assume the very simple case where the dissociations of HR and HCl are complete, and let us further assume that the system is divided by the membrane into two equal parts. Between the initial and the final state of the system chlorine ions have diffused from right to left until the concentration $[\bar{Cl}]_3$ is x. Then $[H^+]_3 = [H^+]_1 + x$ and $[H^+]_4 = [H^+]_2 - x$. Introducing these values into the foregoing equation we have

$$([H^+]_2 - x)^2 = x([H^+]_1 + x)$$

or

$$\frac{[H^+]_2 - x}{x} = \frac{[H^+]_2 + [H^+]_1}{[H^+]_2} \quad \text{or} \quad x = \frac{[H^+]_2^2}{[H^+]_1 + 2[H^+]_2}$$

The following table will give an idea of the magnitude of the effects due to the conditions assumed.

As we have already indicated, the difference of potential between two hydrogen electrodes placed on opposite sides of the membrane must, at the equilibrium state of the system, be equal and opposite to the potential difference at the membrane. Hence

the membrane potential difference may be expressed *in terms of* a hydrogen electrode gas chain:

$$-\frac{RT}{F}\ln\frac{[H^+]_3}{[H^+]_4}.$$

By using this relation we calculate the membrane potential difference given in millivolts in the last column of the following table.

$[R^-]_1=[H^+]_1$	$[H^+]_2$	INITIAL RATIO $\frac{[H^+]_1}{[H^+]_2}$	PER CENT HCl DIFFUSED TO ESTABLISH EQUILIBRIUM	EQUILIBRIUM DISTRIBUTION RATIO $\frac{[H^+]_3}{[H^+]_4}$	MEMBRANE POTENTIAL IN MILLIVOLTS
0.01	1.0	0.01	49.8	1.01	− 0.3
1.0	1.0	1.0	33.3	2.0	− 18.0
1.0	0.01	100.0	0.98	101.0	−120.0

Of course the conditions assumed for purposes of illustration are extremely simple but they suffice to indicate the nature of relations of very great importance in the physiology of the living cell.

References. Donnan (1911), Donnan-Harris (1911), Loeb (1921–22), Michaelis (1922), Moore-Roaf-Webster (1912), Sørensen (1917). See also "Blood," "Comparative and General Physiology."

Dry-Cells.

Reference. Haller-Ritchie (1920).

Electroplating. The potential at which hydrogen is deposited freely upon an electrode is a function of the hydrogen ion concentration of the solution. Therefore pH is important in controlling gassy deposits. In addition it is found that buffer solutions, maintaining the pH within definite limits, aid in the production of desirable qualities in nickel deposits.

References. Bennett-Rose-Tinkler (1915), Blum (1920, 1921), Küster (1900), Thompson (1922).

Electrophoresis (Cataphoresis) and Electro-Osmosis. An electrically charged body placed between an anode and a cathode will tend to move toward the pole having a charge opposite in sign to the charge on the body. If the body is a simple ion, the movement is called ionic migration. If the body is a particle

suspended in a medium such as water, the movement is called electrophoresis. More generally it is known as cataphoresis. The distinction between ionic migration and electrophoresis is not always clear in the case of material in the colloidal state.

We shall not discuss the various theories advanced to account for the experimental facts but shall treat briefly only that point of view which it will be profitable to investigate further with the aid of methods for determining pH.

Since acidic or basic ionization may determine the sign of the charge upon a body of amphoteric nature the sign may be a function of the pH of the medium (aqueous). The direction of electrophoresis is then a function of pH. At the isoelectric point electrophoresis is a minimum. The position of this minimum on the pH scale is a function of the acidic and basic dissociation constants and the zone of the minimum may be narrow or broad according to the relative magnitudes of the constants. See Chapter 1. The method of electrophoresis is useful in determining isoelectric points.

There can be no movement such as that noted above without a reciprocal interaction between suspended or dissolved material and the dispersing medium. If then the charged particles are fixed in position, as in the form of a porous diaphragm, are placed in water and the whole subjected to a potential gradient, the water will tend to move (electro-osmosis). The same *relative* relations indicated above then hold. If the diaphragm is of an amphoteric nature the direction of water flow will depend upon the acidic and basic properties of the diaphragm and upon the pH of the aqueous phase.

In either one of the two cases (particles fixed or free to move) the same end result will be obtained if the particles adsorb hydrogen and hydroxyl ions according to such adsorption isotherms that equality of adsorption and consequently equality of electrical charge is attained at a definite pH value. On either side of this pH value the excess adsorption of one or the other ion will depend upon their concentrations which are a function of pH by reason of the relation $[H^+]\,[OH^-] = K_w$. The position of this "isoelectric" point is a function of the properties of the material and may lie anywhere along the pH scale (according to the nature of the material) with a narrow or broad isoelectric zone.

The converse to the above propositions is that filtration produces a potential difference across the filter which is a function of the acidic and basic nature of the filter and of the pH of the solution filtered.

Obviously the above sketch covers restricted conditions.

References. Barratt-Harris (1912), Briggs (1918), Freundlich (1921) Gyemant (1921), Michaelis (1914, 1922), Perrin (1904–1905), Porter (1921), Putter (1921), Steigmann (1920), Szent-Györgyi (1920, 1921), Svedberg (1916), Svedberg-Anderson (1919). See also "Isoelectric Point."

Enzymes. The activity of enzymes as influenced by the hydrogen ion concentration of the solution has occupied the attention of many investigators since the publication of Sørensen's paper (1909). The analogy between the activity curves of several enzymes and the curves relating the "dissociation residues" of amphoteric electrolytes to pH suggested to Michaelis the amphoteric nature of enzymes (cf. Loeb 1909). Northrop has shown important relations of activity to the acid-base nature of the substrate. Holderer's observations on the extraction of enzymes from cells with solvents of different reaction are most suggestive. The necessity of controlling the pH of enzyme solutions for assays as well as in the study of the effect of salts and in experiments having to do with the formulation of the laws of enzyme activity (Van Slyke and Cullen) is now generally recognized. Barendrecht in the development of his radiation theory notes the special importance of the hydrogen ions.

The following is a rough classification of studies on specific enzymes.

Amygdalase. Bertrand-Compton (1921).

Amylase. Ambard (1921), Biederman-Rueha (1921), Euler-Svanberg (1921), Falk-McGuire-Blount (1919), Maestrini (1921), Groll (1920), McGuire-Falk (1920), Sherman (1919), Sherman-Thomas-Baldwin (1919), Sherman-Schlessinger (1915), Sherman-Thomas (1915), Sherman-Walker (1917), Sjöberg (1920), Takamine-Oshima (1920).

Bacterial enzymes. Abderhalden-Fodor (1921), Avery-Cullen (1920), Barthel-Sandberg (1920), Blanc-Pozerski (1920), Clark (1920), Dernby (1917), Dernby-Blanc (1921), Gröer (1912), Itano (1916), Kanitz (1903), Lord (1919), Meyer (1911), Nye (1922), Waksman (1918), West-Stevens (1921).

Carboxylase. Neuberg (1915).

Catalase. Bodansky (1919), Burge (1920), Euler-Blix (1919), Falk-McGuire-Blount (1919), Harvey (1920), Michaelis-Pechstein (1913, 1914), Morgulis (1921), Phragmén (1919), Senter (1905), Sørensen (1909), Sjöberg (1920), Waentig-Steche (1911).

Cellase. Bertrand-Holderer (1910).

"*Diastases*" (Important historical references) Fernbach (1906), Fernbach-Hubert (1900).

Filtration of. Holderer.

Glycogenase. Norris (1913).

Coferments. Biederman (1921), Tholin (1921).

Emulsin. Bayliss (1912), Nordefeldt (1921), Vulquin (1910). Willstätter-Csányi (1921).

Erepsin. Euler (1907), Dernby (1916), Rona-Arnheim (1913).

Esterases (*lipase*). Avery-Cullen (1920), Baur (1909), Davidsohn (1912–1913), Falk, I. (1918), Falk, K. (1916), Groll (1920), Haley-Lyman (1921), Hulton-Frankel (1917), Kastle (1902), Rona (1911), Rona-Bien (1914), Rona-Reinicke (1921), Rona-Michaelis (1911).

Invertase. Bertrand-Rosenblatt-Rosenblatt (1912), Euler (1921), Euler-Laurin (1919, 1920), Euler-Svanberg (1918–21), Fales-Nelson (1915), Falk-McGuire (1921), Fodor (1921), Griffin-Nelson (1916), Hudson (1910), Hudson-Paine (1910), Kanitz (1911), Langefeldt (1921), Michaelis (1921), Michaelis-Davidsohn (1911), Michaelis-Menten (1913), Michaelis-Pechstein (1914), Michaelis-Rothstein (1920), Nelson-Griffin (1916), Nelson-Hitchcock (1921), Nelson-Vosburgh (1917), Rona-Bach (1921), Rona-Bloch (1921), Sjöberg (1921), Sørensen (1909), Vosburgh (1921).

Lactase. Davidsohn (1913).

Levanase. Kopeloff-Kopeloff-Welcome (1920).

Maltase. Adler (1916), Kopaczewski (1912, 1914, 1915), Michaelis-Rona (1913, 1914), Rona-Michaelis (1913).

Oxidases, etc. Bunzel (1915), Bunzell (1916, 1917), Ohlsson (1921), Menten (1919, 1920), Reed (1916), Rose-Kraybill-Rose (1920).

Oxynitrilase. Krieble-Wieland (1921).

Pectase. Euler-Svanberg (1919).

Optimum temperature. Compton (1915, 1921). Euler-Laurin (1920).

Papain. Frankel (1917). Chesnut (1920).

Peroxidase. Bouma-Van Dam (1918).

Pepsin. Christiansen (1912), Van Dam (1915), Davidsohn (1912), Funk-Niemann (1910), Gies (1902), Graber (1921), Groll (1920), Gyemant (1920), Loeb (1909), Michaelis (1918), Michaelis-Mendelsohn (1914), Michaelis-Rothstein (1920), Northrop (1919, 1920, 1921), Okada (1916), Pekelharing-Ringer (1911), Ringer (1918), Rohonyi (1912), Sørensen (1909).

Phosphatase. Adler (1915).

Rennet. Allemann (1912), Van Dam (1908, 1909, 1912, 1915), Funk-Niemann (1910), Madsen-Walbum (1906), Michaelis-Mendelsohn (1913), Michaelis-Rothstein (1920), Milroy (1915), Thaysen (1915).

Salivary diastase (ptyalin). Cole (1903), Hahn-Harpuder (1920) Michaelis-Pechstein (1914), Ringer-Trigt (1912). See amylase.

Taka-diastase. Okada (1916).

Trypsin. Auerbach-Pick (1913), Hahn-Mickalik (1921), Kanitz (1902), Michaelis-Davidsohn (1911), Northrop (1921, 1922), Palitzsch-Walbum (1912), Ringer (1921), Robertson-Schmidt (1908).

Theory of action. Barendrecht (1920), Euler (1920), Falk (1921), Loeb (1909), Michaelis (1909, 1914), Michaelis-Davidsohn Mad 1911), Rohonyi (1911), Van Slyke-Cullen (1914).

Urease. Barendrecht (1920), Lövgren (1921), Onodera (1915), Rona-György (1920), Rona-Petrov (1920), Van Slyke-Cullen (1914), Van Slyke-Zacharias (1914).

EQUILIBRIA. The hydrogen electrode and indicators in the determination of affinity constants, free energy, hydrolysis, etc.

References. Adolf-Pauli (1921), Bjerrum (1907–21), Böeseken-Kerstjens (1916), Bogue (1920), Chow (1920), Denham (1908), Eucken (1907), Ellis (1916), Ferguson (1916), Frary-Nietz (1915), Fricke (1920), Hardman-Lapworth (1911), Harned (1915–1922), Heyrovsky (1920), Jahn (1900, 1901), Kanitz (1921), Lewis (1908, 1912, 1913), Lewis-Brighton-Sebastian (1917), Lewis-Randall (1914), Linhart (1919), Löffler-Spiro (1919), Loomis-Acree (1911), Loomis-Essex-Meacham (1917), Löwenherz (1896), Margaillan (1913), McBain-Coleman (1914), MacInnes (1919), Merrill (1921), Nernst (1889), Newbery

(1914), Noyes-Ellis (1917), Noyes-Freed (1920), Richards-Dunham (1922), Rosenheim-Leyser (1921), Tizard (1910), Tizard-Boeree (1920), Tolman-Greathouse (1912). See also Chapters IV, VI, XVI.

Explosives.

References. Farmer (1920), Angeli-Errani (1920).

Feces. See "Digestive System."

Filtration. Hydrogen ion concentration, through its influence upon the dispersion of certain colloids and upon the conditioning of filter material, may control the filterability of a substance. Holderer's thesis from Perrin's laboratory presents in admirable form many of the theoretical aspects of the subject. A republication of this rare thesis is desired. The subject is not only of considerable theoretical interest but also of great practical importance. Buffer control with indicator tests may in many instances facilitate filtrations upon an industrial as well as a laboratory scale.

References. Aubel-Colin (1915), Holderer (1909, 1910, 1911, 1912), Homer (1917), Loeb (1919), Schmidt (1914), Strada (1908), Wilson (1921), Wilson-Copeland-Heisig (1921), Wilson-Heisig (1921). See also "Electrophoresis."

Foods, pH of. The National Canners' Laboratory has made a number of determinations of the pH of canned foods. See "Canning." See also "Milk," "Cheese," "Wine," "Beer," "Vinegar," references given by Clark and Lubs (1917)[1] and the paper by McClendon and Sharp (1919). The influence of the pH upon the stability of a "vitamine" has been studied by La Mer (1921), and Campbell, LaMer and Sherman (1922). cf. Harden and Zilva (1918). For sterilization of canned goods see "Disinfection" under "Bacteriology" and "Canning".

Glass, effect of, on reaction of solutions.

References. Esty-Cathcart (1921), Ewe (1920), Fabian-Stull (1921), Levy-Cullen (1920), Russell-Nichols-Stimmel (1920).

Glucose, decomposition of, as influenced by pH.

References. Elias-Kolb (1913), Euler-Hedelius (1920), Henderson (1911), Mathews-McGuigan (1907), Michaelis-Rona

[1] Some of the pH values given by Clark and Lubs for acidified or alkalinized extracts have been misquoted as the pH values of the original material.

(1909–1912), Nef (1913), Rona-Arnheim (1913), Rona-Doblin (1911), Rona-Wilenko (1914). Also references in Woker.

Hemolysis.

References. Atkin (1911, 1914), Cook-Mix-Culvyhouse (1921), Coulter (1921), Fenn, (1922), Fühner-Neubaur (1907), Gros (1910), Haffner (1920), Hellens (1913), Jodlbauer-Haffner (1920, 1921), Jordan (1903), Kozawa (1914), Krogh (1909), Lagrange (1914), Michaelis-Skwirsky (1909), Michaelis-Takahashi (1910), Stevens-Koser (1920), Teague-Buxton (1907), Walbum (1914, 1915).

Hydrolysis. The reaction between an acid and a base is reversible.

$$HA + BOH \rightleftharpoons BA + H_2O$$

There are present then both free acid and free base even when the two are mixed in equivalent proportions. This last condition can be duplicated by making up the solution in the first place with the pure salt. The above reaction then goes from right to left until the equilibrium state is reached and the process is called hydrolysis, because it may be regarded as a splitting of water molecules.

Now the resulting acid and base ionize, the one tending to increase the hydrogen ion concentration, the other tending to increase the hydroxyl ion concentration. If the acid is more highly dissociated than the base the solution will contain more hydrogen ions than hydroxyl ions; and if the base is more highly dissociated than the acid the solution will contain more hydroxyl than hydrogen ions. Since the magnitude of a dissociation constant is a measure of dissociation tendency the reaction of a salt solution will depend upon the relative magnitudes of the K_a and K_b constants of the component acid and base.

In a solution of the salt, BA, we have present BA, B^+, A^-, HA, BOH, H^+ and $OH.^-$

By the rule of electrical neutrality $[A^-] + [OH^-] = [B^+] + [H^+]$. Since total acid = total base, $[HA] + [A^-] + [BA] = [B^+] + [BOH] + [BA]$. Introducing the acid and the base equilibrium equations and the relation $[H^+]\,[OH^-] = K_w$ and combining these equations we have

$$[H^+] = \sqrt{K_w \frac{K_a}{K_b} \frac{(K_b + [B^+])}{(K_a + [A^-])}}$$

If now K_b and K_a are small in relation to $[B^+]$ and $[A^-]$, and if the solution is sufficiently dilute so that $[B^+]$ and $[A^-]$ each approximate the salt concentration [S], then approximately

$$[H^+] = \sqrt{K_w \frac{K_a}{K_b}}$$

Cf. formula for isoelectric point of ampholyte. When $K_a = K_b$, $[H^+] = 10^{-7}$, pH = 7.

If we are dealing with a salt, the acid component of which is very "strong" we may regard the acid set free by the hydrolysis of the salt as completely dissociated. [HA] in the above equations is placed equal to zero and we then derive

$$[H^+] = \sqrt{\frac{K_w}{K_b}(K_b + [B^+])}$$

If now K_b is small in relation to $[B^+]$ and if $[B^+]$ approximates [S]

$$[H^+] \text{ approximates } \sqrt{\frac{K_w}{K_b}[S]}$$

Conversely when the base is very strong and when the same assumptions made above are maintained

$$[H^+] \text{ approximates } \sqrt{\frac{K_a K_w}{[S]}}$$

References. See treatment by Bjerrum (1914), example by Denham (1908), and numerous references under "Equilibria."

Indicator Constants. See Prideaux and Chapters IV and VIII.

References. Clark-Lubs (1917), Gillespie (1920), Paulus-Hutchinson-Jones (1915), Kolthoff (1918–1922), Michaelis (1920), Michaelis-Gyemant (1920), Michaelis-Krüger (1921), Rosenstein (1912), Schaeffer-Paulus-Jones (1915), Salm (1904), Tizard (1910).

Indicators, natural.

References. Bribaker (1914), Crozier (1916, 1918), Haas (1916), Pozzi-Escot (1913), Sacher (1910), Scheitz (1910), Stephanides (1916), Trillat (1916), Walbum (1913), Watson (1913). See also Perkin and Everest.

INDUSTRIAL PROCESSES. See every subject in this chapter. Also the following special references.

References. Brewster-Raines (1921), Clark-Zoller-Dahlberg-Weimar (1920), Keeler (1922), Lubs (1920), Searle (1920), Wilson-Copeland-Heisig (1921), Wilson-Heisig (1921), Zoller (1921), and references on "Water Works" and "Leather."

ISOELECTRIC POINTS. See Chapter I.

References. Brossa (1915), Cohn (1920–1922), Cohn-Gross-Johnson (1920), Eckweiler-Noyes-Falk (1920), Fodor (1920), Loeb (1918–1922), Michaelis (1911–1920), Michaelis-Bien (1914), Michaelis-Davidsohn (1910–1913), Michaelis-Grineff (1912), Michaelis-Mostynski (1910), Michaelis-Pechstein (1912), Michaelis-Rona (1919), Michaelis-Takahashi (1910), Mills (1921), Rona-Michaelis (1910), Sørensen (1912, 1917), Stuber-Funck (1921), Szent Györgyi (1921), Thomas-Kelley (1922).

LEATHER AND TANNING.

References. Atkin (1922), Atkin-Atkin (1920), Atkin-Thompson (1920), Balderston (1913), Procter (1921), Procter-Wilson (1916), Povarnin (1915), Sand-Law (1911), Thomas-Baldwin (1919), Thomas-Foster (1921), Thomas-Kelly (1921, 1922), Wilson (1917, 1921), Wilson-Daub (1921), Wilson-Kern (1921), Wood-Sand-Law (1911). See also "Proteins."

MILK.

References. Allemann (1912), Aron (1914), Baker-Breed (1920), Baker-Van Slyke (1919), Chapman (1908), Clark (1915), Clark-Cohen (1922), Cooledge-Wyant (1920), van Dam (1908, 1918), Davidsohn (1912, 1913), Foà (1905, 1906), Hastings-Davenport (1920), Jones (1921), Kramer-Green (1921), Laqueur-Sackur (1903), Milroy (1915), Palmer-Dahle (1922) Rogers-Deysher-Evans (1921), Rona-Michaelis (1909), Schultz-Chandler (1921), Schultz-Marx-Beaver (1921), Sommer-Hart (1919, 1920), Stutterheinn, Szili (1917), Taylor (1913), Terry (1919), Tillmans-Obermeier (1920), Van Slyke-Baker (1918, 1919). See also "Cheese" and "Protein."

NEURO-PHYSIOLOGY.

References. Adrian (1920), Bottazzi-Craifaleanu (1916), Chiò (1907), Garry (1920), Grant (1920), Mansfield-Szent Györgyi (1920), Mayer (1916), Moore (1919), Neugarten (1919), Zotterman (1921). See also "Blood," (the respiration phase) and "Comparative and General Physiology."

PERMEABILITY of cells.

References. Bethe (1922), Clowes-Smith (1922), Collander (1920), Donnan (1911), Haas (1916), Harvey (1911, 1913), Haynes (1921), Holderer (1911), Jacobj (1920), Lillie (1909), Moore-Roaf-Webster (1912), Odén (1916), Reemelin-Isaacs (1916), Snapper (1913), Stiles-Jorgensen (1915), compare Filtration.

PHAGOCYTOSIS.

References. A. Evans (1921, 1922), Hamberger-Heckma (1908), Koltzoff (1914), Radsma (1920), Sawtchenko-Aristovsky (1912), Schwyzer (1914).

PHARMACOLOGY, etc. pH in relation to properties, activity, deterioration, and assay or detection of drugs.

References. Adams (1917), Crane (1921), Evers (1921), v. Gröer-Matula (1920), Hanzlik (1920, 1921), Kolthoff, (1920, 1922), Leech (1922), Levy-Cullen (1920), Macht-Shohl (1920), Meier-Krönig (1921), Mellon-Slagle-Acree (1922), Menten (1920), Moore (1920), Rippel (1920), Rona-Bach (1920), Shohl-Deming (1921), Snyder-Campbell (1920), Sollmann (1917), Tsakalotos-Horsch (1914), Williams-Swett (1922), Zoccola (1918).

PHYTO-PATHOLOGY AND PHYSIOLOGY.

References. Atkins (1922), Chambers (1921), Clevenger (1919), Crozier (1919), Harvey (1920), Hixon (1920), Lapicque (1921), MacDougal (1921), Martin (1921), Schmitz (1919), Schmitz-Zeller (1919), Webb (1919), Wherry (1918–22), Wolf-Foster (1921), Wolf-Shunk (1921), Zeller-Schmitz (1919).

See "Plant Distribution," "Comparative and General Physiology," "Soil."

PLANT DISTRIBUTION. Wherry, working with a simple field kit, has carried indicators into the field and has correlated the habitats of several plant species with the pH of their soils.

Investigations by O. Arrhenius in Sweden, by Olsen in Denmark and by Atkins in England and India have confirmed Wherry's observation that the pH of the soil is of great significance.

Such information has contributed toward methods of cultivating the blueberry and wild-flowers hitherto unknown or uncommon in garden and greenhouse.

References. Arrhenius (1920, 1922), Atkins (1921, 1922), Comber (1921), Emerson (1921), Fisher (1921), Gail (1919), Olsen (1921), Wherry (1920–1922). See also "Phytopathology and Physiology," "Soils," "Water," and especially "Bacteriology."

PROTEINS, by reason of their chemical structure, are amphoteric. As such they are subject to the pH of aqueous dispersing media as are the simple ampholytes. Though complete equilibrium equations are difficult to formulate we should expect the occurrence of pH points and zones comparable to the isoelectric points and zones of simple ampholytes. Experimentally these have been found. These are also points of optima, or minima, for various properties of protein solutions (e.g. minimal electrophoresis, viscosity and osmosis). If the solubility of the protein itself is less than that of its acid or basic salts, the protein can be precipitated at or near the isoelectric point (e.g. analysis and commercial preparation of casein). Closely related is the adjustment of pH favoring separation of crystals. Proteins are unable to penetrate many membranes but are able to enter into an acid-base equilibrium and thus exhibit many interesting relations in Donnan equilibria (Sørensen, Loeb).

The outstanding difficulty in treating proteins as electrolytes is the establishment of exact quantities for concentrations or activities which must necessarily be used in formulating equilibrium equations. The mathematical treatment by Michaelis and by Sørensen, and especially the painstaking experimental investigations to which Sørensen and his coworkers have devoted several years have advanced the subject beyond dependence on mere *analogy* to the conduct of simple ampholytes.

References. Adolf-Spiegel (1920), Agostino-Quagliariello (1912), Atkin (1920), Bogue (1921), Bovie (1920), Bugarszky-Liebermann (1898), Burrows-Cohn (1918), Chiari (1911), Chick (1913), Chick-Martin (1910–13), Clark-Zoller-Dahlberg-Weimar (1920), Cohn (1920–22), Cohn-Gross-Johnson (1920), Davis-Oakes-Browne (1921), Ferguson-France (1921), Field (1921), Fodor (1920–21), Haas (1918), Handovsky (1910), Hardy (1899, 1905), Henderson-Cohn-Cathcart-Wachman-Fenn (1919), Henderson-Palmer-Neuburgh (1914), Hill (1921), Hitchcock (1922), Laqueur-Sackur, (1903), Lloyd (1920, 1922), Loeb (1918–22), Manabe-Matula (1913), Michaelis (1909), Michaelis-Airila (1921), Michaelis-Mostynski (1910), Michaelis-Rona (1910, 1919), Michaelis-Szent Györgyi (1920), Mills (1921), Okuda-Zoller (1921), Oryng-Pauli (1915), Palmer-Atchley-Loeb (1921, 1922), Patten-Johnson (1919), Patten-Kelems (1920), Pauli (1903–1922), Pauli-

Handovsky (1908–10), Pauli-Matula (1919), Pauli-Samec (1909–14), Pauli-Wagner (1910), Pechstein (1913), Procter-Wilson (1916), Quagliariello (1912), Resch (1917), Robertson (1907–1918), Rohonyi (1912), Ryd (1917, 1918), Sharp-Gortner (1922), Schmidt (1916), Schorr (1911), Sollmann (1917), Sørensen (1917–1921), Sørensen & coworkers (1917), Sørensen-Jürgensen (1911), Spiro (1904, 1913), Starke (1900), Szent-Györgyi (1920, 1921), Thomas (1921), Wagner (1921), Wintgen-Krüger (1921), Wintgen-Vogel (1922), Ylppö (1913), Zoller (1921). See also "Isoelectric Point."

Salt-Action, theory and effects in relation to pH. See Chapters I, II, and VII.

References. Abegg-Bose (1899), Arrhenius (1888, 1889), Åkerlöf (1921), Brightman-Meachem-Acree (1920), Chick-Martin (1912, 1913), Falk (1918, 1920), Gillespie-Wise (1918), Harned (1915), Haynes (1921), Hofmeister (1891), Holm-Sherman (1921–1922), Kolthoff (1916–22), Lloyd (1916), Loeb (1906–1922), McBain-Coleman (1914), McClendon-Mitchell (1912), Michaelis (1914, 1920), Michaelis-Rona (1909), Michaelis-Szent Györgyi (1920), Michaelis-Timénez Dias (1921), Northrop (1920), Poma (1914), Poma-Patson (1914), Prideaux (1919), Rose-Kraybill-Rose (1920), Rosenstein (1912), Ryd (1917), Shearer (1920), Sherman-Thomas (1915), Sørensen-Palitzsch (1913), Sørensen-Sørensen-Linderstrøm Lang (1921), Spiro (1921), Szent-Györgyi (1920), Szyszkowski (1907), Thomas Baldwin (1919), Wells (1920), See especially references in Chapter II on "Activity."

Serology. See also Acid Agglutination of Bacteria, Hemolysis, Bacteriology, Proteins, Colloids.

References. Amako (1911), Atzler (1914), Brooks (1920), Buchanan (1919), Coulter (1921), Enlows (1922), Evans (1921, 22), Field-Teague (1907), Hirsch-Peters (1922), Homer (1917, 1918, 1919), Landensteiner (1913), Landensteiner-Prásek (1913), Langenstrass (1919), Lindenschatt (1913), Leschly (1916), Mason (1922), Michaelis-Davidsohn (1912), Neukirch (1920), Noguchi (1907), Ruppel (1920), Tulloch (1914, 1918).

Sewage.

References. Clark-Cohen (1922), Wilson-Copeland-Heisig (1921), Wilson-Heisig (1921).

Soap Solutions.

References. Beedle-Bolam (1921), McBain (1920), McBain-Bolam (1918), McBain-Martin (1918), McBain-Salmon (1920).

Soil Acidity has been confused by the complexities of titrimetric procedures, has been neglected, or has been considered to be an unreality by one or another school. Gillespie (1916) obtained good agreement between pH values of soil extracts determined by means of the hydrogen electrode and again by means of indicators. The practical significance of this is now revealed by studies which show characteristic pH values for well-defined types of soil, which show correlations between the pH of soil extracts and the growth of beneficial or harmful microörganisms, and which show correlations between the natural distribution of plants and the pH of the soils in which they are found.

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Solubility. The true solubility of a compound may be regarded as independent of the hydrogen ion concentration of a solution; but if the compound is an acid, base, ampholyte or salt some of the material present in solution is ionized and this portion is governed by the ionic equilibrium of which the hydrogen ion concentration is a part. Therefore the total solubility which is

generally of more importance than the true, partial solubility is a function of pH.

Consider the equilibrium $\frac{[H^+]\,[A^-]}{[HA]} = K_a$ and assume that the solubility of the acid HA itself is low so that we shall not encounter the difficulties inherent in the treatment of concentrated solutions. If the acid is present in the solid phase [HA] is maintained constant and is the partial solubility, S_p. On combining the constants in the above equation we have $[H^+]\,[A^-] = K_s$ where K_s is the solubility product. The total solubility, S_t is then equal to the true partial solubility, S_p, plus $[A^-]$ or

$$S_t = S_p + \frac{K_s}{[H^+]}, \text{ or } S_t = S_p\left[\frac{[H^+] + K_a}{[H^+]}\right]$$

If there is present no salt of the acid to furnish $[A^-]$

$$[H^+]^2 = K_s$$

or

$$pH = -\tfrac{1}{2}\log K_s$$

For the case of calcium carbonate, the $[CO_3^{-}{}^{-}]$ from which is controlled by $[H^+]$, see "Carbonate Equilibria."

References. See any text on physical chemistry and "Carbonate Equilibria," "Protein," "Equilibria," etc.

STAINING.

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SURFACE TENSION.

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SWEAT.

References. Clark-Lubs, (1917), Talbert (1919).

TAUTOMERISM other than of indicators.

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TISSUE CULTURE.

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URINE AND KIDNEY FUNCTIONS. The excretion of acids and bases in the urine is one of the mechanisms by which the hydrogen ion concentration of the blood is preserved constant. For this reason the determination of the acid-base equilibria in the urine in their relation to the potential acid-base intake in the food and the degree of oxidation of food material is of importance in fundamental physiological researches and in clinical studies. Besides references to be found under "Blood" the following are some of the more special references on urine.[2]

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VINEGAR.

Reference. Clark-Lubs (1917), Brode-Lange (1909), Kling-Lassieur-Lassieur (1922).

WATER (sea and fresh). The carbonate equilibrium maintains sea water at a very constant pH which has doubtless varied with the CO_2 tension of the atmosphere in geological ages and which varies somewhat with the temperature, and locally with accretions from rivers and springs and contact with geologic deposits. The wider aspects of the carbonate equilibria involved have been described in Henderson's *Fitness of the Environment.* The charting of the pH values for different regions of the seas has been of aid in oceanographic surveys and in some instances has been of value in the study of plant and animal distribution. (See "Plant Distribution" and "Comparative Physiology.")

[2] See Clark and Lubs (1917) for some examples of the application of the sulfon phthalein indicators to the determination of the pH of urines.

Fresh waters are influenced chiefiy by the deposits with which they come in contact. pH determinations in the field are of aid to the geologist in demarking waters of limestone origin (Wherry private communication).

In the clarification of water by "alum" or "iron" coagulation the pH of the final mix determines the percentage coagulant thrown out, the time required for floc formation and the efficiency of color- and turbidity-removal. There is also a probable relation to the efficiency of the filtration process itself.

The hydrogen ion enters into every equilibrium of importance to water softening and to corrosion.

References. Auerbach (1904), Baylis (1922), Buswell (1922), Corti-Alvarez (1918), Crozier (1920), Gaarder (1916–1917), Greenfield-Baker (1920), Haas (1916), Henderson (1913), Henderson-Cohn (1916), Heyman (1920), Kolthoff (1921), Loeb (1904), McClendon (1916, 1917), Mayer (1919), Massink (1920), Massink-Heyman (1921), Michaelis (1914, 1921), Palitzsch (1911, 1915, 1916), Powers (1921, 1922), Prideaux (1919), Ringer (1908), Ruppin (1909), Saunders (1921), Shelford (1919), Smith (1919), Snook (1915), Sørensen-Palitzsch (1910–13), Stephanides (1916), Tillmans (1919, 1921), Trillat (1916), Wagner-Enslow (1922), Walker-Kay (1912), Wells (1921), Wolman-Hannan (1921).

Water, pure. Ionization of.

References. Kohlrausch-Heydweiller (1894), Lewis, Brighton and Sebastian (1917), Nernst (1894), Ostwald (1893), Wijs (1893).

Wine Acidity. Besides influencing the fermentations, the pH of wine has been found to correlate in a general way with the acid taste.

References. Casale (1919), Dutoit-Dubroux (1910), Paul (1914, 1915, 1916), Quartaroli (1912).

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Knowledge is of two kinds. We know a subject ourselves, or we know where we can find information upon it.—SAMUEL JOHNSON.

The references in this bibliography are classified either by notations given at the end of each chapter or else by the subjects briefly outlined in Chapter XXI where cross references have been reduced to a minimum.

Abbreviations follow for the most part the system adopted by *Chemical Abstracts*.

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APPENDIX

TABLE A

STANDARD* VALUES FOR CALOMEL ELECTRODES

(Referred to the normal hydrogen electrode)

TEMPERATURE	CONCENTRATION OF KCl		
	M/10	M/1	Saturated (approximate potential)
°C.			
18	0.3380	0.2864	0.2506
20	0.3379	0.2860	0.2492
25	0.3376	0.2848	0.2464
30	0.3372	0.2836	0.2437
37.5	0.3364		
40	0.3360		

* See page 287.

TABLE B

SHOWING RELATION OF $[H^+]$ TO pH (ON THE ASSUMPTION THAT $pH = \text{Log } \frac{1}{[H^+]}$. See Chapter XIX)

pH	$[H^+]$	pH	$[H^+]$
x.00	1.00×10^{-x}	x.52	0.30×10^{-x}
x.02	0.96×10^{-x}	x.54	0.29×10^{-x}
x.04	0.91×10^{-x}	x.56	0.28×10^{-x}
x.06	0.87×10^{-x}	x.58	0.26×10^{-x}
x.08	0.83×10^{-x}	x.60	0.25×10^{-x}
x.10	0.80×10^{-x}	x.62	0.24×10^{-x}
x.12	0.76×10^{-x}	x.64	0.23×10^{-x}
x.14	0.73×10^{-x}	x.66	0.22×10^{-x}
x.16	0.69×10^{-x}	x.68	0.21×10^{-x}
x.18	0.66×10^{-x}	x.70	0.20×10^{-x}
x.20	0.63×10^{-x}	x.72	0.19×10^{-x}
x.22	0.60×10^{-x}	x.74	0.18×10^{-x}
x.24	0.58×10^{-x}	x.76	0.17×10^{-x}
x.26	0.55×10^{-x}	x.78	0.17×10^{-x}
x.28	0.53×10^{-x}	x.80	0.16×10^{-x}
x.30	0.50×10^{-x}	x.82	0.15×10^{-x}
x.32	0.48×10^{-x}	x.84	0.14×10^{-x}
x.34	0.46×10^{-x}	x.86	0.14×10^{-x}
x.36	0.44×10^{-x}	x.88	0.13×10^{-x}
x.38	0.42×10^{-x}	x.90	0.13×10^{-x}
x.40	0.40×10^{-x}	x.92	0.12×10^{-x}
x.42	0.38×10^{-x}	x.94	0.12×10^{-x}
x.44	0.36×10^{-x}	x.96	0.11×10^{-x}
x.46	0.35×10^{-x}	x.98	0.11×10^{-x}
x.48	0.33×10^{-x}	x + 1.00	0.10×10^{-x}
x.50	0.32×10^{-x}	x + 1.02	0.096×10^{-x}

Examples: pH = 7.00; $[H^+] = 1 \times 10^{-7}$

pH = 8.60; $[H^+] = 0.25 \times 10^{-8} = 2.5 \times 10^{-9}$

$[H^+] = 5.3 \times 10^{-5} = 0.53 \times 10^{-4}$; pH = 4.28

Compare Symes (1916).

TABLE C

TEMPERATURE FACTORS FOR CONCENTRATION CHAINS

$$E = 0.000,198,37\ T \log \frac{C_1}{C_2} \text{ (when valence = 1)}$$

t (CENTIGRADE)	T (ABSOLUTE)	0.000,198,37 T	LOG 0.000,198,37 T
0	273.09	0.05416	$\bar{2}$.73364
5	278.09	0.05515	$\bar{2}$.74152
10	283.09	0.05614	$\bar{2}$.74927
15	288.09	0.05713	$\bar{2}$.75687
16	289.09	0.05733	$\bar{2}$.75838
17	290.09	0.05753	$\bar{2}$.75988
18	291.09	0.05772	$\bar{2}$.76137
19	292.09	0.05792	$\bar{2}$.76286
20	293.09	0.05812	$\bar{2}$.76435
21	294.09	0.05832	$\bar{2}$.76583
22	295.09	0.05852	$\bar{2}$.76730
23	296.09	0.05872	$\bar{2}$.76877
24	297.09	0.05892	$\bar{2}$.77023
25	298.09	0.05911	$\bar{2}$.77169
26	299.09	0.05931	$\bar{2}$.77315
27	300.09	0.05951	$\bar{2}$.77460
28	301.09	0.05971	$\bar{2}$.77604
29	302.09	0.05991	$\bar{2}$.77749
30	303.09	0.06011	$\bar{2}$..77892
31	304.09	0.06031	$\bar{2}$.78035
32	305.09	0.06050	$\bar{2}$.78178
33	306.09	0.06070	$\bar{2}$.78320
34	307.09	0.06090	$\bar{2}$.78462
35	308.09	0.06110	$\bar{2}$.78603
36	309.09	0.06130	$\bar{2}$.78742
37	310.09	0.06150	$\bar{2}$.78884
37.5	310.59	0.06159	$\bar{2}$.78954
38	311.09	0.06169	$\bar{2}$.79024
39	312.09	0.06189	$\bar{2}$.79163
40	313.09	0.06209	$\bar{2}$.79302
45	318.09	0.06308	$\bar{2}$.79990
50	323.09	0.06407	$\bar{2}$.80668

TABLE D

CORRECTION OF BAROMETER READING FOR TEMPERATURE

When the mercury in the barometer is at the temperature t subtract the following corrections to obtain the barometric height in terms of mercury at zero degrees centigrade.

t	BAROMETER READINGS IN MILLIMETERS						
	720	730	740	750	760	770	780
17	2.0	2.0	2.1	2.1	2.1	2.1	2.2
18	2.1	2.1	2.2	2.2	2.2	2.3	2.3
19	2.2	2.3	2.3	2.3	2.4	2.4	2.4
20	2.3	2.4	2.4	2.4	2.5	2.5	2.5
21	2.5	2.5	2.5	2.6	2.6	2.6	2.7
22	2.6	2.6	2.7	2.7	2.7	2.8	2.8
23	2.7	2.7	2.8	2.8	2.8	2.9	2.9
24	2.8	2.9	2.9	2.9	3.0	3.0	3.1
25	2.9	3.0	3.0	3.1	3.1	3.1	3.2
26	3.0	3.1	3.1	3.2	3.2	3.3	3.3
27	3.2	3.2	3.3	3.3	3.3	3.4	3.4
28	3.3	3.3	3.4	3.4	3.5	3.5	3.6
29	3.4	3.4	3.5	3.5	3.6	3.6	3.7
30	3.5	3.6	3.6	3.7	3.7	3.8	3.8
31	3.6	3.7	3.7	3.8	3.8	3.9	3.9

TABLE E

BAROMETRIC CORRECTIONS FOR H-ELECTRODE POTENTIALS

(Data for use in plotting correction curves)

$$E_{bar.} = \frac{0.000{,}19837\ T}{2} \log \frac{760}{x}$$

TEMPERATURE	CORRECTED PRESSURE	VAPOR PRESSURE	x	LOG $\frac{760}{x}$	$E_{bar.}$
°C.	*mm.*	*mm.*			*millivolts*
18	780	15.5	764.5	−0.00256	−0.07
	760		744.5	0.00895	0.26
	740		724.5	0.02078	0.60
20	780	17.5	762.5	−0.00143	−0.04
	760		742.5	0.01012	0.29
	740		722.5	0.02198	0.64
25	780	23.8	756.2	0.00218	0.06
	760		736.2	0.01382	0.41
	740		716.2	0.02578	0.76
30	780	31.8	748.2	0.00680	0.20
	760		728.2	0.01856	0.56
	740		708.2	0.03066	0.92
35	780	42.2	737.8	0.01288	0.39
	760		717.8	0.02481	0.76
	740		697.8	0.03708	1.13
40	780	55.3	724.8	0.02060	0.64
	760		704.8	0.03275	1.02
	740		684.7	0.04525	1.41

$$\frac{\text{E. M. F.} + E_{bar.} - E_{cal.}}{0.000{,}19837\ T} = \text{pH}$$

TABLE F

VALUES OF LOG $\frac{\alpha}{1-\alpha}$ AND OF LOG $\frac{\alpha}{1-\alpha}$ MULTIPLIED BY THE TEMPERATURE FACTORS FOR CONCENTRATION CHAINS AT 20°, 25°, 30° AND 37.5°C.

α	LOG $\frac{\alpha}{1-\alpha}$	LOG $\frac{\alpha}{1-\alpha}$ MULTIPLIED BY			
		0.05812 (20)	0.05911 (25)	0.06011 (30)	0.06159 (37.5)
0.001	−2.9996	−0.1743	−0.1773	−0.1803	−0.1846
0.005	−2.2989	−0.1336	−0.1359	−0.1382	−0.1416
0.01	−1.9956	−0.1156	−0.1180	−0.1200	−0.1229
0.02	−1.6902	−0.0982	−0.0999	−0.1016	−0.1041
0.03	−1.5096	−0.0879	−0.0892	−0.0908	−0.0930
0.04	−1.3802	−0.0802	−0.0816	−0.0830	−0.0850
0.05	−1.2787	−0.0743	−0.0756	−0.0769	−0.0788
0.06	−1.1950	−0.0695	−0.0706	−0.0718	−0.0736
0.07	−1.1234	−0.0653	−0.0664	−0.0675	−0.0692
0.08	−1.0607	−0.0617	−0.0627	−0.0638	−0.0653
0.09	−1.0048	−0.0584	−0.0594	−0.0604	−0.0619
0.10	−0.9542	−0.0555	−0.0564	−0.0574	−0.0588
0.11	−0.9080	−0.0528	−0.0537	−0.0546	−0.0559
0.12	−0.8653	−0.0503	−0.0512	−0.0520	−0.0533
0.13	−0.8256	−0.0480	−0.0488	−0.0496	−0.0509
0.14	−0.7884	−0.0458	−0.0466	−0.0474	−0.0486
0.15	−0.7533	−0.0438	−0.0445	−0.0453	−0.0464
0.16	−0.7202	−0.0419	−0.0426	−0.0433	−0.0443
0.17	−0.6886	−0.0400	−0.0407	−0.0414	−0.0424
0.18	−0.6585	−0.0383	−0.0389	−0.0396	−0.0406
0.19	−0.6297	−0.0366	−0.0372	−0.0379	−0.0388
0.20	−0.6021	−0.0350	−0.0356	−0.0362	−0.0371
0.21	−0.5754	−0.0334	−0.0340	−0.0346	−0.0354
0.22	−0.5497	−0.0320	−0.0325	−0.0330	−0.0339
0.23	−0.5248	−0.0305	−0.0310	−0.0315	−0.0323

TABLE F—*Continued*

VALUES OF LOG $\frac{\alpha}{1-\alpha}$ AND OF LOG $\frac{\alpha}{1-\alpha}$ MULTIPLIED BY THE TEMPERATURE FACTORS FOR CONCENTRATION CHAINS at 20°, 25°, 30° AND 37.5°C.

	LOG $\frac{\alpha}{1-\alpha}$	LOG $\frac{\alpha}{1-\alpha}$ MULTIPLIED BY			
		0.05812 (20)	0.05911 (25)	0.06011 (30)	0.06159 (37.5)
0.24	−0.5006	−0.0291	−0.0296	−0.0301	−0.0309
0.25	−0.4771	−0.0277	−0.0281	−0.0287	−0.0294
0.26	−0.4543	−0.0264	−0.0269	−0.0273	−0.0280
0.27	−0.4320	−0.0251	−0.0255	−0.0260	−0.0266
0.28	−0.4102	−0.0238	−0.0243	−0.0247	−0.0253
0.29	−0.3889	−0.0226	−0.0230	−0.0234	−0.0239
0.30	−0.3680	−0.0214	−0.0218	−0.0221	−0.0227
0.31	−0.3475	−0.0202	−0.0205	−0.0209	−0.0214
0.32	−0.3274	−0.0190	−0.0193	−0.0197	−0.0202
0.33	−0.3076	−0.0179	−0.0182	−0.0185	−0.0190
0.34	−0.2881	−0.0167	−0.0170	−0.0173	−0.0178
0.35	−0.2688	−0.0156	−0.0159	−0.0162	−0.0166
0.36	−0.2499	−0.0145	−0.0148	−0.0150	−0.0154
0.37	−0.2311	−0.0134	−0.0137	−0.0139	−0.0142
0.38	−0.2126	−0.0124	−0.0126	−0.0127	−0.0131
0.39	−0.1943	−0.0113	−0.0115	−0.0117	−0.0120
0.40	−0.1761	−0.0102	−0.0104	−0.0106	−0.0109
0.41	−0.1581	−0.0092	−0.0094	−0.0095	−0.0098
0.42	−0.1402	−0.0082	−0.0083	−0.0084	−0.0086
0.43	−0.1224	−0.0071	−0.0072	−0.0074	−0.0075
0.44	−0.1047	−0.0061	−0.0062	−0.0063	−0.0065
0.45	−0.0872	−0.0050	−0.0052	−0.0052	−0.0054
0.46	−0.0696	−0.0040	−0.0041	−0.0042	−0.0043
0.47	−0.0522	−0.0030	−0.0031	−0.0031	−0.0032
0.48	−0.0348	−0.0020	−0.0021	−0.0021	−0.0021
0.49	−0.0174	−0.0010	−0.0010	−0.0011	−0.0011
0.50	±0.0000	±0.0000	±0.0000	±0.0000	±0.0000
0.51	+0.0174	+0.0010	+0.0010	+0.0011	+0.0011
0.52	+0.0348	+0.0020	+0.0021	+0.0021	+0.0021

For values beyond $\alpha = 0.50$ the table progresses inversely as above but with sign +. Example: $\alpha = 0.53$, $(1 - \alpha = 0.47)$, read row for $\alpha = 0.47$, i.e., $\log \frac{\alpha}{1-\alpha} = +0.0522$, etc. If $\alpha = 0.80$, $(1 - \alpha = 0.20)$, read row for $\alpha = 0.20$, i.e., $\log \frac{\alpha}{1-\alpha} = +0.6021$, etc.

TABLE G

Ionization Constants

The following lists, compiled from Scudder (1914) and miscellaneous sources are for purposes of illustration and approximate uses only. $pK_a = \log \frac{1}{K_a}$.

	K_a	pK_a
Acetic	1.8×10^{-5}	4.7
Alloxan	2.3×10^{-7}	6.6
Arsenic	5.0×10^{-3}	2.3
Arsenious	6.0×10^{-10}	9.2
Benzoic	6.0×10^{-5}	4.2
Boric	6.5×10^{-10}	9.2
Butyric	1.6×10^{-5}	4.8
Cacodylic	6.4×10^{-7}	6.2
Carbonic, first	3.4×10^{-7}	6.5
Carbonic, second	4.8×10^{-11}	10.3
Citric, first	8.2×10^{-4}	3.1
Citric, second	3.2×10^{-5}	4.5
Citric, third	7.0×10^{-7}	6.2?
Formic	2.1×10^{-4}	3.7
Fructose	6.6×10^{-13}	12.2
Glucose	3.6×10^{-13}	12.4
Glutamic	4.1×10^{-5}	4.4
Glycerine	7.0×10^{-15}	14.2
Hippuric	2.2×10^{-4}	3.6
Hydrogen sulfid	5.7×10^{-8}	7.2
Lactic	1.4×10^{-4}	3.9
Mucic	6.3×10^{-4}	3.2
Nitrous	6.0×10^{-4}	3.2
Oxalic, first	1.0×10^{-1}	1.0
Oxalic, second	4.1×10^{-5}	4.4
β-Oxybutyric	3.9×10^{-5}	4.4
Phosphoric acid, first	1.0×10^{-2}	2.0
Phosphoric acid, second	8.8×10^{-8}	7.1
Phosphoric acid, third	3.6×10^{-13}	12.4
Phthalic, first	1.2×10^{-3}	2.9
Phthalic, second	3.9×10^{-6}	5.4
Phenol	1.0×10^{-10}	10.0
Salicylic, first	1.0×10^{-3}	3.0
Salicylic, second	?	
Succinic, first	6.8×10^{-5}	4.2
Succinic, second	2.7×10^{-6}	5.6
Sucrose	1.1×10^{-13}	13.0
Sulfurous	6.0×10^{-4}	3.3
Tartaric, first	9.7×10^{-4}	3.1
Tartaric, second	4.5×10^{-6}	4.4
Uric	1.5×10^{-6}	5.8

Bases

$$pK_b = \log \frac{1}{K_b}$$

	K_b	pK_b	$14 - pK_b$
Ammonium	1.8×10^{-5}	4.8	9.2
Aniline	4.6×10^{-10}	9.3	4.7
Caffeine	4.1×10^{-11}	10.4	3.6
Creatinine	3.7×10^{-9}	8.6	5.4
Diethyl amine	1.3×10^{-3}	2.9	11.1
Ethyl amine	5.6×10^{-4}	3.3	10.7
Guanine	8.4×10^{-12}	11.1	2.9
Pyridine	2.3×10^{-14}	13.6	0.4
Urea	1.5×10^{-14}	13.8	0.2
Xanthine	4.8×10^{-14}	13.3	0.7

Ampholytes

	K_a	pK_a	K_b	$14 - pK_b$
Alanine	2.0×10^{-10}	9.7	3.0×10^{-12}	2.5
Asparagine	1.4×10^{-9}	8.9	1.5×10^{-12}	2.2
Aspartic acid	1.4×10^{-4}	3.9	1.2×10^{-12}	2.1
Glycine	1.8×10^{-10}	9.8	2.8×10^{-12}	2.4
Histidine, first	2.2×10^{-9}	8.7	5.7×10^{-9}	5.8
Histidine, second			5.0×10^{-13}	1.7
Lysine	1.0×10^{-11}	11.0	1.0×10^{-7}	7.0
Tyrosine, first	4.0×10^{-9}	8.4	2.6×10^{-12}	2.4
Tyrosine, second	4.0×10^{-10}	9.4		

NATURAL NUMBERS	0	1	2	3	4	5	6	7	8	9	PROPORTIONAL PARTS								
											1	2	3	4	5	6	7	8	9
10	0000	0043	0086	0128	0170	0212	0253	0294	0334	0374	4	8	12	17	21	25	29	33	37
11	0414	0453	0492	0531	0569	0607	0645	0682	0719	0755	4	8	11	15	19	23	26	30	34
12	0792	0828	0864	0899	0934	0969	1004	1038	1072	1106	3	7	10	14	17	21	24	28	31
13	1139	1173	1206	1239	1271	1303	1335	1367	1399	1430	3	6	10	13	16	19	23	26	29
14	1461	1492	1523	1553	1584	1614	1644	1673	1703	1732	3	6	9	12	15	18	21	24	27
15	1761	1790	1818	1847	1875	1903	1931	1959	1987	2014	3	6	8	11	14	17	20	22	25
16	2041	2068	2095	2122	2148	2175	2201	2227	2253	2279	3	5	8	11	13	16	18	21	24
17	2304	2330	2355	2380	2405	2430	2455	2480	2504	2529	2	5	7	10	12	15	17	20	22
18	2553	2577	2601	2625	2648	2672	2695	2718	2742	2765	2	5	7	9	12	14	16	19	21
19	2788	2810	2833	2856	2878	2900	2923	2945	2967	2989	2	4	7	9	11	13	16	18	20
20	3010	3032	3054	3075	3096	3118	3139	3160	3181	3201	2	4	6	8	11	13	15	17	19
21	3222	3243	3263	3284	3304	3324	3345	3365	3385	3404	2	4	6	8	10	12	14	16	18
22	3424	3444	3464	3483	3502	3522	3541	3560	3579	3598	2	4	6	8	10	12	14	15	17
23	3617	3636	3655	3674	3692	3711	3729	3747	3766	3784	2	4	6	7	9	11	13	15	17
24	3802	3820	3838	3856	3874	3892	3909	3927	3945	3962	2	4	5	7	9	11	12	14	16
25	3979	3997	4014	4031	4048	4065	4082	4099	4116	4133	2	3	5	7	9	10	12	14	15
26	4150	4166	4183	4200	4216	4232	4249	4265	4281	4298	2	3	5	7	8	10	11	13	15
27	4314	4330	4346	4362	4378	4393	4409	4425	4440	4456	2	3	5	6	8	9	11	13	14
28	4472	4487	4502	4518	4533	4548	4564	4579	4594	4609	2	3	5	6	8	9	11	12	14
29	4624	4639	4654	4669	4683	4698	4713	4728	4742	4757	1	3	4	6	7	9	10	12	13
30	4771	4786	4800	4814	4829	4843	4857	4871	4886	4900	1	3	4	6	7	9	10	11	13
31	4914	4928	4942	4955	4969	4983	4997	5011	5024	5038	1	3	4	6	7	8	10	11	12
32	5052	5065	5079	5092	5105	5119	5132	5145	5159	5172	1	3	4	5	7	8	9	11	12
33	5185	5198	5211	5224	5237	5250	5263	5276	5289	5302	1	3	4	5	6	8	9	10	12
34	5315	5328	5340	5353	5366	5378	5391	5403	5416	5428	1	3	4	5	6	8	9	10	11
35	5441	5453	5465	5478	5490	5502	5514	5527	5539	5551	1	2	4	5	6	7	9	10	11
36	5563	5575	5587	5599	5611	5623	5635	5647	5658	5670	1	2	4	5	6	7	8	10	11
37	5682	5694	5705	5717	5729	5740	5752	5763	5775	5786	1	2	3	5	6	7	8	9	10
38	5798	5809	5821	5832	5843	5855	5866	5877	5888	5899	1	2	3	5	6	7	8	9	10
39	5911	5922	5933	5944	5955	5966	5977	5988	5999	6010	1	2	3	4	5	7	8	9	10
40	6021	6031	6042	6053	6064	6075	6085	6096	6107	6117	1	2	3	4	5	6	8	9	10
41	6128	6138	6149	6160	6170	6180	6191	6201	6212	6222	1	2	3	4	5	6	7	8	9
42	6232	6243	6253	6263	6274	6284	6294	6304	6314	6325	1	2	3	4	5	6	7	8	9
43	6335	6345	6355	6365	6375	6385	6395	6405	6415	6425	1	2	3	4	5	6	7	8	9
44	6435	6444	6454	6464	6474	6484	6493	6503	6513	6522	1	2	3	4	5	6	7	8	9
45	6532	6542	6551	6561	6571	6580	6590	6599	6609	6618	1	2	3	4	5	6	7	8	9
46	6628	6637	6646	6656	6665	6675	6684	6693	6702	6712	1	2	3	4	5	6	7	7	8
47	6721	6730	6739	6749	6758	6767	6776	6785	6794	6803	1	2	3	4	5	5	6	7	8
48	6812	6821	6830	6839	6848	6857	6866	6875	6884	6893	1	2	3	4	4	5	6	7	8
49	6902	6911	6920	6928	6937	6946	6955	6964	6972	6981	1	2	3	4	4	5	6	7	8
50	6990	6998	7007	7016	7024	7033	7042	7050	7059	7067	1	2	3	3	4	5	6	7	8
51	7076	7084	7093	7101	7110	7118	7126	7135	7143	7152	1	2	3	3	4	5	6	7	8
52	7160	7168	7177	7185	7193	7202	7210	7218	7226	7235	1	2	2	3	4	5	6	7	7
53	7243	7251	7259	7267	7275	7284	7292	7300	7308	7316	1	2	2	3	4	5	6	6	7
54	7324	7332	7340	7348	7356	7364	7372	7380	7388	7396	1	2	2	3	4	5	6	6	7

NATURAL NUMBERS	0	1	2	3	4	5	6	7	8	9	PROPORTIONAL PARTS								
											1	2	3	4	5	6	7	8	9
55	7404	7412	7419	7427	7435	7443	7451	7459	7466	7474	1	2	2	3	4	5	5	6	7
56	7482	7490	7497	7505	7513	7520	7528	7536	7543	7551	1	2	2	3	4	5	5	6	7
57	7559	7566	7574	7582	7589	7597	7604	7612	7619	7627	1	2	2	3	4	5	5	6	7
58	7634	7642	7649	7657	7664	7672	7679	7686	7694	7701	1	1	2	3	4	4	5	6	7
59	7709	7716	7723	7731	7738	7745	7752	7760	7767	7774	1	1	2	3	4	4	5	6	7
60	7782	7789	7796	7803	7810	7818	7825	7832	7839	7846	1	1	2	3	4	4	5	6	6
61	7853	7860	7868	7875	7882	7889	7896	7903	7910	7917	1	1	2	3	4	4	5	6	6
62	7924	7931	7938	7945	7952	7959	7966	7973	7980	7987	1	1	2	3	3	4	5	6	6
63	7993	8000	8007	8014	8021	8028	8035	8041	8048	8055	1	1	2	3	3	4	5	5	6
64	8062	8069	8075	8082	8089	8096	8102	8109	8116	8122	1	1	2	3	3	4	5	5	6
65	8129	8136	8142	8149	8156	8162	8169	8176	8182	8189	1	1	2	3	3	4	5	5	6
66	8195	8202	8209	8215	8222	8228	8235	8241	8248	8254	1	1	2	3	3	4	5	5	6
67	8261	8267	8274	8280	8287	8293	8299	8306	8312	8319	1	1	2	3	3	4	5	5	6
68	8325	8331	8338	8344	8351	8357	8363	8370	8376	8382	1	1	2	3	3	4	4	5	6
69	8388	8395	8401	8407	8414	8420	8426	8432	8439	8445	1	1	2	2	3	4	4	5	6
70	8451	8457	8463	8470	8476	8482	8488	8494	8500	8506	1	1	2	2	3	4	4	5	6
71	8513	8519	8525	8531	8537	8543	8549	8555	8561	8567	1	1	2	2	3	4	4	5	5
72	8573	8579	8585	8591	8597	8603	8609	8615	8621	8627	1	1	2	2	3	4	4	5	5
73	8633	8639	8645	8651	8657	8663	8669	8675	8681	8686	1	1	2	2	3	4	4	5	5
74	8692	8698	8704	8710	8716	8722	8727	8733	8739	8745	1	1	2	2	3	4	4	5	5
75	8751	8756	8762	8768	8774	8779	8785	8791	8797	8802	1	1	2	2	3	3	4	5	5
76	8808	8814	8820	8825	8831	8837	8842	8848	8854	8859	1	1	2	2	3	3	4	5	5
77	8865	8871	8876	8882	8887	8893	8899	8904	8910	8915	1	1	2	2	3	3	4	4	5
78	8921	8927	8932	8938	8943	8949	8954	8960	8965	8971	1	1	2	2	3	3	4	4	5
79	8976	8982	8987	8993	8998	9004	9009	9015	9020	9025	1	1	2	2	3	3	4	4	5
80	9031	9036	9042	9047	9053	9058	9063	9069	9074	9079	1	1	2	2	3	3	4	4	5
81	9085	9090	9096	9101	9106	9112	9117	9122	9128	9133	1	1	2	2	3	3	4	4	5
82	9138	9143	9149	9154	9159	9165	9170	9175	9180	9186	1	1	2	2	3	3	4	4	5
83	9191	9196	9201	9206	9212	9217	9222	9227	9232	9238	1	1	2	2	3	3	4	4	5
84	9243	9248	9253	9258	9263	9269	9274	9279	9284	9289	1	1	2	2	3	3	4	4	5
85	9294	9299	9304	9309	9315	9320	9325	9330	9335	9340	1	1	2	2	3	3	4	4	5
86	9345	9350	9355	9360	9365	9370	9375	9380	9385	9390	1	1	2	2	3	3	4	4	5
87	9395	9400	9405	9410	9415	9420	9425	9430	9435	9440	0	1	1	2	2	3	3	4	4
88	9445	9450	9455	9460	9465	9469	9474	9479	9484	9489	0	1	1	2	2	3	3	4	4
89	9494	9499	9504	9509	9513	9518	9523	9528	9533	9538	0	1	1	2	2	3	3	4	4
90	9542	9547	9552	9557	9562	9566	9571	9576	9581	9586	0	1	1	2	2	3	3	4	4
91	9590	9595	9600	9605	9609	9614	9619	9624	9628	9633	0	1	1	2	2	3	3	4	4
92	9638	9643	9647	9652	9657	9661	9666	9671	9675	9680	0	1	1	2	2	3	3	4	4
93	9685	9689	9694	9699	9703	9708	9713	9717	9722	9727	0	1	1	2	2	3	3	4	4
94	9731	9736	9741	9745	9750	9754	9759	9763	9768	9773	0	1	1	2	2	3	3	4	4
95	9777	9782	9786	9791	9795	9800	9805	9809	9814	9818	0	1	1	2	2	3	3	4	4
96	9823	9827	9832	9836	9841	9845	9850	9854	9859	9863	0	1	1	2	2	3	3	4	4
97	9868	9872	9877	9881	9886	9890	9894	9899	9903	9908	0	1	1	2	2	3	3	4	4
98	9912	9917	9921	9926	9930	9934	9939	9943	9948	9952	0	1	1	2	2	3	3	4	4
99	9956	9961	9965	9969	9974	9978	9983	9987	9991	9996	0	1	1	2	2	3	3	3	4

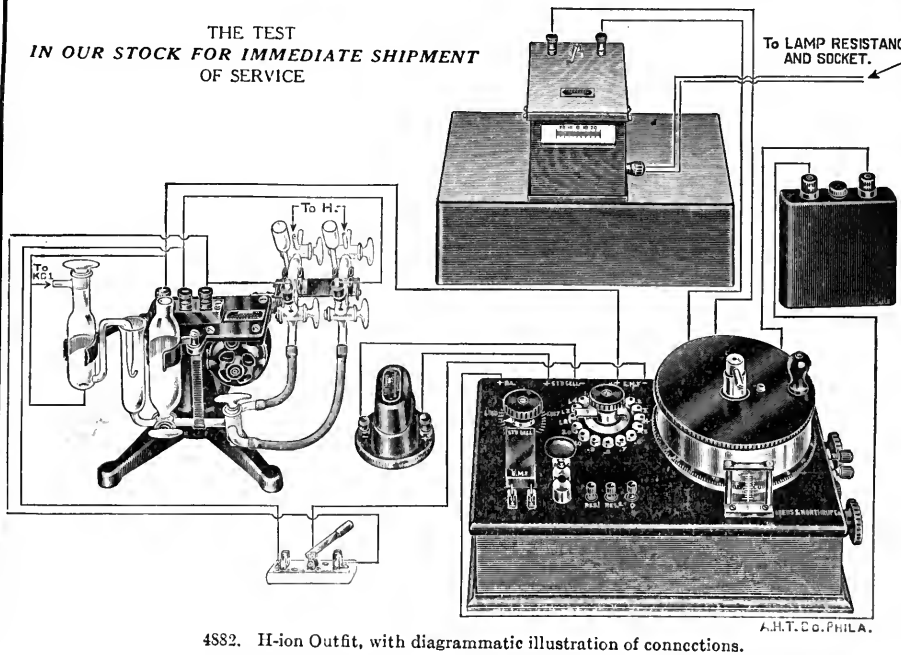
INDEX TO AUTHORS

INDEX TO SUBJECTS

SCIENTIFIC SUPPLIES

ELECTROMETRIC H-ION OUTFIT

THE TEST
IN OUR STOCK FOR IMMEDIATE SHIPMENT
OF SERVICE



4882. H-ion Outfit, with diagrammatic illustration of connections.

This outfit provides for the measurement of H-ion concentration in solutions with a precision sufficient for most researches in bacteriology, physiology and biochemistry. It is also adapted to routine measurements since no adjustments requiring high manipulative skill are involved.

The arrangement is as recommended and used by competent investigators and is based on the Clark Hydrogen Electrode—the most widely used for biological solutions—and upon the Leeds & Northrup Type K Potentiometer and Galvanometer of the enclosed lamp and scale type. This combination of Potentiometer and Galvanometer is recognized as standard throughout the world and measures the difference in potential between the calomel and hydrogen electrode to from .02 to .04% (± .2 millivolts) which in ordinary solutions is the equivalent of 0.001 of the pH unit.

The outfit is as follows, the component parts of which are sold separately under their respective catalogue numbers. The prices indicated are those in effect May 15th, 1923, and are subject to change without notice:—

4790.	Type K Potentiometer, L. & N..	275.00	4820.	Weston Standard Cell, Model 4.....	25.00
4796.	Lamp and Scale Galvanometer, L. & N.....	50.00	2092.	Storage Battery	6.00
4797.	Lamp Resistance for 4796, for 110 volts	5.00	4854.	Clark Hydrogen Electrode Vessels (2)..	18.00
4878.	Combined Support and Shaker, L. & N.....	75.00	4855.	Platinum Electrodes for same (2).....	5.50
			4834.	Calomel Electrode Vessel.....	3.85
			4866.	Clark Connecting Vessel.....	6.50
			9374.	Switch, single pole double throw.....	.40

			Code Word
4882.	**Complete Outfit,** as above listed, with motor for 110 volts, 60 cycle, a.c.....	470.00	*Ezimq*
4883.	ditto, but with motor for 110 volt d.c.....	470.00	*Ezjef*

The new 828-page edition of our catalogue LABORATORY APPARATUS AND REAGENTS lists a modern and complete selection of apparatus for both the electrometric and the colorimetric determination of Hydrogen ions, as well as for other measurements of small potential differences and electrolytic conductivities, all of which is carried in our stock for immediate shipment at manufacturers' prices.

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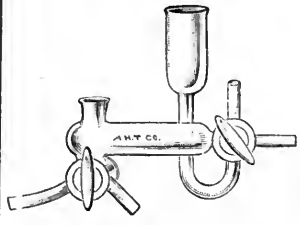
LABORATORY APPARATUS AND REAGENTS
WEST WASHINGTON SQUARE — CABLE ADDRESS, "BALANCE" — PHILADELPHIA, U. S. A.

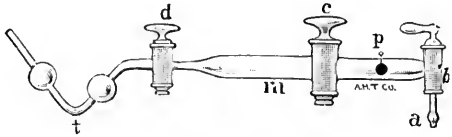
HYDROGEN ELECTRODES

THE TEST
IN OUR STOCK FOR IMMEDIATE SHIPMENT
OF SERVICE

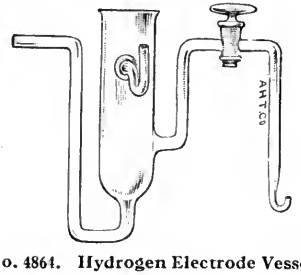


No. 4854. Clark Hydrogen Electrode Vessel

No. 4862
Bailey Hydrogen Electrode

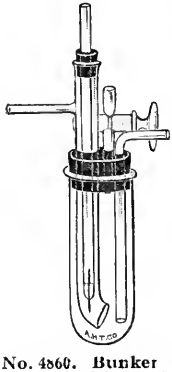


No. 4858. McClendon Hydrogen Electrode



No. 4864. Hydrogen Electrode Vessel

No. 4852. Hildebrand Hydrogen Electrode



No. 4860. Bunker Hydrogen Electrode

			Code Word
4854.	**Hydrogen Electrode Vessel, Clark,** without metallic electrode. See *Journal of Biological Chemistry,* ***23:*** *475, 1915*	9.00	*Eytyi*
4855.	**Platinum Electrode,** for use with Clark Hydrogen Electrode Vessel	2.75	*Eyuob*
4862.	**Hydrogen Electrode, Bailey,** complete with gold electrode and platinum wire connection. See *Journal of the American Chemical Society,* ***42:*** *45, 1920*	3.00	*Eywoz*
4858.	**Hydrogen Electrode, McClendon,** complete with gold electrode and platinum wire connection. See *Journal of Biological Chemistry,* ***25:*** *669, 1916*	9.00	*Eyvev*
4864.	**Hydrogen Electrode Vessel,** improved Ostwald, without platinum electrode	4.00	*Eyxoy*
4852.	**Hydrogen Electrode, Hildebrand,** with S-shaped platinum electrode, but without metallic adapter shown in illustration. See *Journal of American Chemical Society,* ***35:*** *847, 1913*	6.00	*Eysnf*
4860.	**Hydrogen Electrode, Bunker,** complete with platinum electrode, stopcock and rubber stoppers. See *Journal of Biological Chemistry,* ***41:*** *11, 1920*	4.75	*Eyvtq*

The above prices are those in effect May 15th, 1923 and are subject to change without notice

The Colorimetric Determination of H-ions

We list and stock at manufacturers' prices the Clark and Lubs Indicators, Standardized Buffer Solutions, Color Standards, H-ion Comparator Set, Indicator Field Set, etc., as prepared in the laboratories of Hynson, Westcott and Dunning, and the Lamotte Chemical Products Company.

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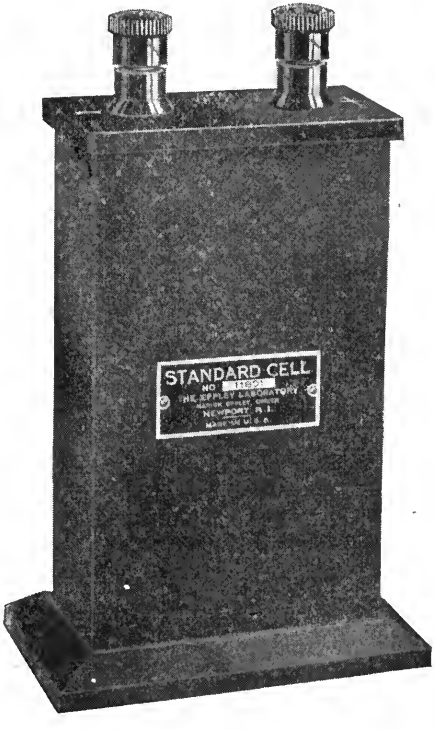
THE EPPLEY STANDARD CELL

A precision standard of highest accuracy



An unsaturated cadmium standard cell (Weston Type) accurate to 0.02% with a negligible temperature coefficient.

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For The Determination Of

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Of

BACTERIOLOGICAL CULTURE MEDIA

And For

Other Bio-Chemical Investigations

As recommended by William Mansfield Clark and Herbert A. Lubs, Dairy Division, Bureau of Animal Industry, U. S. Department of Agriculture. See the Journal of Bacteriology, January, March, May, 1917, Volume II, Nos. 1, 2 and 3.

COVERING A CONTINUOUS RANGE FROM P_H 1.2–P_H 9.8

Trade Name	Chemical Name	P_H Range
Thymol Blue	Thymolsulphonephthalein—Acid Range	1.2—2.8
Brom-Phenol Blue	Tetrabromphenolsulphonephthalein	2.8—4.6
Methyl Red	Orthocarboxybenzeneazodimethylaniline	4.4—6.0
Propyl Red	Orthocarboxybenzeneazodipropylaniline	4.8—6.4
Brom-Cresol Purple	Dibromocresolsulphonephthalein	5.2—6.8
Brom-Thymol Blue	Dibromthymolsulphonephthalein	6.0—7.6
Phenol Red	Phenolsulphonephthalein	6.8—8.4
Cresol Red	Orthocresolsulphonephthalein	7.2—8.8
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Colorimetric Determination of Hydrogen-Ion Concentration

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METHYL RED	Red-Yellow	4.4-6.0
BROMCRESOL PURPLE	Yellow-Purple	5.2-6.8
BROMTHYMOL BLUE	Yellow-Blue	6.0-7.6
PHENOL RED	Yellow-Red	6.8-8.4
CRESOL RED	Yellow-Red	7.2-8.8
THYMOL BLUE (alkaline range)	Yellow-Blue	8.0-9.6
CRESOL PHTHALEIN	Colorless Red	8.2-9.6
PHENOL PHTHALEIN	Colorless Red	8.4-9.2

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POTASSIUM CHLORIDE—BORIC ACID MIXTURE

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D—COLOR STANDARDS: For comparing with unknowns. Supplied in 10cc ampoule sets covering any part of the range from pH 1.0 to 10.0.

E—LaMOTTE H-ION COMPARATOR SET: A complete compact outfit for accurate determination of H-Ion concentration. Equally useful in both laboratory and field work. The set consists of comparator and case, color standards, prepared indicator, graduated test-tubes and pipette together with booklet containing full instructions for use.

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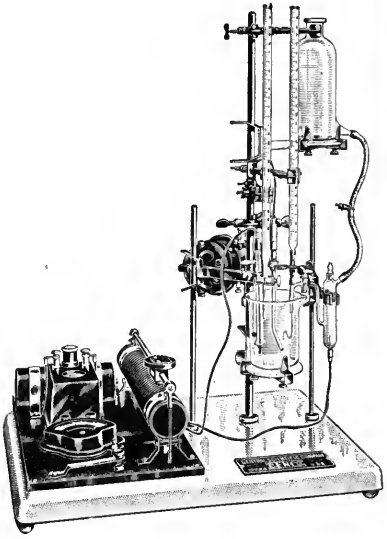

No. 10312

The **Improved Wendt Apparatus** was developed by our manufacturing department from the original apparatus suggested by Dr. Gerald L. Wendt, of the University of Chicago, to provide a simple and economical outfit with all parts arranged and connected so as to enable accurate results to be obtained free from the errors of troublesome current leaks and which could be readily used by the average analytical chemist or bacteriologist, unfamiliar with the complicated and sensitive apparatus used by the physical chemist for his more precise determinations.

The **galvanometer** and **millivoltmeter** were especially designed to enable readings to be made to 0.1 PH value over a range wide enough to permit measurements to be made on alkaline solutions as well as acid. The **galvanometer** is of rugged construction and high sensibility responding readily to extremely small potentials; one volt through 2,500,000 ohms giving a deflection of one scale division. The **millivoltmeter** has a scale 110 mm. long graduated from 0 to 1200 millivolts by 10 millivolt divisions and is provided with a mirror scale and knife edge pointer, which enables the operator to estimate readings to 1 millivolt. These instruments are mounted on a **Bakelite instrument board** elevated ½ inch above the section of the base under the burettes, thus eliminating the possibility of conducting liquids seeping under the instruments and causing short circuits.

The **glazed, porcelain covered cast iron base,** in addition to rendering the apparatus more stable, presents a smooth white surface against which color changes may be instantly observed when checking colored indicator solutions, such as the Clark and Lubs indicators. It does not absorb reagents, is unaffected by practically all ordinary reagents, and is easily cleaned. The base is mounted on rubber feet.

10312. **ELECTRO-TITRATION APPARATUS Wendt's,** with white enameled iron base and with galvanometer, voltmeter, rheostat, battery, key and switch assembled on a Bakelite table, with heavy central rod to support burette, reservoir, beaker and motor stirrer; and with lighter rods for supporting the gas and calomel electrodes. All metal parts are polished and nickel plated. The apparatus is furnished complete, with glass reservoir, hydrogen and calomel electrodes, and with necessary clamps and supports but without burettes, burette clamps, beaker or motor stirrer.....$150.00

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