2. The dissolution rate parameter [k_1, k_2, or k_3 (Eqs. 1–3)] is constant and the same for all particles. According to the theory on which the models are based, this condition can only be achieved experimentally if the temperature and composition of the dissolution liquid are maintained constant and the flow rate is constant or uniform in the cross section of the dissolution cell where the particles are placed.

The first two conditions are easily met. With regard to the flow rate, the dissolution cell used has a very useful feature: the process can be stopped and the cell rapidly disconnected, allowing the particles to be inspected at any stage of the dissolution process. Such inspections showed (after microscopic measurements) uniform dissolution over the whole particle layer, indicating a uniform flow rate. The fact that the particles in the dissolution cell can be inspected in this way makes it possible for dissolution data to be combined with particle-size measurements (2).

3. The initial particle-size (diameter) distribution can be approximated by a truncated log-normal distribution function. Figure 7 shows that this is the case for the 500 particles *measured*. However, it does not guarantee the correctness of the assumption that this small sample represents the particle-size distribution in the samples used for the dissolution tests, although optical investigations of the uniformity of the powder support this assumption.

This paper has shown that it is possible to describe mathematically the dissolution of a multiparticulate system with a high degree of accuracy by considering both the particle-size distribution effect and the particle shape effect discussed earlier (2, 3). It is evident from the dissolution data obtained that, of the three models investigated, the single-particle dissolution model (Eq. 1) describes the kinetics best.

More complex and flexible models for single-particle dissolution possibly could describe the dissolution more adequately. The fact that the *ss* values for the cube root and the square root model are almost the same suggests a model with properties between these two. The Danckwerts model, as discussed by Goyan (7), is given by:

$$-dw/dt = A[(Dp)^{1/2} + D/a]C_a \qquad \text{(Eq. 15)}$$

where w = weight undissolved, A = surface area, D = diffusion coefficient, p = quantity related to stirring, a = radius of the particle, and C_a = steady-state concentration.

This model is very flexible. When $(Dp)^{1/2}$ predominates, the apparent model would be the cube root model. As the quantity D/a becomes more important, the square root model becomes the apparent model. Finally, as the quantity D/a predominates, the squared cube root model becomes the apparent model. However, when applied to the log-normal case, the Danckwerts model results in a mathematical expression that is much more complex than Eqs. 6*a*–6*c*.

The fit of the dissolution data to the cube root model is excellent. Therefore, if an application of the Danckwerts model results in an even better fit, this improvement will most likely be statistically insignificant considering the magnitude of the experimental errors. The Hixon–Crowell model should be the preferred model in such a case because of its simplicity.

REFERENCES

(1) P. Veng Pedersen and K. F. Brown, *J. Pharm. Sci.*, **64,** 1192(1975).

(2) *Ibid.*, **64,** 1981(1975).

(3) *Ibid.*, **65,** 1437(1976).

(4) A. W. Hixson and J. H. Crowell, *Ind. Eng. Chem.*, **23,** 923(1931).

(5) P. J. Niebergall, G. Milosovich, and J. E. Goyan, *J. Pharm. Sci.*, **52,** 236(1963).

(6) W. I. Higuchi and E. N. Hiestand, *ibid.*, **52,** 67(1963).

(7) J. E. Goyan, *ibid.*, **54,** 645(1965).

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Antioxidant Efficiency of Polyhydric Phenols in Photooxidation of Benzaldehyde

DOUGLAS E. MOORE

Abstract □ An experimental system is described for the observation of the kinetics of photochemically initiated oxidation reactions; this system is based on the measurement of oxygen consumption with a polarographic oxygen electrode. The photooxidation of benzaldehyde in dilute aqueous solution was examined and appears to conform to a free radical chain mechanism. The antioxidant efficiency of some polyhydric phenols was determined kinetically and found to be catechol > pyrogallol > hydroquinone > resorcinol > *n*-propyl gallate for the benzaldehyde photooxidation.

Keyphrases □ Antioxidants—efficiency, various polyhydric phenols, photooxidation of benzaldehyde □ Phenols, polyhydric—antioxidant efficiency in photooxidation of benzaldehyde □ Photooxidation—benzaldehyde, antioxidant efficiency of various polyhydric phenols □ Benzaldehyde—photooxidation, antioxidant efficiency of various polyhydric phenols □ Oxidation, photochemical—benzaldehyde, antioxidant efficiency of various polyhydric phenols

The oxidative deterioration of pharmaceuticals can be initiated by UV light, both in the presence and absence of sensitizers. Two general classes of photooxidation reactions are recognized. The first is the free radical chain mechanism of autoxidation (1) initiated by sensitizers (such as benzophenone) that abstract a hydrogen atom from the oxidant. The free radical thus formed adds molecular oxygen and propagates the chain mechanism by abstracting hydrogen from a further oxidant molecule, giving rise to a hydroperoxide. The hydroperoxide is the major product but usually reacts further by a slower nonradical disproportionation mechanism. For example, it was shown (2) that the peroxy acid formed during aldehydic autoxidation undergoes an acid–base-catalyzed decomposition reaction to the carboxylic acid, which is the final reaction product.

The second class is the dye-sensitized (*e.g.*, methylene blue) photooxygenation of acceptors (3). These acceptors are either: (*a*) cyclic dienes, polycyclic aromatic compounds, or heterocyclic compounds producing cyclic

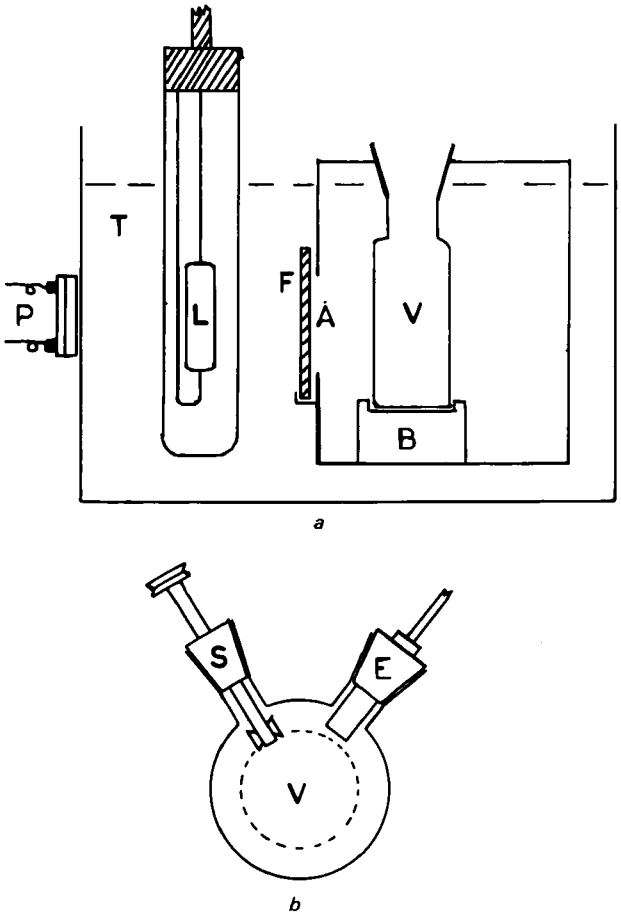


Figure 1—(a) *Schematic view of apparatus for observing the kinetics of photooxidation.* (b) *Reaction vessel, side elevation.* (*For explanation of symbols, see text.*)

peroxides, or (*b*) olefins containing allylic hydrogen atoms whereby allylic hydroperoxides are the products. As an example of the difference between the two types of oxidation, isopropyl alcohol is a very reactive acceptor for radical oxidation but is inert in the dye-sensitized photooxygenations (4). Furthermore, some acceptors can act as their own sensitizers for both classes of reaction.

An understanding of the kinetics and mechanism by which a compound is oxidized clearly enables the more judicious selection of an inhibitor or antioxidant. The traditional means of observing the kinetics of an oxidation reaction has been manometry (5), in which the rate of oxygen absorption from the gas phase into a vigorously stirred solution is measured. Only in specific cases where suitable differences in the absorption spectrum of the reactant and product exist, *e.g.*, ascorbic acid (6), can a spectrophotometric procedure be used. The major disadvantage of manometry is the requirement of a reaction slow enough that the diffusion of oxygen into the reacting solution is not rate determining. Particularly for photooxidation reactions, it is imperative that the sample be irradiated uniformly; thus the vigorous mixing of the gas phase with the solution, creating a vortex and bubbles of variable intensity, is undesirable.

One technique devised for measuring oxygen consumption in the solution directly involves following the decay of the charge transfer spectrum produced by a solution of oxygen in an organic solvent (7). More widely applicable, however, are the variants of the polarographic technique such as the oxygen electrode (8) and the galvanic cell oxygen analyzer (9). A manometric system was compared to a galvanic cell system, and the latter was more sensitive and responded faster in the measurement of oxygen-consuming enzyme systems (10).

In this paper, an experimental system is described for studying the kinetics of photooxidation reactions. The system is based on the rate of oxygen consumption measured with a polarographic oxygen electrode. As a model for the evaluation of the method, the photooxidation of benzaldehyde and its inhibition in aqueous solution were studied. Previously, benzaldehyde oxidation initiated by metal ions was used to study the role of surfactants in retarding oxidation (11–14).

The photooxidation of benzaldehyde was studied at concentrations of the order of 20% in *n*-decane solutions; a free radical mechanism was found to apply (15). The purposes of the present investigation were to determine whether the same mechanism applies in aqueous solution where the benzaldehyde solubility limit is 0.65% (16) and to determine the antioxidant efficiency of several polyhydric phenols.

EXPERIMENTAL

Materials—The oxidation of aldehydes has been stated to be very sensitive to traces of impurities (17); therefore, all chemicals were carefully purified and the reaction vessels were exceedingly clean.

Benzaldehyde[1] was distilled at a pressure of 10 mm of nitrogen, and the fraction with a boiling point of 61–63° was collected. This fraction was then twice distilled *in vacuo* using an all-glass apparatus; the middle 80% was collected each time. The final product was stored under vacuum and protected from light. Whenever a new solution of benzaldehyde was prepared, the required amount of benzaldehyde was vacuum distilled from the reservoir.

Water, first deionized and then double distilled, was used in the preparation of all solutions. Hydroquinone, resorcinol, catechol, and pyrogallol[1] were recrystallized from distilled water under an atmosphere of nitrogen. Solutions (10^{-4}–10^{-3} *M*) of these antioxidants were freshly prepared in air-saturated water when required.

Glassware was cleaned by rinsing with permanganic acid, water, and acidified hydrogen peroxide, followed by numerous rinsings with double-distilled water.

Apparatus—The reaction vessel and the arrangement of the apparatus are shown in Fig. 1. The reaction vessel (V) was constructed[2] from borosilicate glass of uniform thickness and had parallel circular faces 2.5 cm apart and 6 cm in diameter. Two standard taper sockets carried a belt-driven turbine stirrer[3] (S) and the oxygen electrode[4] (E). The UV light source (L) was a medium pressure mercury arc[5] with maximum output at 365 nm.

A glass filter[6] (F), positioned in front of the aperture (A), absorbed wavelengths below 310 nm and above 420 nm, with 28% transmission at 365 nm. The dotted circle in Fig. 1*b* indicates the approximate size of the aperture in relation to the reaction vessel. The volume of reaction mixture contained when the stirrer and oxygen electrode were in place was 135 ml.

The reaction vessel, filter, and aperture were mounted on a block (B), which could be fixed at different distances from the UV lamp to vary the light intensity incident on the reaction mixture. The whole

[1] British Drug Houses Analar grade.
[2] In the workshop at the University of Sydney.
[3] Machined from Teflon (du Pont).
[4] Radiometer (Copenhagen) E5046 oxygen electrode controlled with Radiometer PHM927 gas monitor.
[5] Engelhard Hanovia, United Kingdom.
[6] Corning filter CS 7-37.

Table I—Comparison of Rates of Photooxidation Measured by Manometry and Oxygen Electrode at 30°

Benzaldehyde, M	Reaction Rate, moles of Oxygen/liter/min × 10^6	
	Manometry[a]	Oxygen Electrode[b]
0.0540	—	15.70 ± 0.63
0.0422	—	12.52 ± 0.42
0.0309	9.07 ± 0.80	9.31 ± 0.36
0.0265	7.80 ± 0.54	8.00 ± 0.32
0.0207	6.10 ± 0.45	6.25 ± 0.20
0.0156	4.68 ± 0.47	4.75 ± 0.20
0.0112	3.36 ± 0.36	3.40 ± 0.15

[a] Means of three determinations. [b] Means of five determinations.

apparatus was immersed in a water bath thermostated to ±0.05°. Positioned on the outside face of the thermostat directly opposite the mercury arc was a photovoltaic cell[7] (P); this cell was used to monitor the lamp output.

The signals from the oxygen electrode (*via the gas monitor*) and the photocell were fed into a two-pen recorder[8].

Procedure—Benzaldehyde solutions were prepared by dissolving the required quantity of freshly vacuum-distilled benzaldehyde in air-saturated water. The benzaldehyde was weighed when cold (~0°) and transferred quickly into a volumetric flask, with copious washings of air-saturated water. Solutions were stored in the dark, with intermittent shaking, to ensure that dissolution was complete (3 hr). If this procedure was not followed, rates of photooxidation were anomolously high, corresponding to the accelerated rate of oxidation in emulsions compared to solutions observed by Carless and Nixon (11).

The reaction vessel was filled with the benzaldehyde solution, preequilibrated to the desired temperature, and the stirrer and oxygen electrode were set in place, excluding air bubbles. The oxygen electrode was calibrated each day at the appropriate temperature with air-saturated water and oxygen-free water prepared by adding sodium sulfite to a borax solution. A shutter was positioned between the lamp and the aperture, and the lamp was switched on.

In each experiment, there was no observable change in P_{O_2} (oxygen partial pressure) prior to shutter removal, indicating insignificant thermal oxidation and negligible stray light. When the lamp reached a steady intensity (7 min), the shutter was removed and the P_{O_2} was recorded. In the absence of antioxidants, there was no indication of an induction period and the change in P_{O_2} stopped immediately when the UV lamp was switched off.

A linear trace of P_{O_2} was recorded down to 50 mm Hg. For most purposes, the slope of this trace in millimeters of mercury per minute was used as a measure of the rate of oxidation. When required, the absolute rate of reaction, R, in moles of oxygen per liter per minute was obtained from[9]:

$$R = \frac{(\text{slope in mm Hg/min}) \times (\text{solubility of oxygen in water})}{P_{O_2}} \qquad \text{(Eq. 1)}$$

For the investigation of the effects of antioxidants, the desired amount of antioxidant solution (1–5 ml) was mixed with 200 ml of benzaldehyde solution immediately prior to filling the reaction vessel.

RESULTS AND DISCUSSION

Kinetic Parameters—The linearity of P_{O_2} *versus* time indicates that photooxidation of benzaldehyde is independent of oxygen concentration (for $P_{O_2} > 50$ mm). Variation of the initial benzaldehyde concentration to the near-saturated solution level showed a first-order dependence with a linear correlation coefficient of 0.9998 (Table I). As part of the preliminary testing of the technique, some photooxidation experiments were performed with a manostat similar to that described by Carlsson and Robb (18); it was connected to the reaction vessel in place of the oxygen electrode. The measured rates (Table I) compared favorably except at high initial benzaldehyde concentration (*i.e.*, high rate), emphasizing the inadequacy of manometry for following rapid oxygen absorption.

There is, of course, an upper limit to rates that can be followed by the oxygen electrode, depending as it does on the diffusion of oxygen across the 20-μm polypropylene membrane and efficient stirring of the system. Since thermal oxidation was insignificant in relation to photooxidation in the temperature range studied (15–40°), the upper limit to rate measurements was arbitrarily defined as the point at which the reading monitored from the oxygen electrode continued to decrease after the light was switched off. The upper limit for the apparatus in this study was thereby set at approximately 20 mm Hg/min or 4×10^{-5} mole/liter/min.

The catalytic interference by metal ions was declared unimportant when the presence of edetic acid (10^{-4} M) in the reaction mixture had no effect on the rate of oxidation.

Variation of light intensity gave the results in Fig. 2. The log rate *versus* log intensity plot with a slope of 0.51 ± 0.03 established that the rate of reaction depended on the square root of the incident light intensity, as determined by the ferrioxalate chemical actinometer system (19).

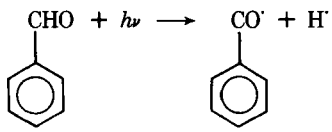
From the experiments in which the temperature was varied in the range of 15–40°, the overall activation energy of the reaction was 28 ± 1 kJ/mole as determined from the Arrhenius plot in Fig. 3.

Iodometric analysis of the reaction mixture immediately after irradiation was stopped for four benzaldehyde concentrations (Table I) showed a 1:1 correlation between moles of peroxide product and moles of oxygen consumed.

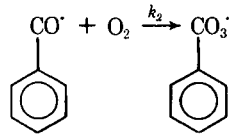
The unretarded rate data can be summarized by:

$$R_u = k[\text{benzaldehyde}](I^{1/2}) \qquad \text{(Eq. 2)}$$

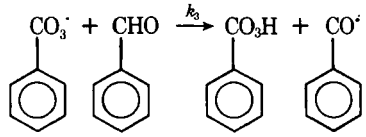
which is the same form as found by Ingles and Melville (15) in decane solution. Thus, one can assume the same chain mechanism (Schemes I–IV).



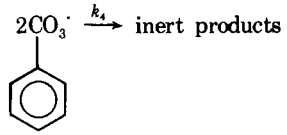
Scheme I—Initiation



Scheme II—Propagation



Scheme III—Propagation



Scheme IV—Termination

Assumptions made in applying the steady-state treatment to this reaction sequence are: (*a*) the reaction chains are long, (*b*) k_3 is the rate-determining step at the oxygen pressure used, and (*c*) termination steps involving $C_6H_5CO\cdot$ radicals are unimportant. The overall rate equation is, therefore, given in Eq. 3, which is of the same form as Eq. 2:

$$R_u = k_3(R_i/k_4)^{1/2}\,[\text{benzaldehyde}] \qquad \text{(Eq. 3)}$$

[7] Evans Electroselenium Ltd., United Kingdom.
[8] Rikadenki B-241, Tokyo, Japan.
[9] The solubility of oxygen in water was obtained from Ref. 5.

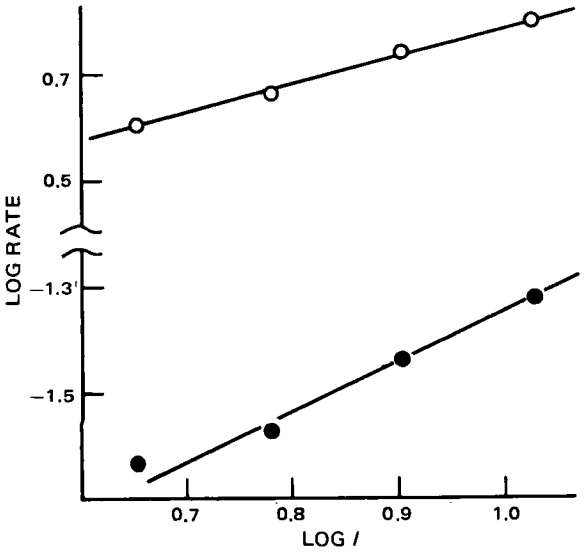


Figure 2—*Benzaldehyde photooxidation in water solution: effect of variation of light intensity on the rate (in mm Hg/min) at 30°. Each data point is the mean of four determinations. Key: ○, 0.0290 M benzaldehyde; and ●, 0.0315 M benzaldehyde + 1 × 10^{-6} M catechol.*

where R_i, the rate of the initiation step, would be expected to be directly proportional to the intensity of incident UV light, a factor considered independent of temperature.

Therefore, the overall energy of activation would be essentially that of k_3, the propagation step, since radical combination steps, such as k_4, have near zero activation energies (20).

Effects of Antioxidants—The polyhydric phenols, hydroquinone, catechol, resorcinol, pyrogallol, and *n*-propyl gallate, slowed the rate of oxidation of benzaldehyde, although none inhibited the reaction completely. These compounds were classed as retarders rather than inhibitors of this system. According to Scott (1), phenol and amine antioxidants act in radical autoxidation by terminating the chain

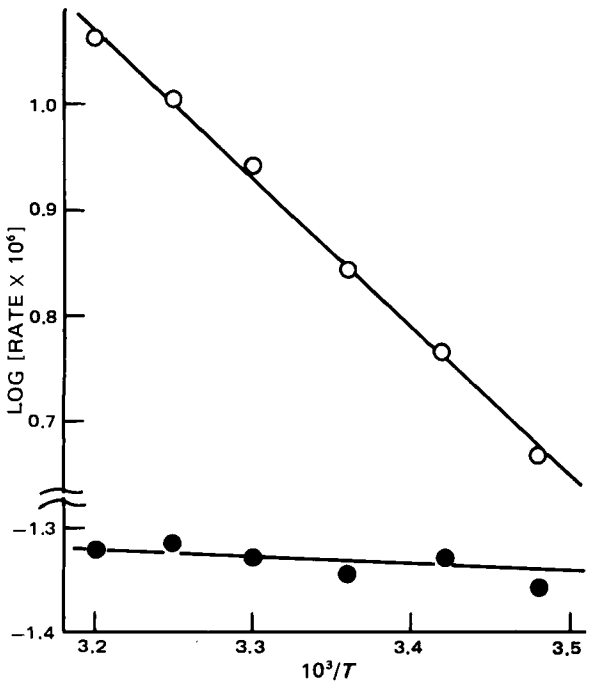


Figure 3—*Benzaldehyde photooxidation in water solution: effect of temperature variation on the rate (in moles/liter/min). Each data point is the mean of four determinations. Key: ○, 0.029 M benzaldehyde; and ●, 0.0315 M benzaldehyde + 1 × 10^{-6} M catechol.*

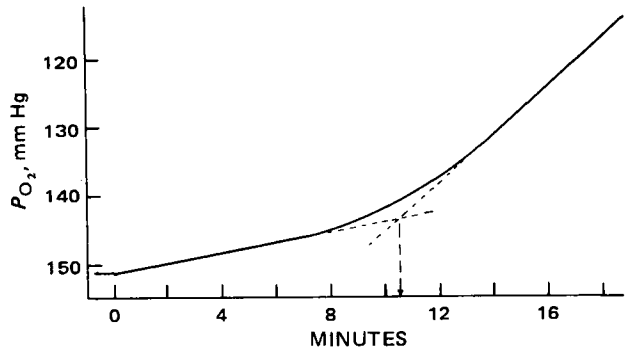
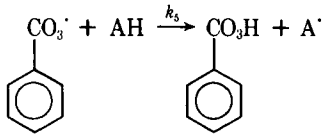


Figure 4—*Recorder trace of oxygen partial pressure during catechol-retarded photooxidation of benzaldehyde at 30° with 0.0315 M benzaldehyde + 1 × 10^{-6} M catechol. The intersection of the linear portions gives the length of the induction period as indicated by the arrow.*

carrying radical (Scheme V).



Scheme V

The antioxidant radical A· is presumed to be relatively stable and capable of reacting only with species of the same kind. In the early stages of the reaction, Scheme IV may be assumed to be negligible compared to Scheme V, so kinetic analysis gives for the rate of the retarded reaction:

$$R_a = R_i \left(1 + \frac{k_3}{k_5}\frac{[\text{benzaldehyde}]}{[\text{AH}]}\right) \qquad \text{(Eq. 4)}$$

A plot of the rate of the retarded reaction *versus* the reciprocal of retarder concentration at constant benzaldehyde concentration should be linear. As required for Eq. 4, the rates recorded here are presteady state; *i.e.*, if the reaction is allowed to proceed for a long enough time, the rate increases to the value obtained in the absence of retarder (Fig. 4). An induction period can be defined as the time at which the changeover to the unretarded rate occurs (Fig. 4). The induction period was found to be approximately proportional to retarder concentration but not sufficiently accurate to be used as a measure of antioxidant efficiency.

The presteady-state rate of the retarded reaction also was directly dependent on the incident UV light intensity but essentially independent of temperature, as shown in Figs. 2 and 3, respectively, for catechol at 1 × 10^{-6} *M*. This behavior conforms to that predicted by Eq. 4 in that R_i depends directly on the lamp intensity, and the rate constants k_3 and k_5 are both for reactions of the free radical with acceptor molecules. Thus, their activation energies are probably similar, producing an overall energy of activation near zero.

Figure 5 displays the results plotted in the form of Eq. 4. Catechol, hydroquinone, and resorcinol all appeared to act in the same way, as evidenced by the linear plots with a common intercept. The value for R_i, the rate of initiation, was 1.8 × 10^{-7} mole/liter/min.

The results for pyrogallol deviated from linearity above 2 × 10^{-6} *M*. Control experiments, at a concentration of 10^{-3} *M*, indicated that pyrogallol is itself slowly photooxidized. Whether pyrogallol oxidation contributes to the apparent oxygen uptake when benzaldehyde is present is impossible to tell without product analysis, a difficult task since the extent of reaction is less than 1%. Similar controls for hydroquinone, catechol, and resorcinol up to 5 × 10^{-3} *M* showed no oxygen uptake. Other commonly employed antioxidants, such as butylated hydroxytoluene and *n*-propyl gallate, were much less active than resorcinol. Full data over a range of concentrations for butylated hydroxytoluene could not be obtained due to its limited water solubility.

A value for the rate constant ratio k_5/k_3 can be obtained from the slopes of the lines in Fig. 5 and used as a measure of antioxidant efficiency. Bolland and tenHave (21) developed the following method for estimating antioxidant efficiencies; rate data at only one antioxidant concentration are required. At a low antioxidant concentration,

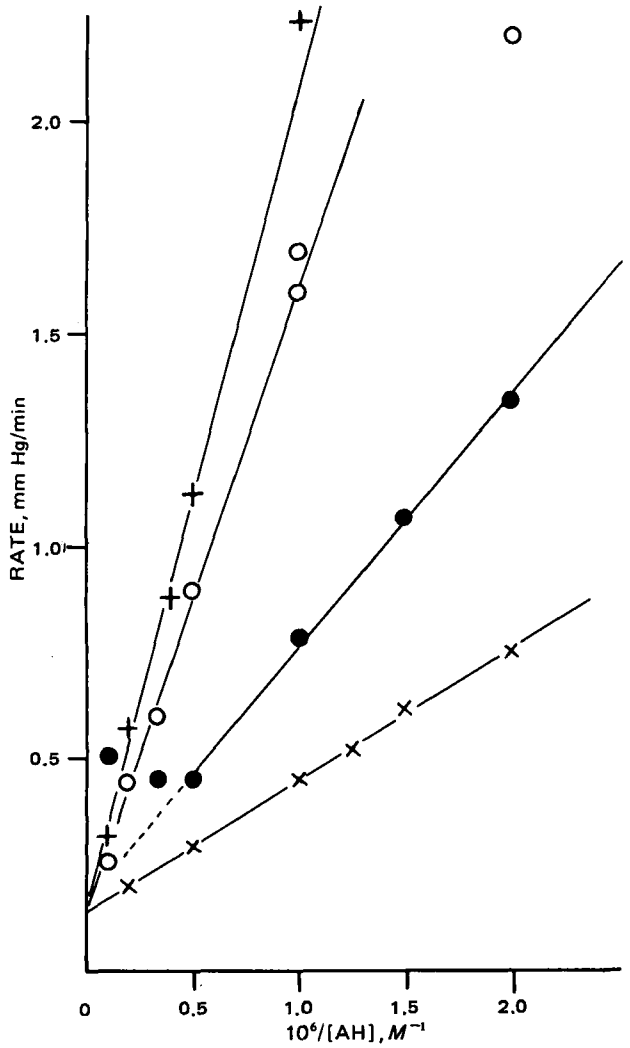


Figure 5—*Rate of the retarded reaction as a function of the reciprocal of the retarder concentration at 30°. Each data point is the mean of three determinations. Key: ×, catechol; ●, pyrogallol; ○, hydroquinone; and +, resorcinol. The benzaldehyde concentration was 0.0315* M *in all experiments.*

Eq. 4 approximates to:

$$R_a \approx R_i \frac{k_3}{k_5} \frac{[\text{benzaldehyde}]}{[\text{AH}]} \qquad \text{(Eq. 5)}$$

Combination of Eqs. 3 and 5 leads to:

$$R_a = \frac{R_u^2 k_4}{k_3 k_5 [\text{benzaldehyde}][\text{AH}]} \qquad \text{(Eq. 6a)}$$

or:

$$\frac{k_4}{k_3 k_5} = \frac{R_a}{R_u^2} [\text{benzaldehyde}][\text{AH}] = K \qquad \text{(Eq. 6b)}$$

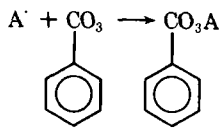
Bolland and tenHave (21) calculated $1/K$ values for polyhydric phenol antioxidants in the benzoyl peroxide-initiated oxidation of ethyl linoleate. In Table II, their values are compared with values of $1/K$ calculated for the same antioxidants in the photooxidation of benzaldehyde. Clearly, either $1/K$ or k_5/k_3 produces the same antioxidant efficiency ranking, but the latter ratio is preferable because it is calculated over the full range of antioxidant concentrations.

Bolland and tenHave (21) also showed that a log-linear relationship existed between their measured antioxidant efficiency and the normal reduction potential of these phenolic antioxidants measured in a neutral aqueous ethanol solution (22). However, we found catechol to be a more effective antioxidant than hydroquinone and pyrogallol. Whether this result was due to a different reactivity of the semiquinone radical (A·) produced in Scheme V could not be determined at present. It is not possible from kinetic studies alone to decide the fate of the free radical A· since Schemes V and VI both give kinetics of the form of Eq. 4.

Table II—Antioxidant Efficiencies of Polyhydric Phenols in the Photooxidation of Benzaldehyde

Antioxidant	$10^2 \times k_5/k_3$	$10^2/K$, liters/mole/min[a]	$1/K$[b]
Catechol	7.5 ± 0.2	28.0 ± 1.2	0.63
Pyrogallol	4.0 ± 0.15[c]	16.0 ± 1.3	3.0
Hydroquinone	1.9 ± 0.2	7.7 ± 1.0	1.0
Resorcinol	0.85 ± 0.05	2.0 ± 0.2	0.016
n-Propyl gallate	0.48 ± 0.06	1.1 ± 0.2	—

[a] Average value from four lowest antioxidant concentrations. [b] From Ref. 21. Values referred to hydroquinone as unity. [c] From linear portion of graph, pyrogallol $< 2 \times 10^{-6}$ *M*.



Scheme VI

REFERENCES

(1) G. Scott, "Atmospheric Oxidation and Antioxidants," Elsevier Publishing Co., Amsterdam, The Netherlands, 1965, chaps. 3, 4.

(2) W. A. Waters, *J. Chem. Soc.*, **1951**, 812.

(3) C. S. Foote, *Acc. Chem. Res.*, **1**, 104(1968).

(4) G. O. Schenck, H.-D. Becker, K.-H. Schulte-Elte, and C. H. Krauch, *Ber.*, **96**, 509(1963).

(5) W. W. Umbriet, R. H. Burris, and J. F. Stauffer, "Manometric Techniques," 4th ed., Burgess, Minneapolis, Minn., 1964.

(6) S. M. Blaug and B. Hajratwala, *J. Pharm. Sci.*, **61**, 556(1972).

(7) J. Betts and J. C. Robb, *Nature*, **215**, 274(1967).

(8) L. C. Clark, *Trans. Am. Soc. Artif. Intern. Organs*, **2**, 41(1956).

(9) K. H. Mancy, D. A. Okun, and C. N. Reilley, *J. Electroanal. Chem.*, **4**, 65(1962).

(10) H. Lipner, L. R. Witherspoon, and A. Wahlborg, *Anal. Chem.*, **37**, 347(1965).

(11) J. E. Carless and J. R. Nixon, *J. Pharm. Pharmacol.*, **9**, 963(1957).

(12) J. E. Carless and J. Swarbrick, *ibid.*, **14**, 97T(1962).

(13) J. Swarbrick and J. E. Carless, *ibid.*, **16**, 596(1964).

(14) *Ibid.*, **16**, 670(1964).

(15) T. A. Ingles and H. W. Melville, *Proc. Roy. Soc. (London)*, **A218**, 175(1953).

(16) A. G. Mitchell, L. S. C. Wan, and S. G. Bjaastad, *J. Pharm. Pharmacol.*, **16**, 632(1964).

(17) M. Niclause, J. Lemaire, and M. Letort, "Advances in Photochemistry," vol. 4, Interscience, New York, N.Y., 1966, p. 27.

(18) D. J. Carlsson and J. C. Robb, *Trans. Faraday Soc.*, **62**, 3404(1966).

(19) J. G. Calvert and J. N. Pitts, Jr., "Photochemistry," Wiley, New York, N.Y., 1966, p. 783.

(20) F. S. Dainton, "Chain Reactions," 2nd ed., Methuen, London, England, 1966, chap. 1.

(21) J. L. Bolland and P. tenHave, *Discuss. Faraday Soc.*, **2**, 252(1947).

(22) L. F. Fieser, *J. Am. Chem. Soc.*, **52**, 5204(1930).

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