semicarbazones. *p*-Methoxy-benzoylacrylic acid reacts with semicarbazide to form a semicarbazone.

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THE INFLUENCE OF HYDROGEN CONCENTRATION ON THE AUTO-OXIDATION OF HYDROQUINONE. A NOTE ON THE STABILITY OF THE QUINHYDRONE ELECTRODE

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It is well known from the work of Manchot,[1] and of Euler and Bolin[2] that the addition of alkali to a solution of hydroquinone causes the latter to act as an auto-oxidizable substance according to the equation: $C_6H_4(OH)_2 + O_2 \longrightarrow C_6H_4O_2 + H_2O_2$. This property of absorbing oxygen in alkaline solution is not confined to hydroquinone alone but is shared in common by a number of substances, particularly those containing phenolic groups. It has been the subject of considerable biological interest as a possible mechanism for respiration in plants.

When the solution is acid, molecular oxygen is unable to effect any detectable oxidizing action upon hydroquinone, although the latter is rapidly oxidized by certain inorganic agents possessing less powerful oxidizing potentials than oxygen, such as ferric ions. This stability of hydroquinone to molecular oxygen in acid solution affords a very simple and accurate means of determining hydrogen-ion concentration in the presence of oxygen, which is impossible with the hydrogen electrode. The method consists simply in dissolving a few crystals of quinhydrone in the liquid and measuring the resulting oxidation potential with an inert electrode of gold or platinum. Under these conditions the following equilibrium[3,4] is set up: $C_6H_4(OH)_2 \rightleftharpoons C_6H_4O_2 + 2H^+ + 2\oplus$, to which the following equation applies: $\pi = \pi_0 + \frac{0.59}{2} \log \frac{[C_6H_4O_2]}{[C_6H_4(OH)_2]} + 0.059 \log H^+$.

The method depends essentially upon the fact that quinhydrone yields equimolecular proportions of quinone and hydroquinone in solution so that the term involving the concentration ratio vanishes and we obtain: $P_H = -\log H^+ = \pi_0 - \pi$ obs./0.059, an equation which differs from the usual hydrogen electrode formula only in the value of π_0. The values of

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[1] Manchot, "Habilitationschrift," Göttingen, **1899**.

[2] Euler and Bolin, *Z. Physiol. Chem.*, **57**, 81 (1908).

[3] Haber and Russ, *Z. physik. Chem.*, **47**, 294 (1904).

[4] Granger, *Dissertation*, Columbia University, March, **1920**. Granger and Nelson, THIS JOURNAL, **43**, 1401 (1921).

π_0 have been found[5,6] to be 0.6990 ± 0.0002 volt at 25° with a temperature coefficient of —0.00077 volt per degree, provided the activities of hydroquinone, quinone, and hydrogen ion are equal to unity.

Biilman has further shown that the method possesses the distinct advantage of being applicable to systems where the usual hydrogen electrode cannot be employed owing to the complications arising from the hydrogenation of unsaturated substances by platinum black. He has obtained excellent results with acrylic, crotonic, fumaric and maleic acids as well as with 0.1 *M* nitric acid. On the other hand, he and Granger and Nelson report that they were unable to get constant potentials in neutral and alkaline solutions. This would seem to be connected with the fact that hydroquinone is a weak dibasic acid, and since it is in the range where the ionization of hydroquinone becomes appreciable that auto-oxidation commences, it was considered of interest to investigate the functional relation existing between the hydrogen-ion concentration and the velocity of absorption of molecular oxygen. This appeared to be one of the most probable of the complications that might change the equimolecular ratio between quinone and hydroquinone and thereby invalidate the method.

A search of the literature showed that, although the velocity of oxygen absorption had been studied before, none of the experiments had been made on buffer solutions, the concentrations of alkali present being only in terms of total alkali without regard to the other substances present, so that the actual hydrogen-ion concentration may have varied within wide limits as well as changed markedly during the course of the reaction. The results of our experiments with solutions of known and definite hydrogen-ion concentration have shown that the rate of auto-oxidation is highly sensitive to changes in hydrogen-ion concentration and that the alkalinity at which the rate is sufficient to interfere with the use of the quinhydrone electrode is fairly well defined.

Experimental Part

The oxygen absorption was measured by rapidly shaking definite quantities of hydroquinone, dissolved in Palitsch's[7] borax plus boric acid buffer mixtures, in an atmosphere of oxygen or of air. The shaking apparatus was immersed in a water-bath at 25° and consisted of a metal swing carrying a 200cc., thick-walled glass cylinder, whose major axis extended along the line of motion. An electric motor drove the cylinder through a distance of approximately 10 cm. at a rate of 140 complete cycles per minute. This was sufficient to keep the solution in a state of froth, and create thereby a large gas-liquid interface.

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[5] Biilman, *Ann. chim.*, [9] **15**, 109 (1921).

[6] La Mer and Baker, THIS JOURNAL, **44**, 1954 (1922). Lillian Baker, *Dissertation*, Columbia University, where more details are given. Conant and Fieser, THIS JOURNAL, **44**, 2484 (1922). La Mer and Parsons, *J. Biol. Chem.*, **57**, 613 (1923).

[7] Mansfield Clark, "Determination of Hydrogen Ions," Williams and Wilkins Co., Baltimore, **1922**, p. 115.

The experiments were performed by introducing 90 cc. of the buffer solution, consisting of the proper quantity of 0.05 *M* borax solution, and the necessary quantity of 0.20 *M* boric acid solution minus 10 cc. of that necessary to form a mixture of the hydrogen-ion concentration under examination. Oxygen was then passed through to displace the air in the apparatus and the shaking continued for a few minutes to saturate the buffer solution with oxygen. The remaining 10 cc. of 0.20 *M* boric acid solution containing 0.108 g. of hydroquinone (0.001 mole) was then introduced through a side tube, and the shaking resumed at once. The absorption of oxygen commenced promptly and was read on an attached gas buret, at intervals appropriate to the velocity of the action. Inasmuch as the boric acid has a Sörensen value (*P*H) of about 5.34, no oxidation of the hydroquinone dissolved in it occurred until it was added to the borax portion of the buffer solution. The hydrogen-ion concentrations of the buffers were determined by performing a careful hydrogen electrode titration of the boric acid portion with the borax solution.

The results of these experiments at Sörensen values extending from 7.36 to 8.69 are shown on the curves in Fig. 1, the volume in cubic centimeters of oxygen absorbed being plotted along the *y* axis and the time in minutes along the *x* axis. The stoichiometrical volume of oxygen at 25° necessary to convert this quantity of hydroquinone into quinone is 24.4 cc., since half of the oxygen goes to form hydrogen peroxide, as was demonstrated by Manchot. From the curves it is evident (at least for solutions more alkaline than about *P*H 7.56) that secondary reactions are increasing the absorption of oxygen. Thus at *P*H 8.56 more than 24.4 cc. was absorbed in ten minutes, and on continuing the experiment it was found that a slow absorption continued for some time, totaling nearly twice this amount at the end of two days.

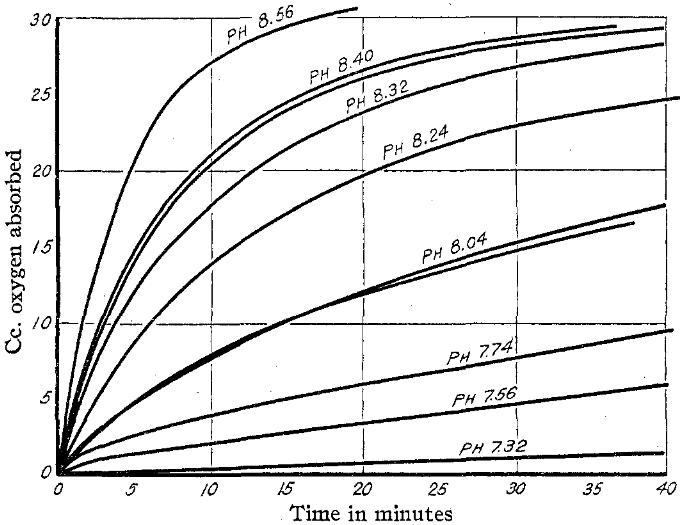

Fig. 1.—Hydroquinone concentration, 0.01 *M;* temp., 25°; gas oxygen total volume, 100 cc.

As a result of the interference of these secondary reactions, which may be either side reactions or of the consecutive type, it is impossible to calculate a velocity constant which will express the reaction velocity over any considerable period of time by the use of a theoretical equation.

In order to get a measure of the reaction velocity, recourse was taken to plotting the increment of oxygen absorbed per unit of time ($\Delta O_2/\Delta t$) against the time and then determining the initial rate by a graphical extrapolation. In so doing we believe that we obtain a better measure of the true initial rate, before it becomes complicated by side reactions, than if we attempt to express the rate as a monomolecular reaction as Euler and Bolin did. Except in some cases for the first reading, which is very likely in error on account of difficulties in determining the exact beginning of the reaction, the values of ($\Delta O_2/\Delta t$) when plotted against t fell very nearly upon a straight line until about 10 cc. was absorbed, when the rate commenced to decrease. The values of the initial rate as so determined are given in Table I, Col. 3.

TABLE I

EFFECT OF CHANGE OF *P*H ON THE INITIAL RATE OF OXYGEN ABSORPTION AT 25° BY HYDROQUINONE (0.01 *M*) DISSOLVED IN 100 CC. OF BORATE BUFFER SOLUTION

1 *P*H	2 $[H^+] \times 10^{10}$	3 $\left(\frac{\Delta O_2}{\Delta t}\right) \frac{\text{Cc.}}{\text{Min.}}$	4 $\left(\frac{\Delta O_2}{\Delta t}\right) \cdot [H^+]$ $\times 10^9$	5 $\left(\frac{\Delta O_2}{\Delta t}\right) \cdot [H^+]^2$ $\times 10^{18}$	6 $\left(\frac{\Delta O_2}{\Delta t}\right) \cdot \sqrt{[H^+]^3}$ $\times 10^{14}$
8.56	27.6	6.00	17.0	46	87
8.40	40.1	3.60 (?)	14.5	57	91
8.40	40.1	3.90	15.6	62	98
8.32	47.9	2.95	14.0	68	98
8.24	57.5	2.00	11.5	66	87
8.04	91.2	1.00	8.6	80	84
8.04	91.2	0.90	8.6	80	84
7.74	182.0	.34	6.2	112	84
7.56	276.0	.19	5.0	145	87
7.32	480.0	.04 (?)	1.9 (?)	91	42 (?)

If the rate of reaction depended upon the number of primary hydroquinone ions present, then the initial velocity multiplied by the hydrogen-ion concentration should yield a constant value. The figures in Col. 4 of the table show very definitely that this is not the case, but instead that the reaction velocity increases much faster than does the primary ionization with increase of *P*H. Likewise, if the velocity is determined by the concentration of completely dissociated (secondary) ions then the product of the initial rate and the square of the hydrogen-ion concentration should be constant. The values given in Col. 5 show a marked rise with increasing *P*H, indicating that in this case the velocity does not keep pace with the growing concentration of secondary ions. Instead we have found (Col. 6) that the product of the initial rate by $[H^+]^{3/2}$ yields a constant within the limits of experimental error for the apparatus, for seven of the eight different Sörensen values.

That this relationship involving the three-halves power of the hydrogen-ion concentration is sensibly correct over the range of Sörensen values which we have investigated can be inferred from the data of Euler and Bolin.[2] These authors studied the effect of minute additions of sodium hydroxide on the velocity of oxygen uptake of 0.10 *M* hydroquinone in the presence of 0.0005 to 0.02 equivalent of manganous acetate as catalyst, the manganous (Mn^{++}) ions causing oxidation at apparently much lower hydrogen-ion concentration than when they were not present. On the assumption that the addition of sodium hydroxide to make the solutions 0.0005, 0.001 or 0.002 *M* will cause a change in hydrogen-ion concentration in proportion to the quantity of alkali added we find from their figures for the apparent monomolecular constants that doubling the alkali added produces approximately a threefold (average 3.18) increase in the rate of reaction. When their data are recalculated as we have done for the initial rate of reaction, we obtain the figures given in Table II, where it will be seen that doubling the quantity of alkali leads to an increase of 2.82- while one would calculate an increase of 2.83-fold from the three-halves power relation.

TABLE II

CALCULATIONS FROM DATA OF EULER AND BOLIN[8] ON THE VELOCITY OF ABSORPTION OF OXYGEN BY HYDROQUINONE IN THE PRESENCE OF SODIUM HYDROXIDE AND MANGANOUS ACETATE

The velocity is calculated for the initial portion of the curve but over a sufficient period to yield significant figures as indicated in Col. 1. The concentration of hydroquinone was 0.2 *N*, and total volume 50 cc.

1	2	3	4	5	6	7
Time interval used in calculating Min.	Concn. of $Mn(Ac)_2$. *N*	O_2 Absorbed. Cc. per minute Concn. of NaOH present 0.002 *N* Cc.	0.001 *N* Cc.	0.0005 *N* Cc.	Ratio of Cols. $\frac{3}{4}$	$\frac{4}{5}$
3– 5	0.02	4.1	1.6	0.65	2.6	2.5
3– 8	.005	2.3	.57	.24	4.2	2.3
3–10	.002	1.1	.39	.16	2.7	2.5
3–10	.001	0.94	.30	.13	3.1	2.3
3–10	.0005	.81	.27	.09	3.0	3.0

Average of Cols. 6 and 7 = 2.82.

Col. 6 gives the ratio of the velocity for 0.002 *N* NaOH to that at 0.001 *N* NaOH, Col. 7 gives the same for 0.001 and 0.0005 *N* NaOH.

Euler and Bolin considered that the reaction involved the ionization of hydroquinone, but offer no explanation of their results except that they believed it likely that the manganous ion formed a complex with the hydroquinone. They were primarily interested in the manganese catalysis but also determined the primary ionization constant of hydroquinone to

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[8] Ref. 2, p. 87.

be $K_1 = 1.1 \times 10^{-10}$ by conductivity measurements.[9] Recently Sheppard[10] by means of a hydrogen-electrode titration has given 1.75×10^{-10} and 4×10^{-12} for the primary and secondary dissociation constants. These figures show that although the concentration of the secondary ion is much less than that of the primary ion, its concentration is nevertheless sufficient to be of possible kinetic importance as it varies between 10^{15} and 10^{17} molecules per liter. Taylor[11] has recently shown that when the concentration of the reacting component is of the order of magnitude 10^{16} a measureable reaction occurs. If we make the assumption that the rate of oxidation instead of depending on any one ionic form depends on the formation of a complex between the oxygen and the primary and secondary ions, a ready interpretation of the dependence of the rate of oxidation on the $^3/_2$ power of the hydrogen-ion concentration is obtained. The kinetic reaction would then be formulated: $2H_2C_6H_4O_2 \longrightarrow 3H^+ + HC_6H_4O^- + C_6H_4O^= \longrightarrow$ complex ion; complex $+ 2O_2 \longrightarrow 2$ Quinone $+ 2H_2O_2$. That the complex formation is more likely to occur in alkaline solution is supported by the fact that a reddish color develops on the addition of alkali to hydroquinone. The greater part of this color will disappear if the solution is promptly reacidified. The color is usually attributed to complexes of the quinhydrone type. Also Conant, Kahn, Fieser and Kurtz[12] state that they have observed that the oxidation-reduction curves of the homologous anthraquinone depart markedly from the theoretical curves in buffered alkaline solution although they were normal in acid solution. They explain this on the basis of complex formation.

Likewise from structural considerations the possibility of such complex formation must be envisaged. In the case of the acid fluorides the structural formula following Werner may be taken as $K^+(HF_2)^-$ where the two fluorine atoms are held together in a complex ion containing one atom of hydrogen. Similarly, we obtain for the complex of the two hydroquinone ions the structure: $3H^+[C_6H_4O_2HC_6H_4O_2]^{---}$. Lowry[13] has already pointed out that hydrogen can serve as a coördinating atom as in the formula $K^+[F^-H^+F^-]$ and the above complex appears to be an example of this type $3H^+[Q^{--}H^+Q^{--}]^{---}$. Although the electrical structure of these complex organic compounds is as yet uncertain the phenomenon of auto-oxidation as a coupled reaction permits of the development of certain considerations. It is usually considered that these processes can be formulated on the following scheme: $M + O_2 \longrightarrow MO_2$ (such as $Zn + O_2 = ZnO_2$); and $MO_2 + H_2O \longrightarrow MO + H_2O_2$ (such as $ZnO_2 + H_2O = ZnO +$

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[9] Ref. 2, p. 71.

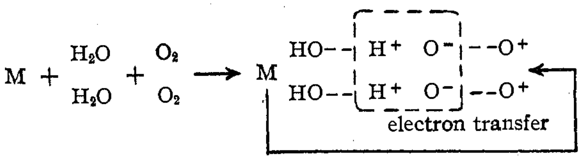
[10] Sheppard, *Trans. Am. Electrochem. Soc.*, preprint, 39, **1921.**

[11] Taylor, *J. Phys. Chem.*, **27,** 322 (1923).

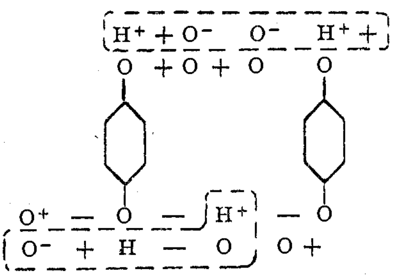
[12] Conant, Kahn, Fieser and Kurtz, THIS JOURNAL, **44,** 1392, 2481 (1922).

[13] Lowry and Burgess, *J. Chem. Soc.*, **123,** 2111 (1923).

H_2O_2). The discovery of Schönbein that substances could be divided into two types, namely, those possessing hydrogen peroxide and those generating ozone during their auto-oxidation process indicates that the above scheme is not sufficient. With van't Hoff one can postulate that atmospheric oxygen is always sufficiently ionized according to the scheme $O_2 \rightleftharpoons O^{++} + O^{--}$ to permit the possible reactions; $M + 2O_2 \longrightarrow M^{++} + O^{--} + O_3$; $M + O_2 + H_2O \longrightarrow M^{++} + O^{--} + H_2O_2$. The degree of ionic dissociation of oxygen is, however, too small under the conditions of auto-oxidation processes to justify this assumption. It appears more probable that the process of auto-oxidation necessitates the intramolecular ionization suggested by Sir J. J. Thomson, $O^{+}-O^{-}$, or polarization of the oxygen molecule, requiring an energy transfer far less than is necessary for complete ionization; thus, in the auto-oxidation of zinc with simultaneous formation of hydrogen peroxide we may depict the course of events as follows,



or the oxidation of phosphorus $P_2 + 2\,[O^{+} - O^{-}] \longrightarrow P_2O + O_3$. The mechanism of hydrogen acceptation or oxygen donation in these and similar biological processes thus ultimately rests on the formation of a complex and an electrical circuit as suggested by Armstrong. The complex undergoes decomposition and rearrangement due to polarization or intramolecular ionization of the oxygen molecule effected by the reactants. In a similar manner, one may formulate the case under consideration as follows,



An interesting corollary to the hypothesis that auto-oxidation is primarily due to an intramolecular ionization of the oxygen molecule produced by the inductor, such as by the absorption of hydroxyl ions of water as in the case of zinc or reaction with an oxygen molecule as in the case of phosphorus or ferric oxide, is that the complex MO_2, for example, $Fe_2O_3 . O_2$ should possess a higher oxidizing potential than molecular oxygen since, although the addition of oxygen to ferric oxide to form the peroxide is accompanied by a decrease in free energy, the intramolecular ionization of the oxygen molecule resulting from such addition raises the oxidizing power. Experimental evidence in favor of this hypothesis will be the subject of a future communication.

In order to test the effect of increasing the concentration of hydroquinone over a wide range of concentrations, it was necessary to work at *P*H 7.56 as the reaction proceeded at such a rapid rate at higher concentrations of hydroquinone that it would have been impossible for the diffusion of oxygen into the liquid to keep pace with the rate at which it was being consumed by the hydroquinone if a more alkaline buffer were used. The two series of experiments given in Fig. 2 and Table III show that the rate increases linearly with the concentration of hydroquinone for the less concentrated hydroquinone solutions but falls off from this relation for the higher concentrations. This decrease from the expected linear relationship between concentration and velocity is due no doubt to the fact that hydroquinone

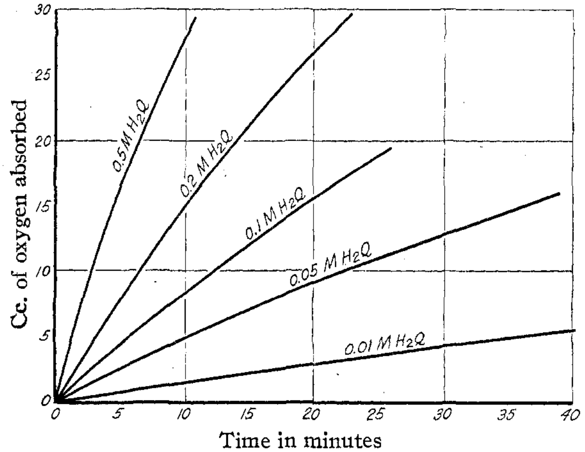


Fig. 2.—The effect of initial concentration of hydroquinone (H_2Q) in borate buffer of *P*H 7.56

TABLE III

THE EFFECT OF VARYING THE CONCENTRATION OF HYDROQUINONE ON THE VELOCITY OF OXYGEN ABSORPTION IN 0.20 *M* BORATE BUFFER SOLUTION

	Series A		Series B	
Concn. of $C_6H_4(OH_2)$ added *M*	O_2 absorbed by 100cc./minute Cc.	O_2 absorbed per minute per mole of hydroquinone Cc.	O_2 absorbed by 100cc./minute Cc.	O_2 absorbed per minute per mole of hydroquinone Cc.
0.01	0.17	170	0.21	210
.025	.42	164	.53	210
.05	.52	110	.87	170
.10	1.97	97	1.1	110
.20	1.6	80	1.8	90

In Series A the buffer solution (0.20 *M*) was exactly *P*H 7.56 previous to the addition of hydroquinone. In series B it was approximately 7.6.

is an acid of almost the same strength as boric acid ($K = 6.6 \times 10^{-10}$) so that the addition of large quantities of hydroquinone reduced the Sörensen value enough to affect the highly sensitive auto-oxidation reaction. The time at our disposal did not permit of us to repeat the experiments with more concentrated and more stable buffers. The concordance in the dilute solutions in lieu of better data may be taken as indicating that the reaction is linear in respect to change in concentration of hydroquinone as the equation demands.

We intend to investigate more fully at a later date this matter as well as the question of the relationship existing between the rate of auto-oxidation and the effect of substitution upon the oxidation-reduction potentials and the ionization constants of diphenolic bodies.

A few experiments were made to determine the rate when air was used in place of oxygen. Under these conditions the rate was found to be slower. These experiments have shown that a rate comparable to that given for oxygen at P_H 7.3 is not reached with air until about P_H 7.8. This corresponds to a rate of change of approximately 2×10^{-5} moles of hydroquinone per minute, or a change of electromotive force of 0.02 millivolt per minute when the concentration of quinone and hydroquinone are both 0.01 *M*. From experiments using the quinhydrone electrode in the presence of air, but without shaking, La Mer and Parsons[14] have found that this rate of electrode potential change occurred at P_H 8.1. They were able to utilize the quinhydrone electrode under such conditions for the determination of the hydrogen-ion concentration of buffer solutions up to an alkalinity of P_H 8.0 with an error of less than 0.02 Sörensen unit.

Summary

The rate at which hydroquinone reacts with molecular oxygen has been followed by shaking borate buffers 0.01 *M* in respect to hydroquinone with oxygen. The rate does not become appreciable until the solution is more alkaline than P_H 7.3 or 7.8 for air. The rate of reaction is highly sensitive to further increase of Sörensen value, being proportional to the three-halves power of the hydrogen-ion concentration for the region investigated. A mechanism involving complex formation of the primary and secondary ions is suggested to explain this relationship.

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[14] La Mer and Parsons, *Proc. Soc. Exptl. Biol. Med.*, Feb., **1923**; *J. Biol. Chem.*, **57**, 613 (1923).