REPORT NO. 58

RATES OF CATALYTIC DECOMPOSITION OF LIQUID HYDROGEN PEROXIDE ON METAL SURFACES

Prepared for the
Office of Naval Research
Contract Nonr 1841(11)
NR-092-008

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The relative activity of 14 heavy metal catalysts in decomposition of liquid H_2O_2 was determined under conditions which permitted recognition and control of oxygen supersaturation phenomena and mass transfer effects. Platinum and silver were studied extensively. At 30°C, the decomposition rate on all metals except silver was found to be a surface-controlled process. The intrinsic activity of silver is at least several times greater than that of any other metal studied and its rate was completely mass-transfer controlled at all H_2O_2 concentrations investigated. The turbulence produced by oxygen bubble evolution as observed on silver surfaces can cause the mass transfer rate to be increased by a factor of as much as 10 over that encountered without forced turbulence or gas evolution.

Most studies of hydrogen peroxide decomposition during past decades have concentrated on problems of mechanism. Very little attention has been given to determining rate data in such a manner that the effectiveness of various heterogeneous catalysts could be compared, although this aspect of hydrogen peroxide decomposition is of much concern in engineering applications. Rate data by different authors, where they have been reported, have often been irreproducible and frequently conflicting, reflecting in part the typical complex nature of a heterogeneous catalytic reaction. However in addition variables such as catalyst surface area, mass transfer effects, supersaturation phenomena and the effects of additives and impurities in H_2O_2 on the reaction have frequently not been controlled, or sometimes even recognized.

The object of this study was (1) tc devise means of comparison of the activity of heavy metal catalysts and to make such comparisons, taking into account the effect of the pertinent variables mentioned above and (2) to determine the conditions under which rate processses other than the true surface reaction as such could be controlling. It was hoped that the catalytic activity of a metal could be correlated with its electronic parameters, e.g., electronic work function, coefficients of electronic specific heat, or the d-band character of the metals. Such correlations have been found by Dowden

and Reynolds (3), Schwab (13) and others with some success over a limited variation of catalyst composition as represented, e.g., by alloy systems. In view of the mechanisms which have been proposed (12) that emphasize the role of an electron transfer in catalytic decomposition of H_2O_2 on metal surfaces, such correlations seemed possible here also.

Two catalysts were studied in detail (1) platinum, one of the very few metals which does not dissolve during the decomposition reaction. It has been extensively studied in the past, but usually in the colloidal state or as platinum black. (2) Silver, one of the most active metals for H_2O_2 decomposition.

Procedure

It is more difficult to compare activities in systems using powdered metal catalysts or colloidal precipitates in a liquid than in those involving plane surfaces of known area and more characterizable fluid dynamics. Therefore, the technique used here consisted in exposing a known area of a pure metal foil catalyst to hydrogen peroxide of known concentration and at a constant temperature of 30°C. The rate of reaction was determined by measuring the rate of oxygen evolution.

The principle apparatus used here involved, in essence, a method of rotating a cylindrical specimen of catalyst mounted on a Teflon cylinder -- termed here the "rotating cylinder technique" (see Figure 1). For relatively slow reactions, a much simpler apparatus was used in which the catalyst was suspended in H_2O_2 from a glass hook or Teflon chip. The H_2O_2 was kept well stirred by a magnetic stirrer using a Teflon-coated magnetic core. This second method is termed here the "foil suspension technique." Close control of the temperature for both types of reactors was achieved through the use of either a thermostatic bath in the case of a relatively inactive catalyst, e.g., platinum, or a cooling coil in the reactor monitored by a thermoregulator in the case of an active catalyst, e.g., silver. In both pieces of apparatus all surfaces in contact with liquid H_2O_2 other than that of the catalyst were either glass or Teflon. Two variations of the rotating cylinder apparatus were used. That for the most active catalysts, which required the most elaborate design, is shown in Figure 2. It used an aluminum plate as a reactor head sealed to the glass vessel against a dual gasket of rubber and Teflon. Certain catalysts, particularly silver, were studied in the form of a foil 0.001" thick which was fastened to form a smooth band on the center of a Teflon cylinder. The cylinder in turn was secured onto an aluminum shaft which

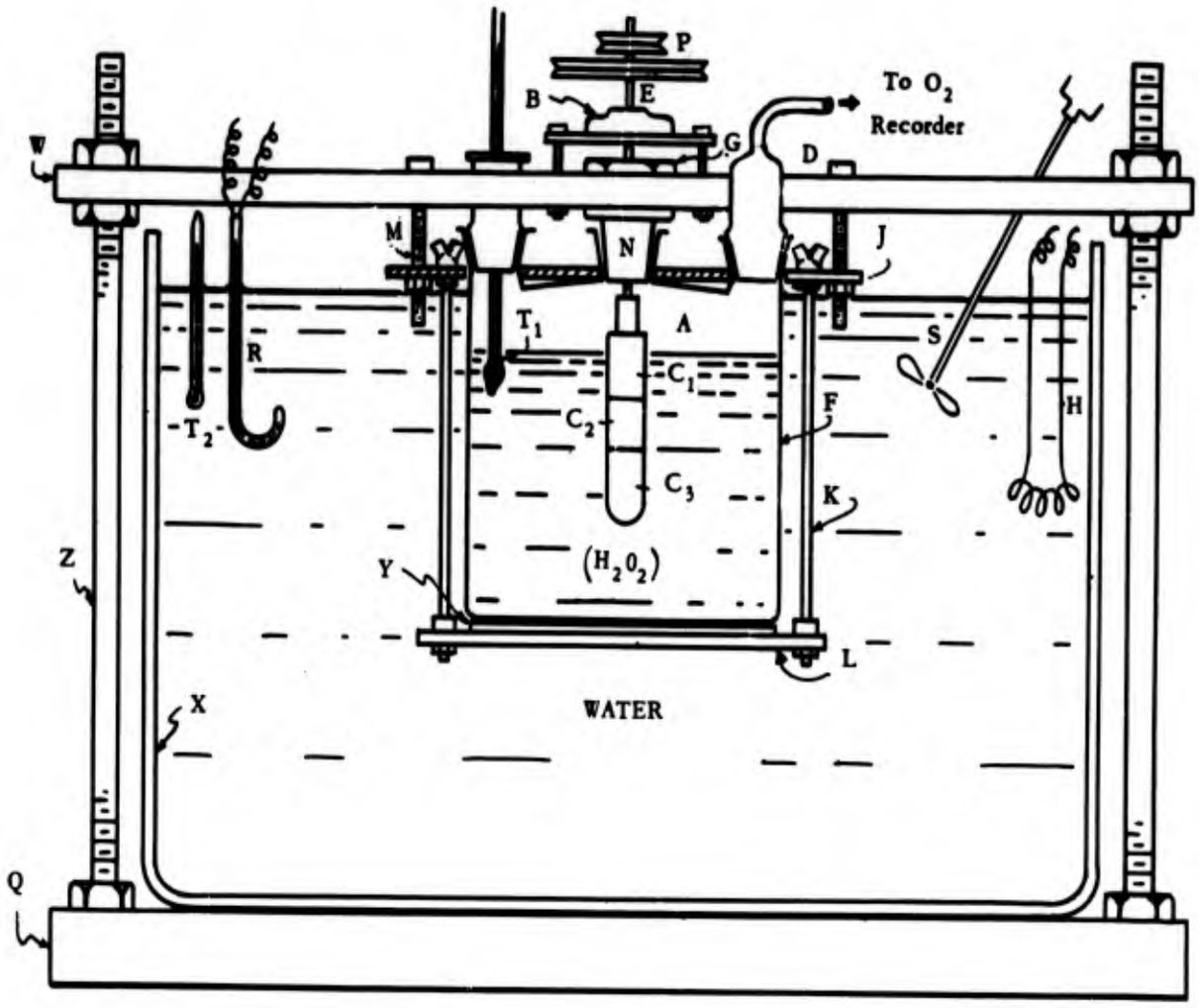

FIGURE 1

ROTATING CYLINDER TECHNIQUE FOR MEASUREMENT OF CATALYSIS OF LIQUID HYDROGEN PEROXIDE BY METALLIC SURFACES.

LEGEND:

- A Aluminum shaft, 3/8 inch dia.
- B Ball-bearing, self aligning type
- C_1 Teflon cylinder, 3/4 inch dia., counter-sunk taper end.
- C_2 Teflon cylinder, 3/4 inch dia., tapered ends, split.
- C_3 Teflon cylinder, 3/4 inch dia., end piece, rounded and closed.
- D Oxygen outlet. 24/40 ⅋ connected by tygon tubing to recorder of O_2
- E Stainless steel shaft, 5/16 inch dia.
- F Pyrex flask, cylindrical, flat bottomed, 3 necked, 24/40 ⅋ joints.
- G Brass packing screw cum oil seal.
- H Heater
- J Plexiglas plate, 1/2 inch thick.
- K Brass tie rods with brass screws at the ends.
- L Aluminum plate 1/4 inch thick.
- M 1/2 inch steel Bolts.
- N Teflon stopper with gland.
- P Fiber pulley.
- Q Wooden base
- R Thermo-Regulator
- S Stirrer
- T_1 Thermometer, with ⅋ joint.
- T_2 Thermometer
- W 3/4 inch steel plate, 3X 16 inches.
- X Pyrex Jar, 12 inch dia.
- Y Rubber gasket.
- Z 3/4 inch steel rods with lock nut.

ran through a Teflon packing gland in the reactor head. The Teflon cylinder rested on a Teflon bearing and the whole assembly was carefully machined to run without eccentricity. Certain other less active catalysts in which mass transfer was not controlling were studied in the form of 3/4" diameter cylinders suspended from above but without a lower bearing surface.

The reactor of Figure 2 was fixed on a Cenco Lab-Jack so it could be quickly raised or lowered at the beginning and end of a run. A pulley and cone reduction system was used to permit infinite variation of stirrer speeds from 0 to 1200 r.p.m. Temperature control was maintained by a glass cooling coil through which water was circulated from either of two baths, one held at 30°C. and the other at 0°C. A thermoregulator switched from one source of water to the other to maintain reaction temperature at 30°C. $\pm$ 0.5°C. even at the highest rates studied. It was possible to measure the rate of oxygen evolution at rates from about 0.01 ml/min to 5 liters/min, using an inclined burette system for the lowest, and various combinations of rotameters checked by a precision wet test meter at higher rates.

When pure metal foils were used they were degreased by successive treatments with carbon tetrachloride, acetone and conductivity water. The hydrogen peroxide was unstabilized SP-90, obtained from Becco Chemicals, Buffalo, New York. This is the purest grade commercially available, comparable to that of doubly

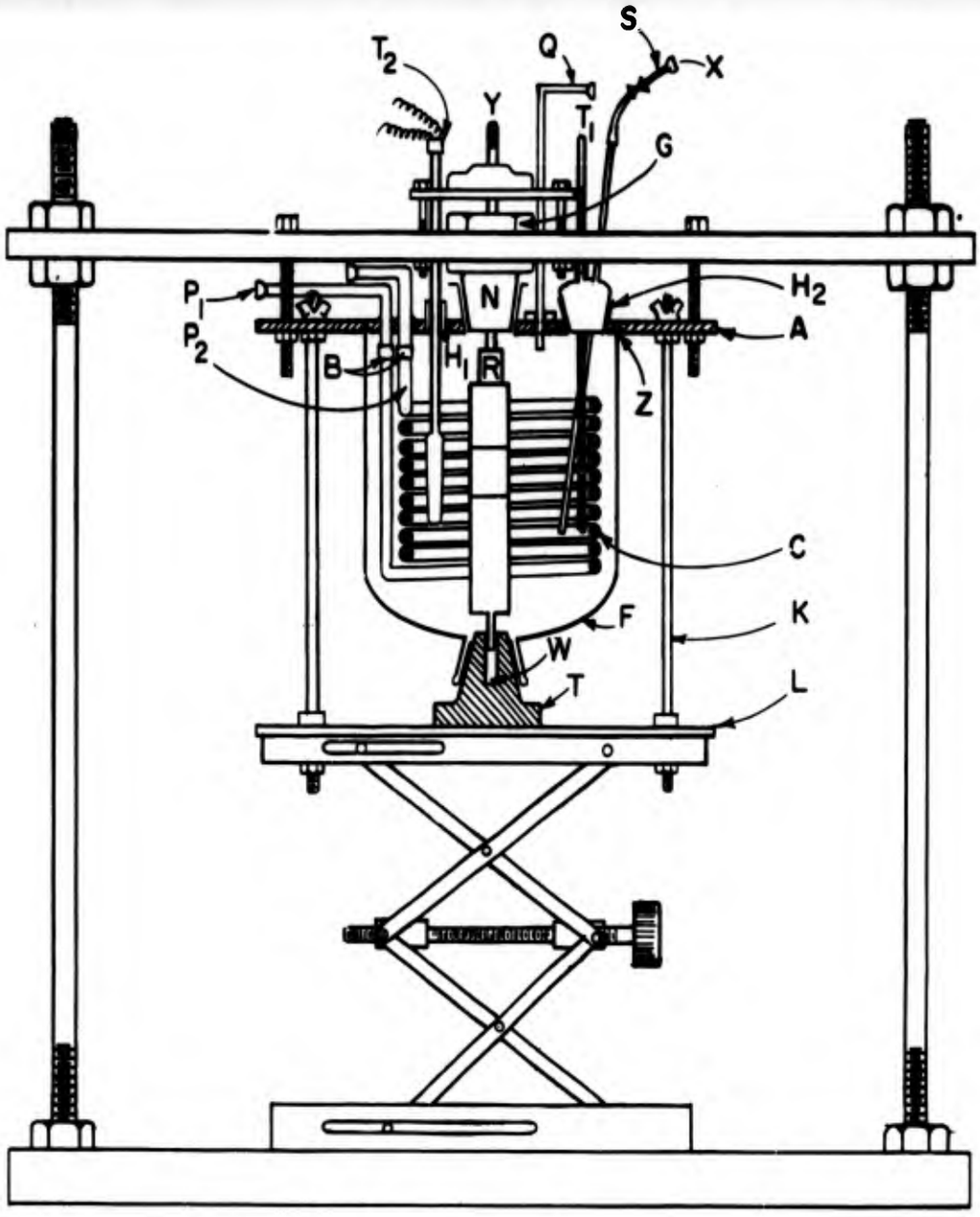

FIGURE 2

ROTATING CYLINDER TECHNIQUE FOR MEASUREMENT OF FAST DECOMPOSITION OF LIQUID HYDROGEN PEROXIDE ON METALLIC SURFACES.

Legend

A = Reactor head, ½" aluminum
B = Std. taper joints
C = Glass coil
F = Reactor, cut bottle
G = Brass packing screw and seal
H_1 = Fitting, aluminum
H_2 = Tapered, fitting aluminum
K = Brass tie rods with brass screws at ends
L = Base plate, brass
N = Teflon gland
P_1 = Inlet for cold water
P_2 = Outlet for water from coil
Q = Outlet for oxygen
R = Rotor
S = Sampling probe
T_1 = Thermometer
T_2 = Thermoregulator
T = Teflon stopper and bearing
W = Well for bearing
X = Spherical std. taper joint
Y = Driving shaft
Z = Teflon, gasket

distilled water and contained not more than a few p.p.m. of stabilizer or impurities. Water for dilution was first distilled and then percolated through a mixed ion exchange column.

For comparative studies on various heavy metals, the simpler magnetically stirred reactor was used with a metal foil suspended in the hydrogen peroxide. In each case the hydrogen peroxide was first brought to a de-supersaturated condition by constant agitation until a steady state rate of oxygen evolution had been reached.

Surface area measurements were made according to the technique proposed by Wagner (17), where a change in the polarization capacity of the surface is measured for a small change in its potential, using a ballistic galvanometer. Electron micrographs provided a visual check on the above results. Electron diffraction and X-ray diffraction studies were also used to gather additional information about the surface and bulk structure of the catalysts. High speed photography was used to observe the bubble activity on the surface of the catalyst.

RESULTS AND CONCLUSIONS

Supersaturation Phenomena

Much past confusion in interpretation of H_2O_2 decomposition rate studies has stemmed from neglect of the fact that oxygen supersaturates in H_2O_2 to a sub-

stantial degree (11). Consequently the initiation or cessation of agitation will cause release or accumulation of oxygen in solution. If the reaction rate is being followed by measurement of the oxygen evolution rate, this can lead to much misinterpretation. The phenomenon is illustrated by Figure 3 which shows a series of runs on one platinum foil using the rotating cylinder technique. The runs were carried out in the order shown, on one batch of H_2O_2, and with a quiescent period following each run. It is seen that each curve obtained of rate versus time consisted of an unsteady state region and a steady state region. The former was a function of rotation rate and the length of the quiescent period before the experiment was started. If a run at low rotation rate follows one at high rotation rate after a relatively short quiescent period, the initial O_2 evolution rate is low and climbs; if a high rotation rate follows a low rotation rate, the initial O_2 evolution rate is high and drops. In this system about 15-20 minutes of agitation is required for a steady-state rate of O_2 evolution to be reached. Without agitation some period greater than 30 min. or so is required. Previous workers had variously interpreted such supersaturation effects as the variations in catalytic activity of platinum or mass transfer effects, leading to the erroneous conclusion that rates on smooth platinum were mass transfer controlled (4, 14, 16).

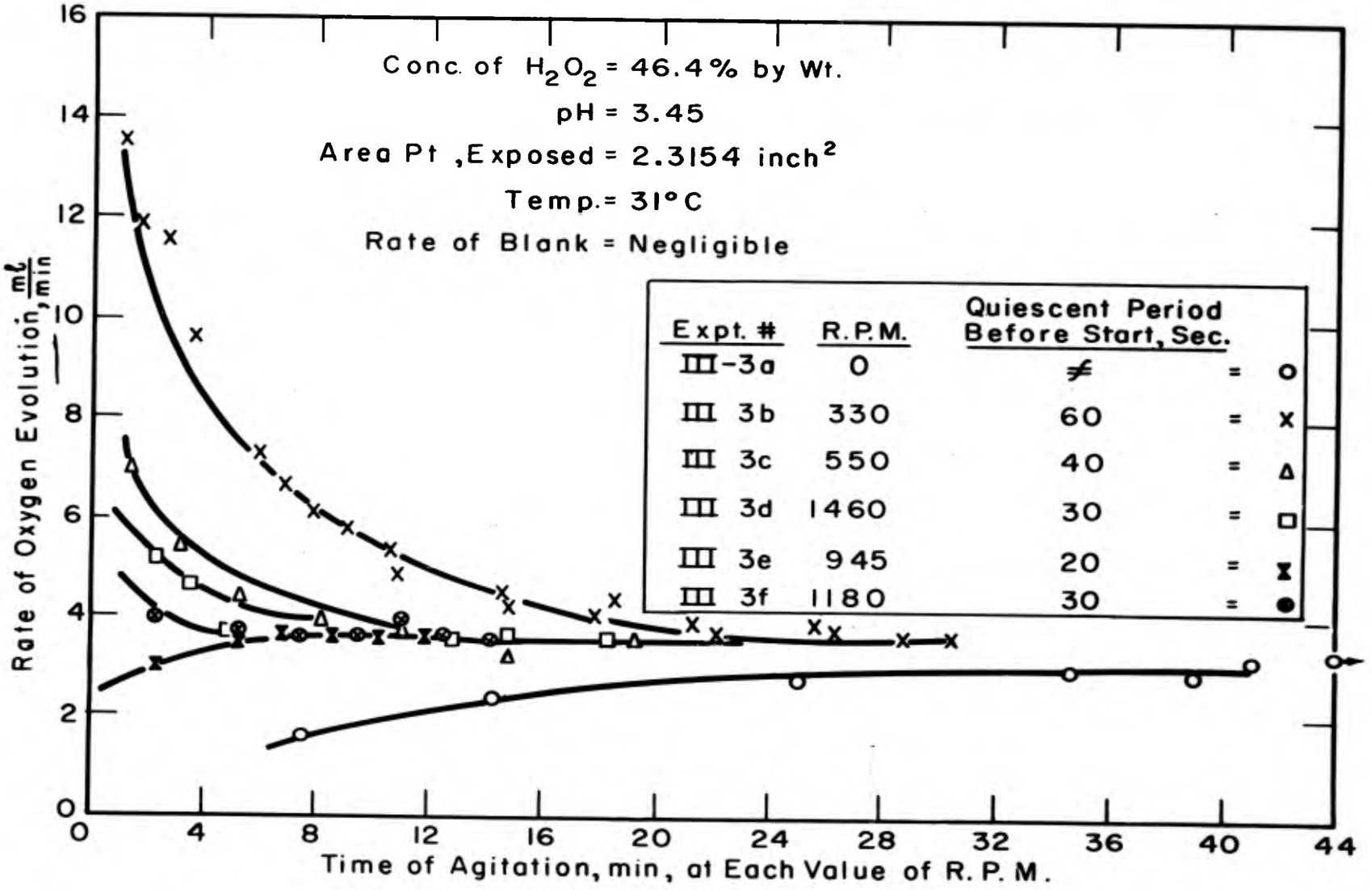

FIG. 3–EFFECT OF AGITATION ON RATE OF OXYGEN EVOLUTION FROM HYDROGEN PEROXIDE IN CONTACT WITH PLATINUM

Results - Platinum

The steady state region of the curves as shown in Figure 3 was found here to be independent of rotation rate in the case of platinum and represents, we believe, the true catalytic rate free of supersaturation effects. Similar steady state data were obtained at concentrations from 0 to 89% w/w hydrogen peroxide, a new, fresh platinum foil being used for each run. (See Figure 4.) The catalytic rates on smooth platinum foils throughout this range are surface rate controlled and follow first order kinetics.

The activity of fresh platinum foils was found not to be affected by heat treatment at temperatures of 400-1300°C., but did vary with the source. (Compare foils A and B in Figure 4. Reports elsewhere of loss of activity of platinum on silica by heat treatment probably reflect a sintering or crystal growth process occurring only with finely divided material. However the activity of smooth platinum did decrease with increased frequency as well as increased length of contact with peroxide, in agreement with earlier reports. Over a considerable series of runs, the platinum activity dropped to as low as a tenth of its original value. For example, the activity of the platinum foil in Figure 3 was substantially below that of the fresh foils represented in Figure 4. The curves of Figure 4 may be taken as representing maximum figures for smooth platinum foils as available to us.

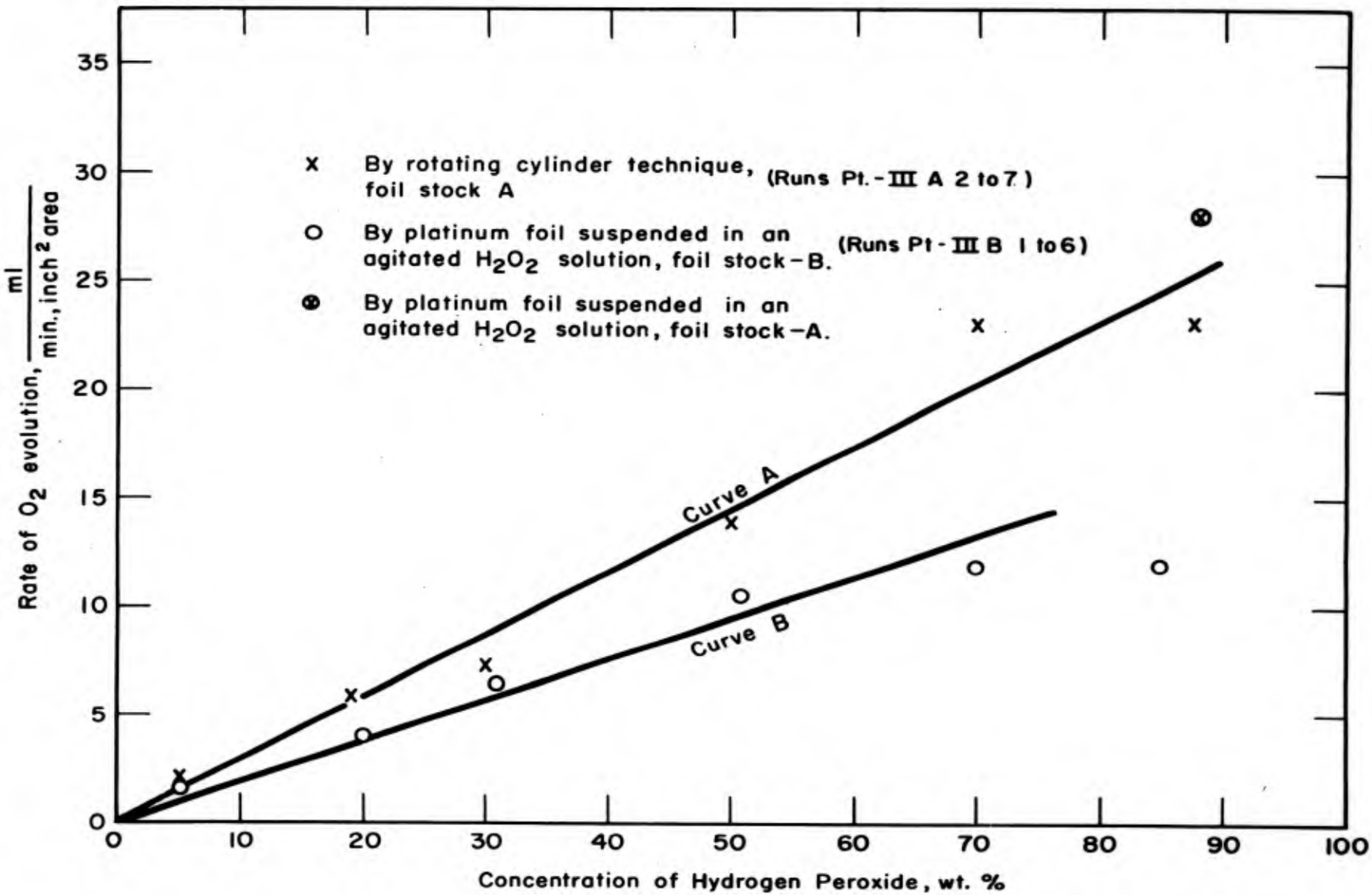

FIG. 4 RATE OF O_2 EVOLUTION vs. THE CONCENTRATION OF HYDROGEN PEROXIDE TEMP. = 30° C

If one estimates the area of the platinum foil of Heymann, who also used a rotating cylinder technique, but did not report catalyst area, his rates are not higher than those obtained here and hence could not have been mass-transport controlled as he reported. Similarly Teletov's results (16) clearly indicate neglect of the desupersaturation phenomena. Likewise in the work of Lopatkin and co-workers (7) the inference that the catalytic rate on platinum supported on silica gel is second order could arise from neglect of desupersaturation, since the basis for this formulation was an oxygen evolution versus time curve.

At relatively low rates of decomposition, e.g., as found with platinum and in the absence of agitation the rate of oxygen release increases with time over the first few minutes even after supersaturation effects have been eliminated. Photographs made at one microsecond exposure during the reaction show that a large fraction of the surface becomes blanketed by bubbles during the first few seconds of contact. As these bubbles grow and detach an increasing fraction of the platinum surface becomes free of gas and a steady state condition becomes apparent in the photographs at the same time that the oxygen evolution rate becomes constant.

X-ray diffraction studies showed that the smooth platinum was highly textured, with a preferential orientation of the 220 and 311 faces of the crystal lattice.

Electron micrograph replicas at about 3000 magnification showed that the platinum foil was quite smooth and plane except for a gradual waviness of the surface, about 0.4μ in height.

Silver

Silver slowly dissolves in H_2O_2 while catalyzing the decomposition, but the dissolved silver is not a significant catalyst. For present purposes the surface decomposition reaction and the solution process may be regarded as separate phenomena. However, a silver-containing solid may be precipitated from solution as the decomposition reaction causes the H_2O_2 concentration to drop, thus causing a rise in the overall decomposition rate.[8, 12] Typically, with massive silver in, say 20 wt. % H_2O_2, the rate rises during the first 3-5 minutes until a steady-state bubble configuration is reached, remains relatively constant until 30-35 minutes have passed and then begins to rise substantially as silver begins to become precipitated from solution. (See Figure 5 for a typical curve at 20 wt. % H_2O_2.) Figure 6 shows the degree of reproducibility obtained through a set of 8 successive runs with 37 wt. % H_2O_2, taking the relatively steady-state period as representative. The curves of rate vs. concentration appear identical and the experimental scatter is within ± 10%. Rate data as a function of concentration on the silver foil catalyst are shown in Figure 7 (1). The observed rate of catalysis shows a

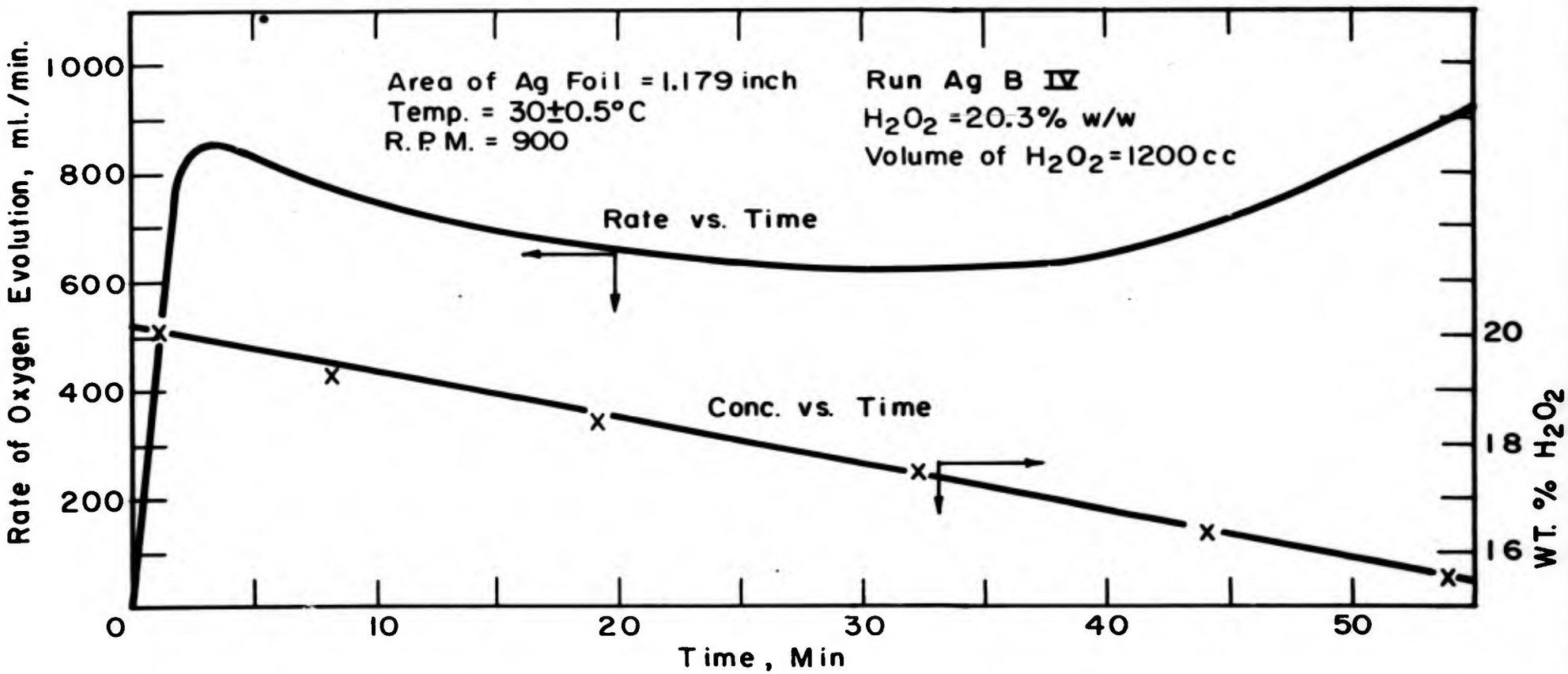

FIG. 5 CATALYTIC DECOMPOSITION OF H_2O_2 BY SILVER FOIL

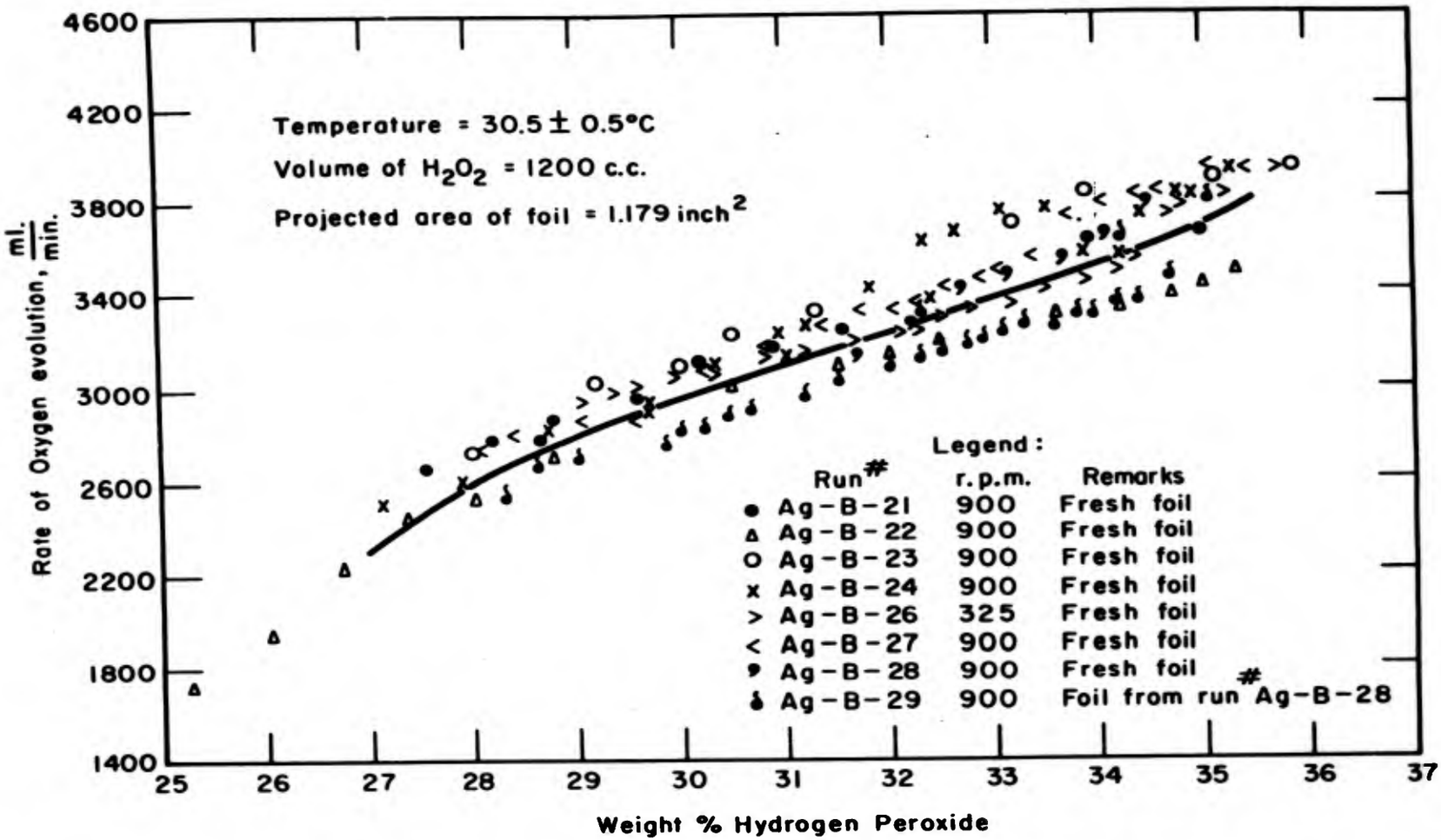

FIG. 6 RATE OF DECOMPOSITION OF H_2O_2 ON SILVER-FOILS vs CONCENTRATION OF H_2O_2

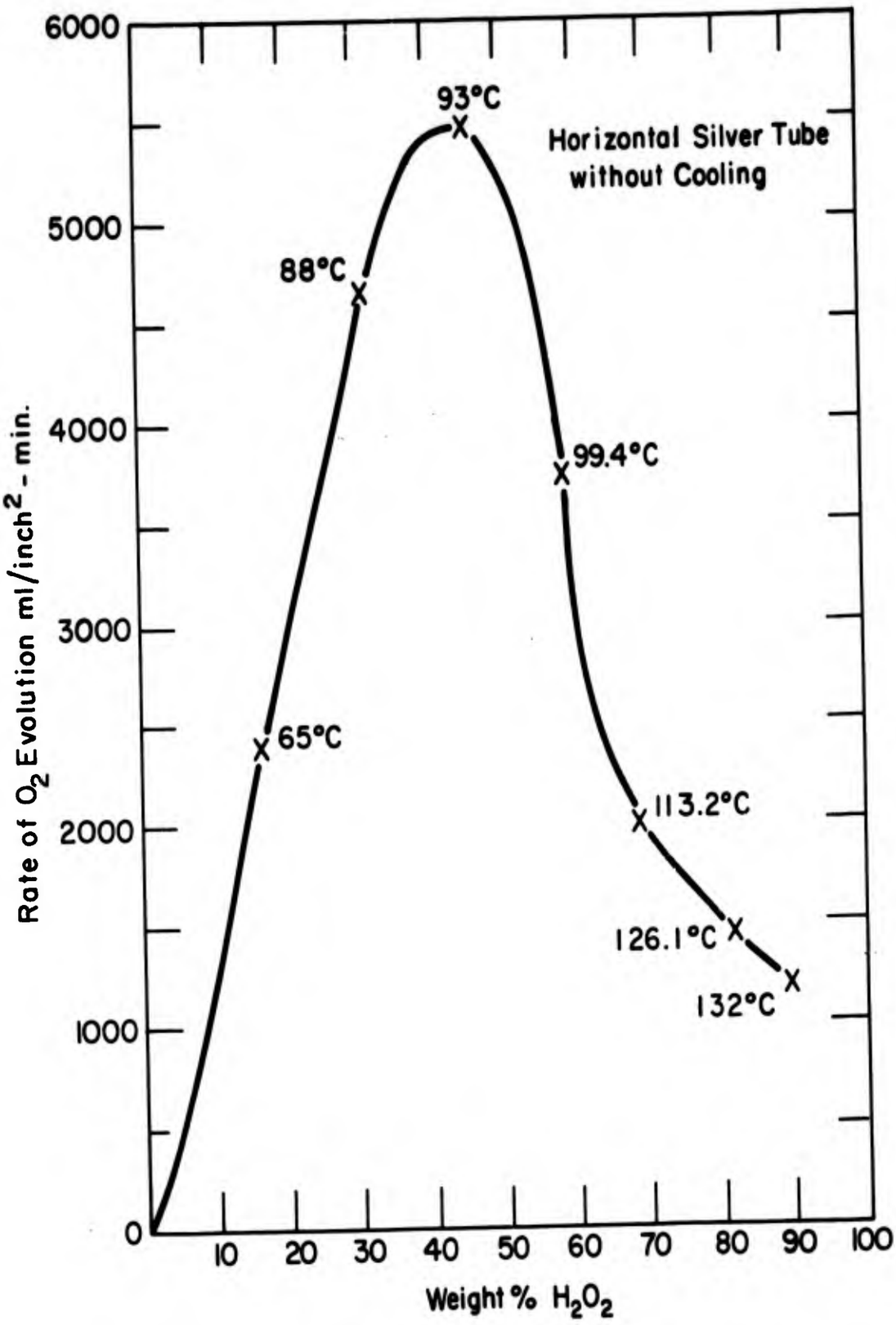

FIG. 7 - DECOMPOSITION RATE AS A FUNCTION OF CONCENTRATION

rise with concentration of hydrogen peroxide up to a maximum at about 45% by weight but then decreases with increased concentration.

Rotation rate had no effect on decomposition rate at any H_2O_2 concentration studied, within the accuracy of the experiments. The decomposition rates were also found to be independent of rotation rate when the silver catalyst foil was made the internal surface of a cylinder, (using 32% wt. H_2O_2) or a set of four baffles were provided at a distance of 0.005" from the catalyst surface.

Theoretical: Consider the conditions under which surface reaction and mass transport to the surface each contribute some resistance to the overall rate. Assume that the resistance to mass transport is confined to that of a stagnant film of thickness δ and that a fraction, f, of the catalyst surface is blocked from reaction by gas bubbles.

The rate of reaction, $\frac{dn}{d\theta}$ may then be expressed as:

$$dn/d\theta = \frac{D \sigma_p (C_b - C_s)(1 - f)}{\delta} = k_s \sigma_t C_s^n (1 - f) \quad (1)$$

If the surface reaction is approximated as first order, the rate can be recast into the form:

$$dn/d\theta = \left(\frac{1}{\frac{\delta}{D} + \frac{1}{k_s R}} \right) C_b \sigma_p (1-f) \quad (2)$$

where δ is the fictitious film thickness for resistance to mass transport

D is the diffusivity

σ_p is the geometrical (projected) area of the catalyst

σ_t is the true catalyst area

C_b = H_2O_2 concentration in bulk

C_s = H_2O_2 concentration on catalyst surface

k_s = the reaction rate constant for surface reaction

f = the fraction of the catalyst surface blocked from reaction by gas bubbles

R, = the roughness factor, = σ_t / σ_p

In liquid-solid systems in which no gas is evolved, δ decreases with increasing turbulence so an observed increase in reaction rate with increased rotation rate provides strong evidence for a mass-transport-limited process. Under such conditions, knowing the diffusivity, D, the value of δ can be calculated and such values have been reported for a variety of solid-liquid systems. When a gas is evolved, however, two new situations develop (1) the gas bubbles may at any given instant block a fraction of the surface from reaction and (2) bubble evolution itself creates turbulence, thereby affecting the value of δ. At sufficiently high reaction rates, the interfacial turbulence produced by the bubbles could presumably be so great that a change in rotation

rate would have no effect even where the process is completely mass-transfer controlled.

Comparison to Other Solid-Liquid Systems with Gas Evolution

A physical system similar to the present one is that of dissolution of metals in acid. However, most workers have used a depolarizer, to avoid the complexities of gas evolution. The only other study in which attention has been focussed on the effects of gas evolution appears to be that of Roald and Beck (10) who studied solution of magnesium in HCl. They used the rotating cylinder technique and found the process to be diffusion-controlled. They reported that the higher the reaction rate, the less the effect of rotation rate on chemical rate and at the highest dissolution rates they studied -- corresponding to the evolution of about 400-600 cc $H_2/(in^2)$ (min) -- variation of rotation rate from 6400 r.p.m. to 0 had negligible effect on dissolution rate, less than 10%. Here turbulence caused by bubble evolution presumably assumes a major role as compared to turbulence imposed by rotation. Their highest rates correspond to about the lowest studied here with silver. From this one can conclude that all of the silver studies here were mass transfer controlled and it is not surprising that rotation rate had no effect in our work.

Another comparison is available from some data reported by Maggs and Sutton (8) on the rate of solution

of silver in H_2O_2 during decomposition. Using short lengths of silver wire, the rate of solution was found to decrease steadily with time and approached zero as the solution became essentially saturated with silver. This behavior, characteristic of a diffusion-controlled system, was found for H_2O_2 concentrations of 50 wt. % and above at 2°C. and for all lower concentrations studied at 25°C. (apparently down to about 20%).

They made further studies by rotating a 1 cm. diameter silver cylinder. Eight runs were reported, seven at 5000 r.p.m. and one at 2000 r.p.m. Unfortunately decomposition rates were apparently not measured. From the rate of solution of silver, the value of δ was calculated to be about 0.33×10^{-3} cm. at 10°C. and 5000 r.p.m. and about 0.6×10^{-3} cm. at 2000 r.p.m. These values may be contrasted to values of approximately 3×10^{-3}cm. reported for various other dissolution systems having no forced turbulence or bubble evolution. In view of the small effect of rotation rate on dissolution rates reported by Roald and Beck at 400-600 ml/(min)(in^2), the pronounced effect of rotation rate reported by Maggs and Sutton for one run is somewhat surprising since the bubble evolution rate in their case was presumably substantially greater. However, the value of δ is probably affected by average bubble size on detachment as well as total gas evolution rate. There are indications from studies of heat transfer to liquids that for a given rate of gas or vapor evolution, the heat transfer rate

increases with the number of bubbles, i.e., with decrease in average size (9). Hence the effect of forced turbulence on mass transfer rates -- for a given rate of gas evolution -- may well vary considerably for different chemical systems.

If it is assumed that our rates at 30°C. over the range of 10-35% H_2O_2 are mass-transfer controlled, using values of H_2O_2 diffusivity extrapolated from the data of Stern (15) one can calculate the corresponding value of δ to be about 0.4×10^{-3} cm. This compares well with those reported by Maggs and Sutton considering that we have a lower rotation rate but a higher temperature which causes faster reactions and bubble formation that presumably create more turbulence. It is particularly significant that bubble evolution can cause the mass transport rate to be increased by a factor of as much as 10 over that encountered without gas evolution.

The decrease in rate with increase in H_2O_2 concentration above about 50% apparently represents increased blockage of the silver surface by gas bubbles. The phenomena seem to be similar to those occurring in heat transfer at the transition region from nucleate to film boiling.

The maximum reaction rate found here, 3500 ml O_2/(min)(in^2) or 1070 ft.^{3} O_2/(hr.)(ft.^{2}) corresponds to a heat release rate of 2.25×10^5 BTU/(hr)(ft.^{2}), all of which must be transferred from the silver surface to the

bulk liquid. The only heat transfer studies in a system physically analogous to the present one are those of Mixon and co-workers[9] who produced bubbles electrolytically on a heated surface immersed in a dilute aqueous electrolyte. They reported heat fluxes for surface boiling with superimposed electrolytic action under conditions where the bath temperature was maintained at temperatures generally of 80-100°F. However, their gas evolution rates ranged up to a maximum of only about 16 $ft.^3/(hr)(ft.^2)$. They reported that heat transfer coefficients were increased by bubble evolution up to rates of about 0.75 $ft.^3/(hr)(ft.^2)$ beyond which, the greater the bubble evolution rate the greater the surface temperature required to produce a fixed heat flux, i.e., an effect analogous to the transition from nucleate to film boiling.

Surface area measurements on the silver catalyst by the polarization capacity method (17) showed that the increase in area (here equals the roughness factor, R) after reaction amounted to a factor of from 1 to 3. Electron micrograph replicas of the silver (10a) showed that before reaction it was essentially plane, with a few minor parallel scratches no greater than about 0.05 μ in dimension. After reaction, the silver surface showed a random etching pattern varying with experimental conditions, the depth of the etching being of the order of 0.15 μ or less.

Comparison of the roughness factor with the rate of decomposition showed a slight correlation between higher rates and lower values of R, contrary to what might have been expected, although the effect is partly hidden in the adventitious variation of results from run to run. It appears that the primary factor is the blocking of a portion of the catalyst surface by gas bubbles and the effect of variation of etching pattern and R is primarily upon the detachment of bubbles.

X-ray diffraction studies showed that the silver surface was highly textured, with a preferred orientation of the 220 and 111 faces of the crystal lattice. As would be expected, there was no change after contact with H_2O_2. The silver foils, which are initially shiny, always become dull white in appearance after exposure to H_2O_2. This might be caused by an oxide, by etching, or by both. Electron diffraction patterns were obtained on silver, before and after contact with H_2O_2, on analytical grade silver oxide, and on scrapings from exposed silver. Comparison showed a faint indication of presence of silver oxide on the exposed silver, but definitely no gross oxide film was present.

The rates of decomposition of H_2O_2 on 12 other heavy metals were studied briefly under a standardized set of experimental conditions so as to provide a comparison of their relative activities. The results are

shown in Table I. A striking observation is the enormously greater intrinsic activity of silver relative to that of any other metal. A number of solid catalysts for hydrogen peroxide show a remarkable periodicity of action (12, p. 498) or rates which vary irregularly with time. Cobalt, copper and lead showed the phenomenon here; the other nine metals showed fairly reproducible activities. The periodicity is probably caused by an alternation between a diffusion-limited process and a true surface-limited process in which the state of oxidation of the metal surface, and hence its activity, changes with the composition of the liquid in immediate contact with the surface. It was attempted to relate the activities with various electronic or other parameters as shown on the table, but no correlation was found.

It is now known that even those metals very low in the electromotive series such as platinum and gold acquire an oxide film in contact with strong oxidizing agents so that the active surface in contact with the H_2O_2 here was probably an oxide layer in most cases. It is possible that an oxide film builds up to varying extents on the different catalysts and that electron transfer through this surface barrier is the rate limiting process.

Literature Citations

1. Audibert, F., M.S. Thesis in Chemical Engineering, M.I.T., 1960.

2. Cullity, B. D. "Elements of X-ray Diffraction," Reading, Mass., Addison Wesley (1956) p. 482.

3. Dowden, D. A., Disc. Faraday Soc., No. 8, 184-8 (1950).

4. Heymann, H., Z. Physik. Chem. 81, 204-22 (1912).

5. Hodgman, C. D., Editor, "Handbook of Chemistry and Physics," 38th Ed., Ohio, Chemical Rubber Pub. Co. (1956-57) p. 2350.

6. Kittel, C., "Introduction to Solid State Physics," New York, Wiley (1956)

7. Lopatkin, A. A., Sekhobalova, V. I. and Lebedev, V. P., Zhur. Fiz. Khim, 28, 2222-31 (1954).

8. Maggs, F. T., and Sutton, D., Trans. Faraday Soc., 54, 1861-1870 (1958).

9. Mixon, F. O., Chon, W. Y., and Beatty, K. O., Jr., Chem. Eng. Progress, 55, 49 (Oct. 1959).

10. Roald, B. and Beck, W., Jl. Electrochem. Soc., 98, 277 (1951).

10a. Sarda, P. K., Sc. D. Thesis in Chemical Engineering, M.I. T., 1959.

11. Satterfield, C. N., "Supersaturation of Oxygen in Aqueous Hydrogen Peroxide Solutions," Special Report No. 57, DSR 7476, M.I.T., 1960.

12. Schumb, W. C., Satterfield, C. N., and Wentworth, R. L., "Hydrogen Peroxide," A.C.S. Monograph No. 128, Reinhold, 1955.

13. Schwab, G. M., Trans. Faraday Soc., 42, 689 (1946).

14. Sieverts, A. and Bruning, H., Z. Anorg. Allgem. Chem. 204, 291-304 (1932).

15. Stern, K. G., Ber., 66, 547 (1933).

16. Teletov, I. S., J. Russ. Phys. Chem. Soc., 39, 1145-6 (1909).

17. Wagner, C., Jl. Electrochem. Soc., 97, 71 (1950).

TABLE I

A Comparison of the Electronic Parameters and Catalytic Activity of Metal Surfaces for Decomposition of Hydrogen Peroxide (31 Wt. % H_2O_2, T = 30°C.)

Metal	Specific Rate ml Oxygen Evolution (Min) (Sq. Im.)	Coefficient of Electronic Specific Heat, γ	Lattice Type*	Lattice Parameter* ° A	Work Function, + Electronic (Contact Method) ø, e. v.
Silver	1500	1.54-1.6	FCC	4.0856	4.21-4.47
Platinum	6	16-16.5	FCC	3.9237	4.52
Cobalt	Varying (highest recorded 150)	12	HCP	2.507	4.21
Palladium	3.7	26-31	FCC	3.8902	4.49
Iron	2.2	12	BCC	2.8664	5.36
Copper	Varying (highest recorded 11.7)	1.73-1.8	FCC	3.6153	4.46
Cadmium	1.0	1.7	HCP	2.9787	4-4.49
Molybdenum	0.3	5.1	BCC	3.1466	4.08-4.48
Vanadium	0.4	15-22	BCC	3.039	4.44
Tungsten	0.1	3.5	BCC	3.1648	4.38
Indium	0.1	4-4.3	FC Tetrag.	4.594	
Lead	Varies with oxidation state of metal	7.5	FCC	4.9495	3.94
Rhodium	0.4		HCP	2.7609	
Gold	0.4	1.8	FCC	4.0783	4.46

NOTES: (γ = coefficient of the linear term γT in the molal heat capacity, data from Kittel (6)).

BCC - Body centered cubic
FC Tetrag. = Face centered tetragonal
FCC - Face centered cubic
HCP = Hexagonal close packed

* Cullity, B. D. (2)
\+ Handbook of Chemistry and Physics (5)

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