Response of a Gas-Liquid Interface to Concentration Pulses

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Frequency response techniques have been applied to the study of the absorption of sulfur dioxide in water contaminated with surface-active material. The response of a 0.004-in. thick liquid film to a sinusoidally varying sulfur dioxide pressure was determined for frequencies ranging from 1 to 10 cycles/sec.

The total quantity of sulfur dioxide dissolved was observed continuously by measuring the amount of light absorbed by the film, which contained a colored hydrogen ion indicator, cresol red. The experimental data indicate the presence of a surface phase of absorbed material through which sulfur dioxide molecules had to pass as they were transferred between gas and liquid. In addition to creating a resistance to the passage of the solute molecules, the surface phase apparently was capable of storing the molecules temporarily.

The problem of gas absorption in a liquid has been studied by many investigators with a wide range of experimental methods. Most of these studies have had in common the use of a relatively fresh surface at which the absorption took place. The contact times in liquid jets and short wetted-wall columns usually range from 0.001 to 1.0 sec. Various early workers (*6*, *10*, *12*) measured absorption rates which were lower than that predicted by the penetration theory and interfacial resistance was suggested as the cause. Recent studies (*3*, *13*, *16*, *19*), making use of the liquid jet apparatus and various forms of the short wetted-wall column, have shown good agreement between the penetration theory and the measured absorption rates for carbon dioxide and sulfur dioxide. It appears likely that in some cases the low absorption rates originally found were caused by deviations from the simple flows assumed and to end effects rather than to interfacial resistance. However, recent studies by Chaing and Toor (*2*) and by Harvey and Smith (*8*) showed low absorption rates for oxygen and carbon dioxide, respectively, in water which could not be explained in terms of fluid dynamics or end effects.

The purpose of this study was to investigate the absorption of a gas in a liquid under conditions such that the fluid was stagnant (thus eliminating the problem of inaccurately known fluid motion), yet subject to the high mass transfer rates necessary to detect nonequilibrium at the interface. Furthermore, the use of a stagnant interface is desirable for a study of the effect of surface-active agents on absorption rates, since the interface will be at an equilibrium state and the excess surface accumulation of the surface active agent will be known.

Recently a similar investigation was carried out by Harvey and Smith (*8*) where the transient concentration gradients produced in a liquid by a step function in carbon dioxide pressure were measured with an interferometer. Measurements made with the stagnant interface contaminated with a surface active agent indicated the presence of an interfacial resistance, while measurements made with a clean interface showed good agreement with the penetration theory.

In this work the average sulfur dioxide concentration in a thin water film was measured continuously as the sulfur dioxide pressure was varied sinusoidally at frequencies ranging from 1 to 10 cycles/sec. The peak-to-peak amplitude of the pressure variation was 0.2 atm. Measurement of the phase lag and amplitude of the liquid concentration fluctuations permitted the determination of two interfacial rate constants.

DESCRIPTION OF APPARATUS

A drawing of the apparatus is shown in Figure 1. It consisted of an optical system for measuring the concentration of dissolved gas in the water film, a reciprocating pump for generating sinusoidal variations in the pressure of pure sulfur dioxide over the liquid, and a pressure transducer for measuring the gas pressure. The concentration changes in the water phase produced changes in the intensity of color in a dissolved hydrogen ion indicator, cresol red. The concentration changes were followed by means of a sensitive photocell which received a light beam that had been totally reflected off the gas-liquid interface and which had passed twice through the liquid film. The observed changes in photocell current were proportional to the changes in the instantaneous average dissolved gas concentration. If $c(x, t)$ is the instantaneous concentration deviation from the time-average value at a depth x below the interface, the light signal was proportional to

$$\int_0^\infty c(x, t)\,dx = A(\omega)\, e^{i(\omega t - \theta)} \tag{1}$$

where ω is the frequency of the impressed pressure pulsations, $A(\omega)$ is the amplitude of the recorded signal, and $\theta(\omega)$ is the phase lag angle in radians. A and θ were the quantities observed experimentally as functions of frequency.

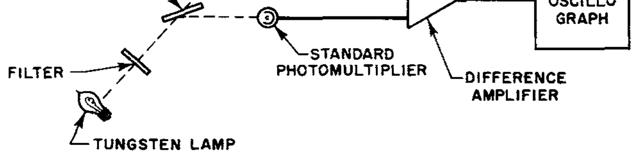

Fig. 1. Schematic drawing of gas absorption apparatus.

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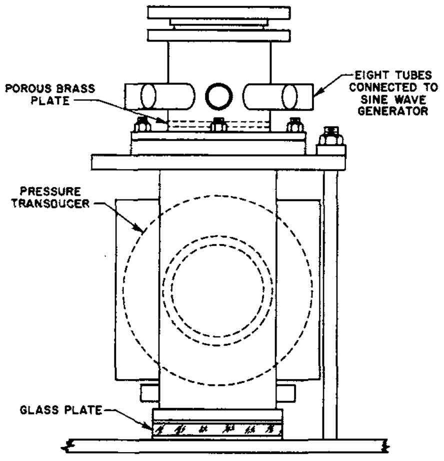

Fig. 2. Gas absorption chamber.

A drawing of the gas absorption chamber is shown in Figure 2 and its incorporation into the optical system is shown in Figure 3. Rubber tubing was used to connect the absorption chamber to the sine wave generator to minimize vibration. Figure 4 gives a detailed view of the liquid film, which was approximately 1 in. in diameter and about 0.004 in. in depth. The light beam was about ¼ in. wide. The stainless steel retaining ring around the liquid film had to be very nearly the same thickness as the liquid layer so that changes in surface tension during absorption and desorption would not give spurious signals owing to fluid motion. The porous bronze plate above the liquid was helpful in reducing motion of the liquid owing to the surging gas flow. The pressure drop through this plate was so small that only negligible differences in pressure could have occurred across it. Further details are given by Whitaker (*20*).

When detecting the rate of gas absorption by means of a hydrogen ion indicator, it is necessary that the indicator reaction rate be much faster than the rate of changes in surrounding hydrogen ion concentration. To test the reaction rate for cresol red a water solution of the indicator was brought into contact with a hydrochloric acid solution in a jet mixer (*15*). The indicator was completely changed from yellow to red in less than 5 msec., indicating a sufficiently fast reaction. Roughton (*17*) has investigated several other sulfone-phthalein indicators and has found them to be suitable for following other fast reactions in solution.

The assumption that the light signal is a measure of the concentration of dissolved sulfur dioxide also implies that the hydrolysis reaction between sulfur dioxide and water is extremely fast. This assumption is difficult to substantiate since the reaction rate constants apparently have not been measured. Lynn, Straatemeir, and Kramers (*13*) were concerned with the same question in their experiments on the absorption of sulfur dioxide into water in their wetted-wall column. They concluded that the hydrolysis reaction could be neglected in composite rate processes completed in 0.1 sec. or more. A period of 0.1 sec. in the present experiments corresponds to $\omega = 62.8$ sec.$^{-1}$, which was about the maximum frequency used.

PHENOMENOLOGICAL TREATMENT OF INTERFACIAL RESISTANCE

We suppose that the water surface is covered with a layer of absorbed molecules as indicated in Figure 5. These may be either solute molecules, which are in the process of transition from gas to dissolved states, or surface-active color producing molecules, which were added originally to the water phase and which have concentrated in the interface region. The rates of addition of solute molecules to the interface layer are assumed to be proportional to p, the partial pressure of the gaseous solute touching the surface, and to the concentration of dissolved solute at the surface of the bulk liquid $c(O, t)$. The rates of escape of surface adsorbed solute into the gas or into the liquid are assumed proportional to the surface concentration of solute $c^*(t)$. The four rate constants are not independent, for at equilibrium the partial pressure and the bulk concentration must satisfy the Henry's law relationship, leading to $H = k_1K_1/k_2K_2$.

The equations to be developed below for the rate processes imply the use of Henry's law to represent the equilibrium condition. This may be seen by putting a zero on the left sides of Equations (2) and (4) and solving for the ratio $c(0, t)/p$ at equilibrium. Whether the same numerical value for H can be employed with the surfactant present as applies with pure water might be questioned, but note that the molar concentration of surfactant in the bulk liquid will be extremely small. It seems quite unlikely that any difference in H could be found. It is not implied that the literature value H applied to concentrations of sulfur dioxide in the assumed surface phase. In fact, as the data will show, the partial equilibrium constant relating surface to bulk compositions is different and is affected by the strong concentration of surfactant molecules at the interface.

A mass balance on adsorbed solute leads to the equation

$$dc^*/dt = k_1p - k_2c^* + K_2c(0, t) - K_1c^* \qquad (2)$$

In the bulk of the liquid phase the diffusion equation

$$\frac{\partial c}{\partial t} = D\frac{\partial^2 c}{\partial^2 x} \qquad (3)$$

governs the concentration distribution $c(x, t)$. The boundary conditions are

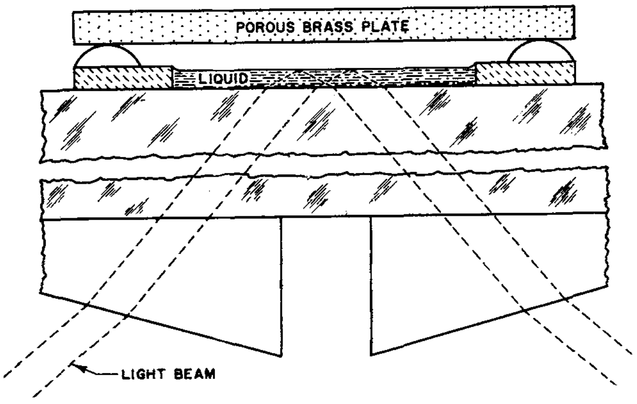

Fig. 3. Gas absorption apparatus, detail A.

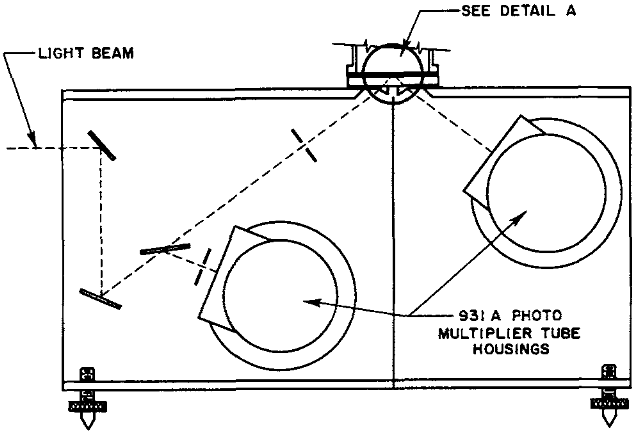

Fig. 4. Gas absorption apparatus.

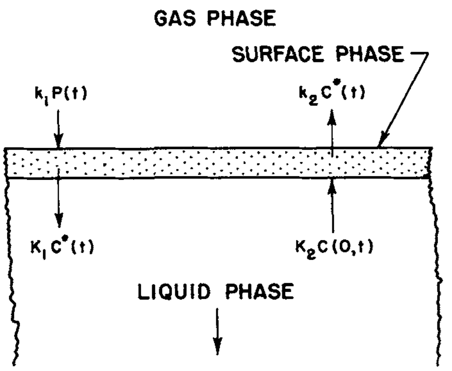

Fig. 5. Interface model used for deriving transient response.

$$-D(\partial c/\partial x) = K_1c^* - K_2c \quad \text{at} \quad x = 0 \qquad (4)$$

$$c = 0 \qquad \text{at} \quad x = \infty \qquad (5)$$

where the concentrations now represent deviations from the corresponding steady state quantities. Setting $dc^*/dt = 0$ in Equation (2) and $\partial c/\partial x = 0$ in Equation (4) gives a relationship between p and c at equilibrium which confirms the expression for H stated earlier.

A suitable solution can be found by introducing the complex quantities $C(x)$ and C^*, defined by

$$c(x, t) = C(x)\, e^{i\omega t} \qquad (6)$$

$$c^*(t) = C^*\, e^{i\omega t} \qquad (7)$$

The gas pressure fluctuation is represented by

$$p(t) = a_p\, e^{i\omega t} \qquad (8)$$

where a_p is the amplitude of the pressure excursions. The solution is

$$c(x, t) = Ha_p \frac{e^{-(i\omega/D)^{1/2}x + i\omega t}}{\dfrac{i\omega + k_2}{k_2} + \dfrac{i\omega + k_2 + K_1}{k_2K_2}(i\omega D)^{1/2}} \qquad (9)$$

and the corresponding line-integral average concentration defined in Equation (1) is equal to $c(o, t)/(i\omega/D)^{1/2}$. Note that an assumption is implied, that the observed color changes are due only to the changes in the dissolved concentration c and not to c^*; that is, that the light beam is totally reflected from the plane $x = 0$ back into the liquid.

The amplitude A and phase lag angle θ defined in Equation (1) are given by

$$A(\omega) = \frac{(D/\omega)^{1/2}Ha_p}{(m^2 + n^2)^{1/2}} \qquad (10)$$

$$\theta(\omega) = \tan^{-1}(m/n) \qquad (11)$$

where

$$m = 1 + (\omega/k_2) + (1 + K_1/k_2)(2D\omega/K_2^2)^{1/2} \qquad (12)$$

and

$$n = 1 - (\omega/k_2)(1 + (2D\omega/K_2^2)^{1/2}) \qquad (13)$$

Note that if all the interfacial rate coefficients are infinitely great, $m = n = 1$ according to Equations (12) and (13). Then the denominator in Equation (10) is constant and the amplitude of the light signal should decrease as the square root of the frequency; the phase lag, given by Equation (11), should be constant and equal to $\pi/4$. These are the results of the bulk diffusion effect alone. Deviations of the experimental results from such behavior imply interface rate limitation.

In order to consider the properties of the A and θ functions more carefully we introduce the abbreviations $u = (\omega/k_2)^{1/2}$, $\alpha = (2Dk_2/K_2^2)^{1/2}$, $\beta = HK_2/k_1$, and put the previous equations in the form

$$m = 1 + \alpha(1 + \beta)u + u^2 \qquad (14)$$

$$n = 1 - (1 + \alpha u)u^2 \qquad (15)$$

The following properties are valuable:

$$(d\theta/du)_{u=0} = \alpha(1 + \beta)/2 \qquad (16)$$

$$(d^2\theta/du^2)_{u=0} = 1 - [\alpha(1 + \beta)/2]^2 \qquad (17)$$

$$(d\ln(u\,A)/du)_{u=0} = -\alpha(1 + \beta)/2 \qquad (18)$$

If $u^* = (\omega^*/k_2)^{1/2}$ = value of u at which $\theta = \pi/2$

$$\alpha = (1 - u^{*2})/u^{*3} \qquad (19)$$

and at high frequencies, provided $\beta \neq 0$

$$\theta \approx \pi - (1/\alpha\, u) \qquad (20)$$

These equations suggest several ways for determining values of the rate constants from the data. The method used in the following was: determine $S_o' \equiv (d\theta/d\sqrt{\omega})_{\omega=0}$, $S_o'' \equiv (d^2\theta/d\sqrt{\omega^2})_{\omega=0}$, and $A_o' \equiv -(d\ln(\omega^{1/2}A)/d\sqrt{\omega})_{\omega=0}$ from the plots of observed quantities against the square root of frequency. Then, based on Equations (16) and (17) compute

$$k_2 = (S_o'' + S_o'^2)^{-1} \qquad (21)$$

and from Equation (18)

$$K_2 = \frac{(2D\omega^{*3})^{1/2}}{k_2 - \omega^*} \qquad (22)$$

From Equation (16)

$$H/k_1 = (D/2)^{1/2}S_o' - (1/K_2) \qquad (23)$$

The value of K_1 follows from

$$K_1 = (H/k_1)k_2K_2 \qquad (24)$$

No use has been made of the amplitude attenuation in formulating Equations (21) to (24). A test of the accuracy of the measurements is provided by this experimentally independent information; that is, by comparing the quantities S_o' and $-A_o'$, which should be numerically equal.

Several assumptions are implied in the foregoing derivation. First, it is assumed that the temperature of the interface is constant, despite changes in gas temperature as the gas is compressed and expanded. The justification for this assumption is too involved to be given here but is available elsewhere (*20*); it depends primarily on the much greater thermal conductivity of the liquid than that of the gas and on the presence of a small heat of solution at the interface. Second, it is assumed that the 0.004-in. layer of liquid is effectively infinitely deep. This introduces considerable mathematical simplification and is justified by the rapid attenuation of the waves of concentration, as they diffuse from the interface into the liquid, at a logarithmic rate equal to $(\omega/2D)^{1/2}$ per unit depth. Note that the ratio of concentrations at points 0.004 in. apart and at the smallest frequency used in the experiments is calculated to be 1/220. Interfacial resistance, if present, should make this ratio even smaller.

EXPERIMENTAL RESULTS

Figure 6 shows the observed phase shift data taken at two slightly different temperatures and at two bulk concentrations of the color indicator. This concentration could not be changed greatly since enough had to be

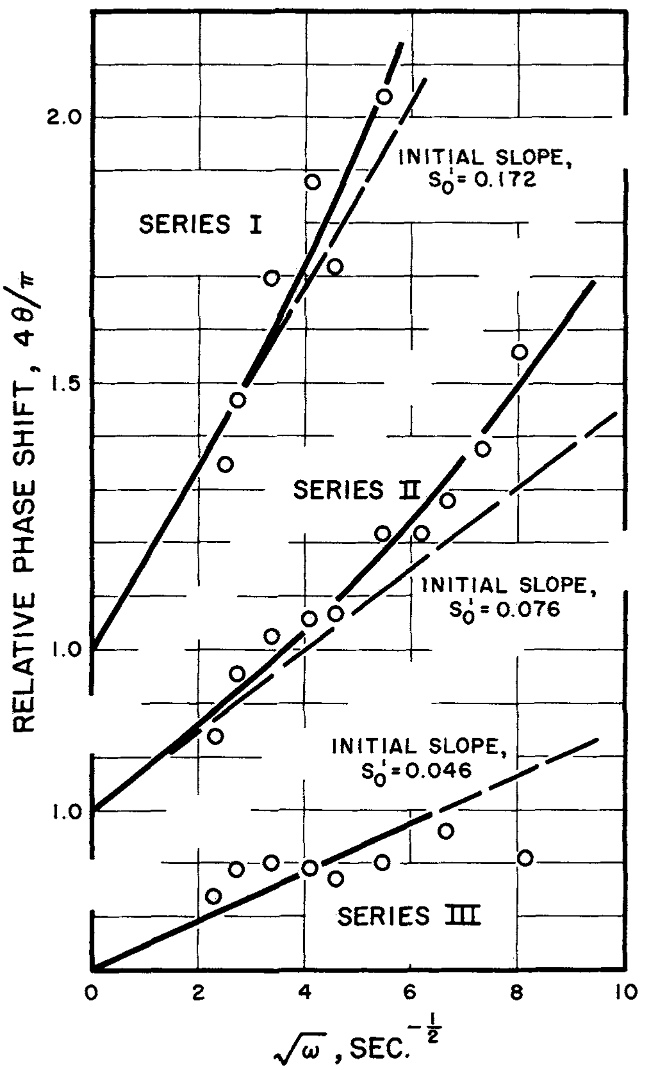

Fig. 6. Phase shift observed in light transmission experiments.

used to give an observable light absorption signal and the solubility limit could not be exceeded. Figure 7 shows the related changes in amplitude.

It is evident from Figure 6 that phase angles greater than 45 deg. were always observed, indicating that some rate process other than bulk diffusion was influencing the results. In one series of experiments a phase angle greater than 90 deg. was observed and a curve through the points in another series is clearly headed for values greater than 90 deg. This suggests that some rate process in addition to the slow transition from surface-adsorbed to dissolved states for sulfur dioxide will also be required to account for the observations.

Table 1 lists values of the four interface rate coefficients found from the data with the method outlined above. Of the values given, k_2 is probably the most reliable and K_1 the least reliable, considering the accumulation of errors and their effect on the calculated quantities. In the third series of experiments the data did not extend to high enough frequencies for all the constants to be evaluated but the value of k_2 must be greater than the quantity listed to account for the initial slope of the data on Figure 6.

The lines shown on Figure 7 passing through the amplitude data are drawn with slopes expected from the numerical values of the constants found from the phase shifts. The good agreement shows that the experimental frequency response data were internally consistent.

DISCUSSION OF RESULTS

The values of k_2 and K_1 listed in Table 1 for series 2 indicate that the average sulfur dioxide adsorbed molecule spends about 0.01 sec. on the interface, whether it leaves by escaping into the gas or into the liquid. When the surface concentration of cresol red is smaller, as in series 3, the adsorbed molecules escape more easily in an average time of about 0.001 sec. When the temperature is reduced slightly, as in series 1, the frequency of escape is reduced by a factor of three or four. The rate of capture of gas molecules by the interface phase is hardly affected but the rate of the transition from dissolved to the adsorbed states is decreased by the reduction of temperature. Obviously the small temperature difference between the experiments of series 1 and 2 was not enough for the effect of temperature on the coefficients to be determined reliably, but it is interesting that all the coefficients change in the right direction and that k_1/H is hardly affected, as is expected. The principal influence shown in the table is that of the surfactant, which appears to be very great.

From the numerical values of the rate coefficients it is possible to compute the total surface concentration of sulfur dioxide that would be at equilibrium with a pure gas at 1 atm. The apparent values are 5×10^{-6} and 2×10^{-6} g.-moles sulfur dioxide/sq. cm. of interface for series 1 and 2, respectively. The surface concentrations of cresol red are about 100 to 1,000 times smaller. (The latter were computed from measured values of the surface tension of cresol red aqueous solutions containing no dissolved sulfur dioxide with the Gibbs adsorption isotherm.) It is believed that the true surface concentrations of the color indicator molecules may have been greater

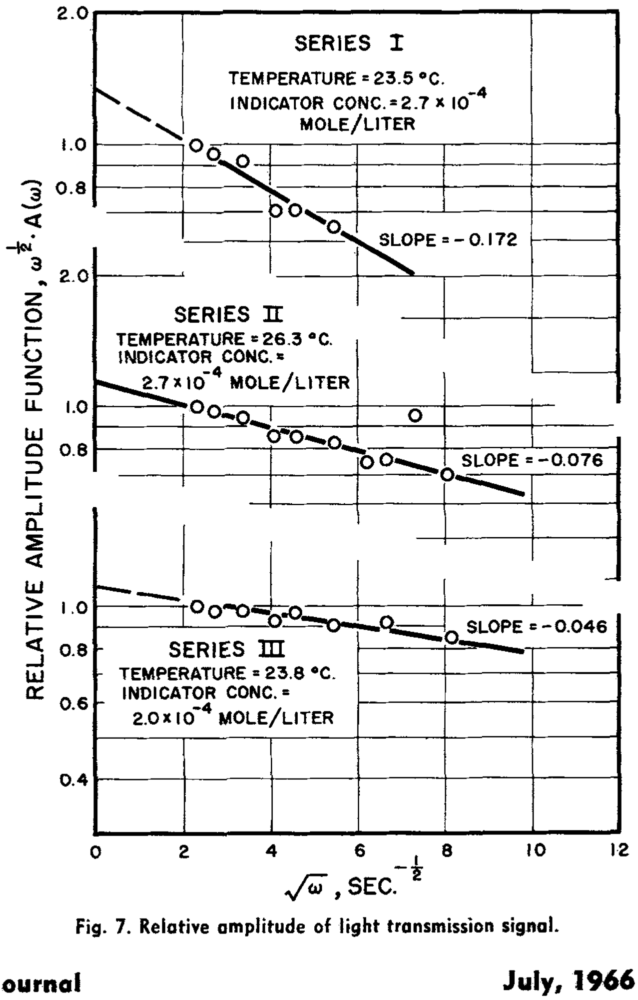

Fig. 7. Relative amplitude of light transmission signal.

TABLE 1. SUMMARY OF CALCULATED RESULTS

		Cresol red conc.		Surface rate coefficients			
				Gas to interface		Interface to liquid	
Series	Temp., °C.	Bulk, g.-mole/liter	Surface, g.-mole/sq. cm.	k_1/H, cm./sec.	k_2, 1/sec.	K_1, 1/sec.	K_2, cm./sec.
2	26.3	2.7×10^{-4}	1.2×10^{-9}	0.18	111	126	0.21
1	23.5	2.7×10^{-4}	1.2×10^{-9}	0.16	42	15	0.06
3	23.8	2.0×10^{-4}	0.6×10^{-9}	—	>800	—	—

for there was a very surprising change in the surface forces during the experiments. As the gas pressure changed the surface of the liquid appeared to move unless great care was exercised to make the surface precisely level with the metal retaining ring. Whether the addition of sulfur dioxide to the interface tended to draw indicator molecules out of the bulk solution or to repel them is not known, however. The derived equilibrium surface concentrations of sulfur dioxide can be divided by the concentration corresponding to a monolayer of sulfur dioxide molecules to find the number of layers needed to account for the total concentrations. The required number of layers is 27 and 12 in series 1 and 2, respectively.

The values of k_1/H in Table 1 may be expressed in terms of an accommodation coefficient, defined as the fraction of gas molecules impinging on the interface which are adsorbed there. The value is about 10^{-4}. Evidently a surface composed of cresol red adsorbed molecules is not able to accept very many of the gas molecules which strike it. Experimental data are available which indicate that a pure water surface does not offer a measurable resistance to the passage of sulfur dioxide molecules. Lynn, Straatemeir, and Kramers (*13*) measured the rates of absorption of sulfur dioxide into water and into aqueous solutions of hydrogen chloride, sodium bisulfite, and sodium chloride in short wetted-wall columns. Their results indicated that the interfacial rate coefficients must have been very large, even when small amounts of a surface active material, Teepol, were added to prevent ripple formation on the falling water film. The contact times in their work ranged from 0.13 to 0.98 sec., which are comparable with the range of cycle periods used here. Possibly the Teepol did not saturate the liquid surface during the short contact times; hence, no interfacial resistance was observed. In the present experiments the surface-active material had 10 to 20 min. in which to diffuse to the surface and to form a stable film.

The carbon dioxide absorption studies made by Harvey and Smith (*8*) also indicate that surface-active agents may cause interfacial resistance. They derived a rate constant comparable to K_2 from their experimental data. The values obtained for two different surface-active agents, each at two concentrations, were on the order of 3×10^{-2} cm./sec. This is within an order of magnitude of the values given in Table 1. The agreement would have been better if k_1 and k_2 had been assumed infinite, as was assumed by Harvey and Smith. On the other hand, the resistance detected by them depended only slightly on concentration of the surfactant, which was varied by a factor of 5, or on the temperature, which was changed by 2°. Lack of dependence on bulk concentration or surfactant may be due to lack of change of surface concentration. Addition of small amounts of surfactant will often saturate the surface and further additions will not alter the surface condition significantly (*1*).

In order to account for the finite values of rate constants K_1 and K_2, there must be a resistance to the passage of sulfur dioxide molecules into the liquid after they have accommodated themselves to the surface layer. When a monolayer of an organic salt is formed at a gas-water interface an electrostatic change difference between the surface and the bulk liquid is established. This is caused by the orientation of the organic molecules at the surface. The organic end of the molecule is repelled by the liquid and extends away from the surface while the ionic end anchors itself in the liquid. This preferential orientation creates a potential difference which may sometimes be as large as 250 mv. (*1*). Havinger (*20*) has discussed the effect of this potential difference on reaction velocities at interfaces and has pointed out that it will influence the energy of activation for a reaction occurring at the interface whenever the initial state and transition state have different charges. The passage of neutral sulfur dioxide molecules across the electric double layer does not suggest a transition state with a charge different than the initial state; the sulfur dioxide molecule is dipolar, however, and it may be that its passage through the electric double layer requires either a high activation energy or a specific orientation of its dipole. If such were the case, the electric double layer would create a resistance to mass transfer across an interface.

CONCLUSIONS

The experimental evidence that the amplitude of concentration pulsations in the liquid phase falls off more rapidly with increasing frequency than can be accounted for by bulk diffusion, and that there is also a greater phase lag, strongly indicates that there is some rate process, in addition to diffusion, which affects the passage of sulfur dioxide molecules into and out of a water interface. Although it cannot be claimed that the model of interface behavior which is postulated is the only one which will account for the observations, it is significant that all characteristics of the observed phase shift and amplitude are in quantitative agreement with the model used. The derived rate coefficients are at least plausibly related to the expected condition of the interface, which was composed partly of surface-active color indicator molecules.

The data imply that a gaseous sulfur dioxide molecule on its way into the aqueous phase first becomes adsorbed in a thin surface layer, after which it may either return to the gas or be dissolved in the bulk liquid at a finite rate. Likewise, dissolved molecules from the aqueous substrate enter the surface phase with some difficulty. As a result, concentration pulsations in the liquid are not as strong as they might be in the absence of interfacial resistance and the concentration peaks in the liquid are deferred. Appreciable quantities of the soluble gas are stored temporarily on the interface before they become dissolved.

NOTATION

A = amplitude of average concentration fluctuations
c = concentration of bulk phase, g.-moles/cc.
c^* = concentration of surface phase, g.-moles/sq.cm.
D = diffusivity, sq.cm./sec.
H = Henry's law constant, g.-mole/(cc.)(atm.)
K_1 = interfacial rate constant, sec.$^{-1}$

K_2 = interfacial rate constant, cm./sec.
k_1 = interfacial rate constant, (cm.)(g.-mole)/(cc.)(atm.)(sec.)
k_2 = interfacial rate constant, sec.$^{-1}$
m = function defined by Equation (12)
n = function defined by Equation (13)
P = pressure, atm.
S_o' = observed value of $d\,\theta/d\sqrt{\omega}$ at $\omega = 0$
S_o'' = observed value of $d^2\theta/d\sqrt{\omega}^2$ at $\omega = 0$
T = temperature, K.
t = time, sec.
x = distance, cm.

Greek Letters

ω = frequency, sec.$^{-1}$
ω^* = value of ω at which the observed value of $\theta = \pi/2$
θ = phase angle, rad.

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Continuous Ion Flotation of Dichromate

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An experimental investigation is presented of the continuous ion flotation of dichromate, with the use of a cationic surfactant, ethylhexadecyldimethylammonium bromide. The feed streams contain concentrations of dichromate ($Cr_2O_7^{2-}$) ranging from 10 to 100 mg./liter, and feed ratios of surfactant to dichromate from 2.5 to 6.0 (on a weight basis) are employed. The effects of feed concentration of dichromate, of feed concentration of surfactant, of liquid detention time, of air rate, and of feed into and above the foam phase are established. Results are presented in terms of the effluent concentrations of dichromate and surfactant, the ratios of effluent flow rate to feed flow rate, removal ratios, and enrichment ratios. For three-column series operation, predictions show that a feed stream containing 100 mg./liter dichromate may be separated (with approximately 400 mg./liter of surfactant) into an effluent stream containing 8 mg./liter of dichromate and a collapsed foam stream containing 468 mg./liter of dichromate.

Ion flotation involves the addition to aqueous solution of a surface-active agent that has a hydrophobic part (generally containing a long hydrocarbon chain) of opposite charge to the inorganic ion that is to be separated. An insoluble compound is formed between the hydrophobic cation or anion and the inorganic anion or cation of interest; the compound, which is surface-active, may then be floated to the surface of the solution and carried into a foam phase by aeration of the solution. The compound is preferentially adsorbed at the air-aqueous solution interfaces associated with the foam-generating bubbles. An extensive discussion of ion flotation has been presented by Sebba (8), and Rubin and Gaden have reviewed a number of applications (7).

The ion flotation of dichromate has been studied previously (3 to 5), but virtually all of the experiments were conducted on a batch basis, foaming a given volume of feed solution with the dependent variables changing as functions of time. The batch studies were carried out with feed solutions containing from 5 to 99 mg./liter of dichromate ion and from 202 to 396 mg./liter (added in several doses) of a cationic surfactant, ethylhexadecyldimethylammonium bromide (EHDA-Br). Most efficient operation was achieved with a molar feed ratio of EHDA ion to dichromate ion from 2.1 to 3.0, and removal ratios (defined as the quantity of dichromate floated per unit quantity of dichromate in the feed) ranging from 80 to 95% were obtained. For the ion flotation of dichromate to be a feasible technique for the treatment of industrial wastes, the process would have to be conducted with continuous flow ion flotation units. The results of the batch experiments were sufficiently promising to merit an experimental investigation of the continuous ion flotation of dichromate.

In the present study, continuous flow experiments include the effects of the dichromate concentration in the feed stream, of the surfactant concentration in the feed stream, of liquid detention time, and of air rate. De-