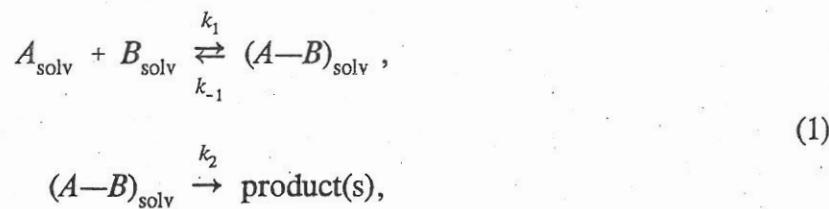


## Experiment 22

### The Kinetics of a Diffusion-Controlled Reaction

**Objective** To determine the rate constant and collision diameter of a diffusion-controlled reaction (photoexcited anthracene with carbon tetrabromide) using the technique of fluorescence quenching.

**Introduction** In this experiment you will study a very fast bimolecular reaction between two different species in solution. Let us assume that the reactants, *A* and *B*, which are electrically neutral, undergo independent, random motion in solution. There is a certain probability that *A* and *B* will encounter each other at some close distance, *R*, where *R* is approximately equal to the sum of their molecular radii. This arrangement is called an *encounter complex*. Because there is a tendency for the solutes *A* and *B* to maintain constant random motion, it is inevitable that, once loosely held in this complex, they will subsequently separate from each other unless a chemical reaction (or other definitive process) first links them together or causes them to react. Since this random motion takes place in a "bath" of solvent molecules (assumed to be unreactive with respect to *A* and *B*), the separation of the *A*—*B* encounter complex will be impeded by the neighboring solvent molecules. This artificial holding together of the two molecules is called the *cage effect*. The important point is that there is a kinetic competition between a net (thus measurable) reaction between *A* and *B* via the collision complex, and the release of *A* and *B* from the solvent cage back into the solvent medium where no reaction occurs. We can represent these processes by the following kinetic scheme:



where  $A_{\text{solv}}$  and  $B_{\text{solv}}$  represent the solvated *A* and *B* molecules, and  $(A-B)_{\text{solv}}$  is the solvated encounter complex. The rate constants for these elementary steps are denoted as  $k_1$ , for the bimolecular formation of the encounter complex (or diffusion into the solvent cage);  $k_{-1}$ , for the unimolecular dissociation of the complex (or diffusion out of the solvent cage); and  $k_2$  for the "unimolecular" reaction between *A* and *B* in the complex to form the product(s).

Because we assume that the encounter complex undergoes rapid deactivation, either by dissociation or via reaction, we can employ the steady-state approximation, according to which the net formation rate of  $(A-B)_{\text{solv}}$  is zero. Thus

$$\frac{d[A-B]}{dt} = k_1[A][B] - (k_{-1} + k_2)[A-B] = 0,$$

and we can approximate the steady-state concentration of the encounter complex as

$$[A-B] \simeq \frac{k_1[A][B]}{k_{-1} + k_2}$$

Now, if we further assume that the reaction rate is given by the slow step, i.e.,  $k_2[A-B]$ , we can express that rate as

$$\text{Rate} \simeq k_1 k_2 [A-B] = \frac{k_1 k_2}{k_{-1} + k_2} [A][B]$$

For convenience, we define the second-order rate coefficient,  $k_{\text{obs}}$ , as

$$k_{\text{obs}} \equiv \frac{k_1 k_2}{k_{-1} + k_2} \quad (2)$$

This analysis leads to two limiting cases with respect to  $k_{\text{obs}}$ : (1) The reaction between  $A$  and  $B$  is very slow compared with their departure (and separation) from the solvent cage, i.e.,  $k_{-1} \gg k_2$ ; and (2) the  $A-B$  reaction is much faster than their separation, or  $k_2 \gg k_{-1}$ . The first case describes a reaction that is under *chemical* control, with  $k_{\text{obs}} \simeq k_1 k_2 / k_{-1}$ , and the second pertains to a *diffusion-controlled* reaction for which  $k_{\text{obs}} \simeq k_1$ . The latter situation is considered in this experiment. More details of the reaction, which involves fluorescence quenching, will be described after a discussion of the relevant theoretical ideas.

### Diffusional Mass Transport

The basic issue in a diffusion-controlled reaction concerns the dynamics of mass transport in the condensed phase. The fundamental equations describing mass transport are embodied by Fick's laws of diffusion. These laws are also encountered and discussed in Experiment 19. Fick's first law says that the number of molecules,  $dn$ , diffusing through a unit area,  $A$ , per unit time,  $dt$ , in the  $x$ -direction (i.e., the flux,  $J_x$ ) is proportional to the concentration gradient at that point, i.e.,

$$J_x \equiv \frac{1}{A} \frac{dn}{dt} = -D \frac{dC}{dx}, \quad (3)$$

where the proportionality constant  $D$  is called the *diffusion coefficient*. The cgs units of  $D$  are  $\text{cm}^2 \text{s}^{-1}$ . The minus sign in equation (3) indicates that transport goes *against* the concentration gradient (i.e., from high to low concentration values).

Fick's second law states that the change in the concentration of molecules,  $dC$ , diffusing across an infinitesimally thin plane per unit time,  $dt$ , is proportional to the gradient of the flux:

$$\left( \frac{dC}{dt} \right)_x = - \left( \frac{dJ_x}{dx} \right)_t \quad (4)$$

Again, the minus sign ensures that concentration increases in time in response to a flux that decreases with increasing  $x$ . If we assume  $D$  to be independent of  $x$ , we may substitute the expression for  $J_x$  from Fick's first law into the second to give

$$\left( \frac{dC}{dt} \right)_x = D \left( \frac{d^2C}{dx^2} \right)_t \quad (5)$$

Because we are concerned with a three-dimensional and isotropic space, i.e., equal forces in all directions, we can write Fick's second law in more general form as

$$\left( \frac{dC}{dt} \right)_{x,y,z} = D (\nabla^2 C)_t, \quad (6)$$

where  $\nabla^2$  is the Laplacian operator  $\partial^2/dx^2 + \partial^2/dy^2 + \partial^2/dz^2$ .

The solution of Fick's second law is an equation that expresses concentration as a function of *space*, i.e., distance,  $r$ , and *time*,  $t$ . We are interested in solving equation (6) with the boundary conditions

$$\begin{aligned} C(r, 0) &= C_0, \\ C(\infty, t) &= C_0, \\ C(r = R, t) &= 0, \end{aligned}$$

where  $r = [x^2 + y^2 + z^2]^{1/2}$ , and  $C_0$  denotes the *bulk* concentration.

The first boundary condition states that initially the bulk concentration prevails throughout the system; the second says that even after the reaction starts ( $t > 0$ ), the concentration very far away from the reactant is constant, i.e., the bulk concentration,  $C_0$ ; and the third indicates that the reactant concentration is zero at a distance equal to the sum of the reactant collision radii. Note that the diffusion coefficient,  $D$ , contained in equations (3), (5), and (6) is the *sum* of the individual diffusion coefficients of the reactants,  $D_A + D_B$ . This expression essentially allows the motion of one of the reactants to be considered *relative* to the other.

The solution of Fick's second law for this set of boundary conditions (first rendered by Smoluchowski in 1917) yields, for the space-time dependence of  $C$ ,

$$C(r, t) = C_0 \left\{ 1 - \frac{R}{r} \operatorname{erfc} \left[ \frac{(r-R)}{2} (Dt)^{1/2} \right] \right\}, \quad (7)$$

in which  $\operatorname{erfc}$  is the *co-error function*,

$$\operatorname{erfc}(x) = 1 - \frac{2}{\pi} \int_0^x \exp(-y^2) dy$$

The time dependence of reactant at the reaction boundary  $r = R$  becomes expressible in terms of the flux of reactant,  $J_R$ , or its rate of transport across a hypothetical spherical surface with radius  $R$ . Thus

$$J_R = \frac{4\pi RDN_A C_0}{1000} \left[ 1 + \frac{R}{(\pi Dt)^{1/2}} \right] \text{ molecules s}^{-1}, \quad (8)$$

where  $N_A$  is Avogadro's number. Equation (8) indicates that the flux (molecules per unit area per second) is actually time-dependent. This time dependence comes from  $R/(\pi Dt)^{1/2}$ , which is sometimes referred to as the *transient term*. Physically, the transient term accounts for the fact that initially nearby reactant molecules do not have to diffuse through the bulk medium in order to react. After a short time, however, these nearby reactant molecules are depleted, and the flux approaches a constant, or steady-state, value,  $4\pi RDN_A C_0/1000$ . This can be seen mathematically in that time is represented in equation (8) as  $t^{-1/2}$ .

Because  $J_R$  represents the rate of passage of one of the reactants (having a bulk concentration  $C_0$  in  $\text{mol dm}^{-3}$ ) through a spherical reaction surface with the other reactant at the center, the rate of the reaction represented by equation (1) is merely  $J_R$  itself. Hence, the bimolecular, diffusion-controlled rate constant  $k_1$  is  $J_R/C_0$ ,

$$k_1 = \frac{4\pi RD}{1000} \left[ 1 + \frac{R}{(\pi Dt)^{1/2}} \right] \text{ (dm}^3 \text{ mol}^{-1} \text{ s}^{-1}). \quad (9)$$

The units shown pertain if  $R$  and  $D$  are expressed in  $\text{cm}$  and  $\text{cm}^2 \text{ s}^{-1}$ , respectively. This rate constant is not a true *constant*, however, because of the transient term. However, as  $t$  becomes large enough,

$$k_1 \rightarrow \frac{4\pi RD}{1000} \text{ (dm}^3 \text{ mol}^{-1} \text{ s}^{-1}), \quad (10)$$

and equation (10) can be used to determine the diffusion-controlled rate constant if the mutual diffusion constant  $(D_A + D_B)$  and collision radii  $(R_A + R_B)$  are known. We emphasize again that the units of  $k_1$  presented above are  $\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ , if  $R$  and  $D$  are expressed in  $\text{cm}$  and  $\text{cm}^2 \text{ s}^{-1}$ , respectively.

Often, however,  $R$  and  $D$  are not known, and an indirect method is used to estimate  $k_1$ . This approach, developed by Einstein using Stokes's law (see Experiment 17), allows the diffusion coefficient to be expressed in terms of a *bulk property*, the solvent viscosity,  $\eta$ . Thus

$$k \approx \frac{8RT}{3000\eta} \quad (\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1}), \quad (11)$$

where  $R$  is the gas constant. Equation (11), which is often referred to as the Stokes-Einstein-Smoluchowski (SES) equation, holds when the reactants  $A$  and  $B$  are different and their sizes are larger than that of the solvent molecules. The indicated units are obtained if  $R$  and  $\eta$  are expressed in cgs units, i.e.,  $8.314 \times 10^7 \text{ erg mol}^{-1} \text{ K}^{-1}$  and poise, respectively.

### Fluorescence Quenching

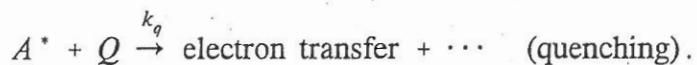
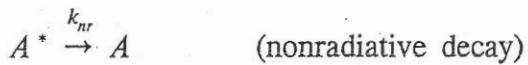
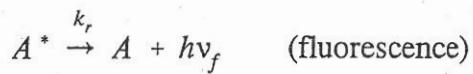
The reaction studied in this experiment takes place between an electronically excited molecule and a quencher, a species that removes the electronic excitation. The advantages of using an electronically excited molecule as one of the reactants are: (1) the reaction takes place only when the system is exposed to light; (2) the concentration of electronically excited molecules can be readily followed fluorometrically (via fluorescence detection); and (3) the reaction itself is intrinsically fast: It has no activation energy per se, and is limited by the encounter between the excited molecule and the quencher. Experiments 27, 29, 31 and 35 also utilize luminescence quenching as a "kinetic" technique. The overall reaction describing fluorescence quenching can be represented as



where  $A^*$  represents the electronically excited (fluorescent) state of a molecule (e.g., anthracene in this experiment),  $Q$  denotes the quencher (e.g.,  $\text{CBr}_4$ ) that "extinguishes" the  $A^*$  fluorescence, and "product(s)" indicates the species into which the fluorescent molecule is eventually transformed. This deexcitation process may occur through a charge-transfer intermediate ( $A^+ - Q^-$ ) in which electronic charge is transferred to the quencher molecule from the excited-state species. For the system encountered in this experiment, electron transfer from  $A^*$  to the quencher occurs very rapidly. In general, it is possible that the consequence of fluorescence quenching is the return of  $A$  to its electronic ground state without any change in structure, i.e.,  $A + Q$ .

In the case of  $\text{CBr}_4$ , the quenching mechanism involves the electron transfer from the excited state,  $A^*$ , to  $\text{CBr}_4$ , which is more electronegative than  $A^*$ . A reverse electron transfer then restores the two molecules to their neutral species,  $A$  being in the ground state. (See Experiment 27 for another example of quenching via electron transfer.)

The bimolecular quenching reaction shown above, however, competes with the intrinsic first-order fluorescence decay of  $A^*$  and thus reduces the probability that the  $A^*$  species will emit a photon. A more complete scheme is thus



If a solution containing  $A$  and  $Q$  is irradiated with a *steady-state* light source (at a wavelength where  $A$  absorbs), a very small time-independent concentration of  $A^*$  is produced. The probability that  $A^*$  will fluoresce,  $P_f$ , is the ratio of the fluorescence (or radiative) rate constant to the *sum* of *all* the rate constants that deplete  $A^*$ , namely,

$$P_f = \frac{k_r}{k_r + k_{nr} + k_q[Q]}, \quad (12)$$

where  $k_r$ ,  $k_{nr}$ , and  $k_q[Q]$  are, respectively, the radiative, nonradiative, and (pseudo-first-order) quenching rate constants. Since  $[Q] \gg [A^*]$ , the diffusion-controlled bimolecular quenching step becomes pseudo first order. If the amount of light absorbed by  $A$  is unaffected by  $[Q]$ , the fluorescence *intensity*,  $I$ , measured by the detector (i.e., the instrument response) is proportional to  $P_f$ . In the absence of quencher,  $I = I_0 = k_r/(k_r + k_{nr})$ , and the ratio of fluorescence intensities of unquenched to quenched samples of  $A$  becomes

$$\frac{I_0}{I} = 1 + \frac{k_q[Q]}{k_r + k_{nr}}. \quad (13)$$

Identifying  $1/(k_r + k_{nr})$  as the reciprocal of the *fluorescence lifetime* of  $A^*$  in the absence of quencher ( $\tau_0$ ), we can write equation (13) as

$$\frac{I_0}{I} - 1 = k_q \tau_0 [Q]. \quad (14)$$

Equation (14) is known as the steady-state *Stern-Volmer* (S-V) relation as applied to fluorescence quenching. It predicts that a plot of  $(I_0/I - 1)$  vs.  $[Q]$  should be linear, having an intercept of zero and a slope equal to  $k_q\tau_0$ . (See also Experiments 27, 29, and 35.) The latter is called the Stern-Volmer constant,  $K_{SV}$  (units of  $\text{dm}^3 \text{ mol}^{-1}$ ). If fluorescence quenching is diffusion-controlled (which is the case in the anthracene/CBr<sub>4</sub> system),  $k_q$  can be identified with  $k_1$  [see equations (1), (2) and (9-11)].

In the absence of the transient effect, which is a type of "static" quenching in that some  $A^*$  molecules are quenched by nearby  $Q$  molecules that do not have to undergo transport fully through the solvent medium in order to approach  $A^*$  within a distance  $R$ , the S-V plot should be linear, and independent knowledge of  $\tau_0$  allows us to determine  $k_q$ . Alternatively, if the process is known to be diffusion-controlled, and if  $k_1$  is determined from the SES (or some other) equation [e.g., equation (10)], the fluorescence lifetime of  $A^*$  can be estimated from the measured  $K_{SV}$  value. This is a case in which a dynamic property (a rate constant) can be obtained from a static experiment ( $I_0/I$  measurements).

The S-V relation can be expressed in terms of the diffusion-controlled rate constant that contains the transient term, e.g., equation (9). The time dependence was treated by Ware and Novros, who applied equation (9) to the condition pertinent to this experiment, namely, that the system is considered to be suddenly exposed to a steady-state excitation source, called a step function (see Figure 1 and *Further Readings*). In this case, the S-V equation becomes more complicated:

$$\frac{I_0}{I} = \frac{1 + 4\pi RDN_A / \{1000 [Q]\tau_0\}}{Y}, \quad (15)$$

where

$$Y = 1 - \frac{b}{a^{1/2}} \exp\left(\frac{b^2}{a}\right) \operatorname{erfc}\left(\frac{b}{a^{1/2}}\right), \quad (16)$$

in which

$$a = \frac{1}{\tau_0} + \frac{4\pi RDN_A}{1000 [Q]},$$

$$b = \frac{4R^2(\pi D)^{1/2}N_A}{1000 [Q]}.$$

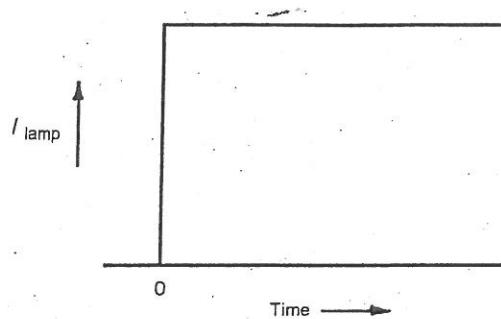


Figure 1. Step excitation function. At time = 0, absorbing radiation is suddenly "switched on."

If  $Y = 1$ , equation (15) reduces to equation (14), in which case  $k_q = k_1$  from equation (10). (Verify this!)

The Stern-Volmer plot based on equation (15) will be curved upward. If the mutual diffusion coefficient of the reacting pair is known, the S-V plot of experimental data can be fit using equation (15) with  $R$  used as a "fitting" parameter. For a satisfactory match between experiment and theory, however, the value of  $R$  must be physically reasonable. That is,  $R$  should be approximately equal to the sum of the molecular diameters of  $A$  (actually  $A^*$ ) and  $Q$ .

## Safety Precautions

- Always wear safety glasses or goggles; these glasses should block ultraviolet light. Ordinary plastic safety goggles or glasses may not be effective in absorbing all the ultraviolet radiation. Check with your instructor.
- Do not allow solid anthracene to come in contact with the skin. Wear gloves when handling anthracene and  $\text{CBr}_4$ .
- Ozone is sometimes produced by ultraviolet light sources. If you detect this gas, which has a sharp, slightly acrid odor, notify your instructor and immediately shut off the source. Increase ventilation, and leave the room immediately.
- The compressed-gas cylinder used in deaerating should be securely strapped to a firm support. The delivery pressure should never exceed a few (e.g., 5) psig.
- The experiment should be performed in an open, well-ventilated laboratory.

## Procedure

1. Using freshly sublimed or purified anthracene and  $\text{CBr}_4$ , prepare 250 mL of an  $\sim 1.00 \times 10^{-4}$  M solution of anthracene (AN) in spectrometric (or fluorometric) quality *n*-hexane. Using this solution as "solvent," prepare 25 mL of an  $\sim 1.50 \times 10^{-2}$  M "stock" solution of  $\text{CBr}_4$ .
2. Prepare at least eight dilutions of the AN/ $\text{CBr}_4$  stock solution using the AN solution as solvent. The most dilute should be 5%, and the rest should be nearly evenly spaced up to the stock-solution value. 10-mL volumes of each solution are appropriate. Include 0 and 100% samples for uniformity. Label these samples.

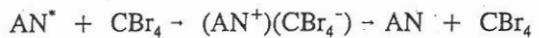
3. Starting with the 0% sample (AN only), introduce the solution into a fluorescence cell and deaerate with dry  $N_2$  for about 2 min. Avoid using an excessive flow rate that will cause the solution to splatter from the cell or to otherwise result in undue solvent evaporation. Promptly stopper the cell.
4. Record the full fluorescence spectrum using the instrumental conditions previously outlined by the instructor. Label the spectrum while it is being recorded (if a strip-chart recorder is used).
5. Using exactly the same excitation wavelength as in step 4, measure the fluorescence intensity of the most dilute and subsequently more concentrated AN/CBr<sub>4</sub> solutions. You must deaerate each sample in a consistent procedure (i.e., identical bubbling time). However, instead of obtaining the entire fluorescence spectrum, you need monitor only the AN maximum (near ~398 nm). Make sure you are at the maximum by first scanning that region slowly. If you have any doubt as to the steadiness of the exciting source, switch back to the AN sample and compare its maximum intensity with that of the first sample run.
6. As you examine increasingly more concentrated samples are examined, the fluorescence intensity will diminish to the point that you will need to increase the instrument gain (i.e. signal amplification) to ensure maximal reading sensitivity. Establish the background, or "dark check," when the instrument sensitivity is changed.
7. Keep in mind the following procedural points: (a) The temperature of the sample should be held constant to 0.5 to 1°C; (b) the exposure times of all of the samples (especially those rich in CBr<sub>4</sub>) to the excitation source (and even room light) should be minimized; and (c) samples should be stored in the dark until they are used.

### Data Analysis

1. Tabulate the data as  $I_f$  (AN fluorescence intensity), and  $[Q]$  (CBr<sub>4</sub> concentration), or enter the data into a spreadsheet.
2. Make a Stern-Volmer plot of the data.
3. Superimpose on this plot the straight line expected if the transient term is ignored. The following information can be used: Ware and Novros reported a value of  $D = 4.35 \times 10^{-5} \text{ cm}^2 \text{ s}^{-1}$  in *n*-heptane at 25°C. To obtain  $D$  for *n*-hexane at some other temperature, assume that  $D$  is proportional to  $T/\eta$ . They also reported a value of  $\tau_0 = 5.52 \times 10^{-9} \text{ s}$ . You can assume this to be temperature-independent.
4. Also indicate the straight line expected on the basis of the SES equation. Tabulate the two calculated  $k_1$  values and their error limits.
5. Using equation (15), calculate several points for the S-V plot using a fixed value of  $R$ . If you are not using a computer to generate these points, choose them judiciously so you can discern the general shape of the calculated plot. Start with a value of 6 Å and increase  $R$  in steps of 0.5 Å until you obtain satisfactory agreement with the data.

**Questions and Further Thoughts**

1. Why is the anthracene fluorescence quenching diffusion controlled, as opposed to chemical controlled? Thus justify why the reaction between  $\text{AN}^*$  and  $\text{CBr}_4$  is very fast.
2. What experiment(s) could you perform to determine whether a stable photoproduct is formed as a result of the fluorescence quenching?
3. What could you do to obtain supporting evidence for the existence of a charge-transfer (or ion-pair) intermediate, e.g.,



in the quenching process?

4. Verify that if  $Y = 1$  in equation (15),  $k_q \rightarrow k_1$ ; see equations (14) and (10).
5. One of the undesirable complications associated with this type of experiment is the potential presence of *ground-state* complexes between  $\text{AN}$  and  $\text{CBr}_4$ . If such a complex competes for light absorption with uncomplexed, free anthracene, it can cause a decrease in fluorescence intensity without there having to be a diffusion between  $\text{AN}^*$  and  $\text{CBr}_4$ . This is sometimes called *static quenching*. For this reason, it is desirable to work with quencher concentrations as low as possible. What experiment could you do to obtain evidence of a ground-state complex? How could you determine the equilibrium constant of such a complex?
6. What effect would a polar solvent, e.g., acetonitrile ( $\text{CH}_3\text{CN}$ ), have on the ground- and excited-state interactions between  $\text{AN}$  and  $\text{CBr}_4$ ?
7. If you wanted to do a fluorescence quenching experiment that demonstrated (and thus maximized) the transient effect in diffusional processes, what qualities of solvent viscosity and fluorescence probe lifetime would you seek (i.e., high or low viscosity; long or short fluorescence lifetime)?
8. The lowering of dissolved oxygen by bubbling the solution with dry  $\text{N}_2$  gas (deairating) is an application of Henry's law. Explain how deaeration works in this context.

**Further Readings**

- R. A. Albery and R. J. Silbey, *Physical Chemistry*, 2d ed., pp. 692-696, Wiley (New York), 1997.  
 A. H. Alwattar, M. D. Lumb, and J. B. Birks, *Organic Molecular Photophysics*, vol. 1, p. 403, Wiley (New York), 1973.  
 P. W. Atkins, *Physical Chemistry*, 5th ed., p. 795, W. H. Freeman (New York), 1994.  
 I. N. Levine, *Physical Chemistry*, 4th ed., p. 469, McGraw-Hill (New York), 1995.  
 J. H. Noggle, *Physical Chemistry*, 3d ed., p. 480, HarperCollins (New York), 1996.  
 J. B. Birks, *Photophysics of Aromatic Molecules*, p. 518, Wiley (New York), 1970.  
 R. M. Noyes, *Progress in Reaction Kinetics*, vol. 1, 131 (1961).  
 N. J. Turro, *Modern Molecular Photochemistry*, p. 311, Benjamin-Cummings (Menlo Park, Calif.), 1978.  
 W. R. Ware and J. S. Novros, *J. Phys. Chem.*, 70:3246 (1966).

## Experiment 22

## Data Sheet A

### Diffusion-Controlled Rate Constant

NAME \_\_\_\_\_

DATE \_\_\_\_\_

Anthracene concentration: \_\_\_\_\_ g in \_\_\_\_\_ mL

$\text{CBr}_4$  stock solution: \_\_\_\_\_ g added

AN/ $\text{CBr}_4$  Solutions  
% stock

Fluorescence Intensity

Temperature: \_\_\_\_\_

0%

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

## Experiment 22

*Data Sheet B*

### Diffusion-Controlled Rate Constant

NAME \_\_\_\_\_

DATE \_\_\_\_\_

Anthracene concentration: \_\_\_\_\_ g in \_\_\_\_\_ mL

$\text{CBr}_4$  stock solution: \_\_\_\_\_ g added

AN/ $\text{CBr}_4$  Solutions  
% stock

Fluorescence Intensity

Temperature: \_\_\_\_\_

0%

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_

\_\_\_\_\_