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Critical Review of Hydrolysis of Organic Compounds in Water Under Environmental Conditions

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This review examines the rate constants for hydrolysis in water of 12 classes of organic compounds with the objective of using these data to estimate the presistence of these compounds in freshwater aquatic systems. Primary data were obtained by literature review through most of 1975 and some of 1976. These data, which include values for acid, base, and water promoted rate constants (k_A, k_B, k_N) and temperature coefficients are presented in 18 tables in section 4. Estimated rate constants for hydrolysis under environmental conditions are presented in 13 tables in section 5, including rate constants at 298 K and pH 7 for acid, base, and water promoted reactions together with values for the estimated rate constant (k_h) and the half-life ($t_{1/2}$).

Key words: Acid; base; environmental conditions; freshwater systems; hydrolysis; organic compounds; rate constants.

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1. Introduction

1.1. Background

During 1975, 293 billion pounds of about 8,000 different organic compounds were manufactured in the United States, exclusive of petroleum products [1].[1] Each year about 5,000 new compounds are synthesized, and 200 to 300 of them come into commercial use [2]. Many millions of pounds of synthetic organic compounds are deliberately introduced into the environment for weed and pest control, and many millions of pounds of other chemicals are introduced into the environmental through dispersive uses.

The effects on the biosphere of these often persistent, often biologically active synthetic chemicals are only now becoming fully apparent: fish kills, species extinction or enervation, and perhaps as many as several hundred thousand human cancers per year have been associated with the increasing production and dispersion of synthetic chemicals [3]. In an effort to avert further, and perhaps tragic, biological consequences of a continuing and unrestricted discharge of chemicals to the air, water, and soil, the government began some years ago to regulate the manufacture and use of selected compounds, such as DDT, which had been shown to be exceptionally hazardous to the biological environment. The Toxic Substances Control Act of 1976 is one further step toward regulating the introduction of hazardous chemicals into the environment.

Fortunately, the environment has the capacity to cleanse itself of many kinds of chemicals through a variety of chemical and biological processes. An understanding of this process, especially how rapidly the environment can degrade a specific chemical structure to a simpler and potentially less harmful chemical, is a key to the rational use of natural resources with minimal abuse. The ability to predict the probable fate of specific compounds in the environment is also essential for screening the thousands of chemicals that may be considered for applications leading to their ultimate introduction into water, soil, or air.

The objective of this review is to provide one kind of predictor of environmental fate—hydrolysis of organic compounds in freshwater systems—based on the best available kinetic data for hydrolysis in water.

There are at least two reasons why hydrolysis may be a significant chemical process in the environment. First, many hydrolyzable chemicals, including pesticides and plasticizers, eventually find their way into groundwater, streams, and rivers through leaching and runoff. Second, rates of hydrolysis in aquatic systems are independent of commonly used but rapidly changeable indicators of the degradative capacity of aquatic systems, such as sunlight, microbial populations, and oxygen supply; rates do depend on pH, temperature, and concentration of chemical—properties that may change only slowly and seasonally.

In undertaking a review of this kind, which claims to predict rates of hydrolysis in aquatic (freshwater) systems on the basis of laboratory data, we are aware of the belief that such estimates are of little value because the rates of hydrolysis obtained in laboratory studies do not reflect the complexity found in the environment. This concern, although understandable, is not well founded. There are several examples where rates of hydrolysis have been measured in both pure and natural waters and which showed good agreement between the two kinds of measurements for a variety of chemical structures [4,5,6] providing that both pH and temperature were measured.

Another objection raised occasionally concerns the validity of the extrapolation of data for hydrolyses measured in the laboratory at concentrations of chemicals that often exceed 0.01 M to environmental conditions where typical concentrations of trace organics rarely exceed 10^{-6} M. The apparent implication of this concern is that at high dilution other complications may arise but in fact it is axiomatic that rate processes found to be simple at high concentrations remain so at low concentrations; moreover, with only one exception, the rate laws for hydrolysis reviewed here show a simple first-order dependence on the chemical. Thus the actual half-life of the chemical is independent of its concentration and depends only on easily and accurately measurable parameters like pH and temperature.

For these reasons we believe that in most cases extrapolation from laboratory to field site is relatively uncomplicated and that estimates of persistence summarized in section 5 provide valuable information on the upper limit for persistence of hydrolyzable dissolved organic compounds. From these estimates it should be possible to also assess possible effects of intervention by other environmental processes. Thus, if field measurements of the half-lives (persistence) of

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[1] Figures in brackets indicate literature references at the end of each section or subsection.

specific chemicals depart significantly (by factors of 5 or more) from those predicted in section 5 after taking into account differences in pH and temperature, then some other process such as biodegradation, or photolysis or insolubilization may have changed the measured half-life. These estimates apply only to chemicals dissolved in water; in most cases any suspended or oil solublized chemicals will hydrolyse much more slowly than predicted.

Only recently has there been much interest in measuring or estimating the rates of hydrolysis of organic compounds under environmental conditions which include very dilute solute concentrations in pure or natural waters in the pH region 4 to 8 at or below 298 K. In the absence of much environmentally-"intended" data, this review has drawn upon a large amount of hydrolysis data from studies concerned with kinetics and mechanism. Some of these data were obtained from experiments in mixed solvents, or at high temperatures (to 393 K) or extreme pH values, and often lack data on temperature dependence.

In some cases data are given for compounds which are not of environmental concern; however these data do describe the range of rate constants for the class of compounds and make it possible to estimate the rate constant(s) for and persistence toward hydrolysis of organic compounds not included in this review.

We hope this review will stimulate more and better studies of hydrolysis of synthetic chemicals in water under carefully defined conditions, both in the laboratory and in the field, to supplement the data presented here and further validate the concept that careful laboratory studies on individual processes can accurately predict the fate and persistence of chemicals in the environment.

1.2. Content of This Review

This review provides information on kinetics of hydrolysis in water of 13 classes or organic structures, with sufficient detail within each class to enable the reader to find either the specific compound or one close enough to estimate reliably the rate constant(s) for hydrolysis. The classes of compounds reviewed are:

(1) Organic halides
  (a) Alkyl halides
  (b) Allyl halides
  (c) Benzyl halides
  (d) Polyhalomethanes
(2) Epoxides
(3) Esters
  (a) Aliphatic acids
  (b) Aromatic acids
(4) Amides
(5) Carbamates
(6) Phosphorus esters
  (a) Phosphonates
  (b) Phosphates and thiophosphates
  (c) Phosphonohalidates
(7) Acylating and alkylating agents and pesticides

Kinetic data are presented in two sections. Section 4, "Hydrolysis Rate Data," summarizes primary kinetic data on hydrolysis of these compounds in water or (in a few cases) in mixed solvents; section 5, "Estimated Hydrolysis Rates under Environmental Conditions," uses data from section 4 to estimate rate constants and half-lives for these same compounds at 298 K, pH 7, and zero ionic strength, conditions typical of a great majority of freshwater systems (section 2.1).

The classes of compounds reviewed here cover a significant fraction of the hydrolyzable structures found in organic molecules. Many of the individual compounds listed in this review are used in quantity in industrial or agricultural applications; others are unique to laboratory studies. Most of them have been subject to study in water solvent under conditions that require few assumptions and relatively short extrapolations to environmental conditions.

Several classes of hydrolyzable compounds have *not* been included in this review because some are hydrolytically inert under ordinary conditions and are best considered as refractory toward water (nitriles, vinyl, and aromatic chlorides are examples), and others are hydrolytically unstable but have not been examined in any quantitatively useful way.

1.3. Error Analysis

Because many different kinds of experimental procedures are used to measure hydrolysis reactions, no one error analysis procedure is applicable to all sets of data. Rate constants for hydrolysis of most compounds appear to be of high precision, often with less than 2% standard deviation. Different investigators have reproduced individual rate constants to within ±50%. Some sets of experimental data are reported with error limits that involve a judgmental factor in selection of data. These error limits may be considered equal to twice the standard deviation.

Most experimental measurements of E or $\Delta H^{\neq}$ are made over temperature spans of 40–80 K and usually around 345 K. Benson's "rule" [1] indicates that with a random error of 2% in rate constant k, activation energy E may be determined with an accuracy of about 5%; however, with a random error of 10% in k, E is only accurate to 100% or a factor of 2. Put another way, if E is known only with an accuracy of ±10%, k is known only with an accuracy of ±60%.

Values for k_h estimated in section 5 at 298 K are probably not more accurate than a factor of 2 (±100%) or less accurate than a factor of 5 (±250%) owing to uncertainties in pH, temperature coefficients, and, in some cases, solvent effects.

1.4. Literature Sources

A thorough search of the current literature through 1975 was made using *Chemical Abstracts.* Data were searched under the major subject headings for specific compounds.

1.5. Format

The review is divided into five sections; to assist the user, references are renumbered as appropriate within each section. Although every effort has been made to use a consistent format throughout, some differences among tables are unavoidable owing to differences in the kinds of compounds and the reliability of data available for various compounds.

Section 4 reports primary kinetic data in the form of rate constants for acid (k_A), neutral (k_N), and base catalyzed (k_B) hydrolysis reactions, together with available temperature coefficients. Data are summarized in 18 tables. Section 5 uses data from section 4 to estimate k_h at 298 K and pH 7 for selected compounds in 13 tables.

In each table, compounds are grouped by class, such as alkyl halides, epoxides, and esters, and are listed in order of increasing complexity. Line formulas are used in most cases to avoid ambiguities of nomenclature.

Rate constants k are expressed in units of s^{-1} or $M^{-1}s^{-1}$, where $M=mol \cdot dm^{-3}$. Very large and very small values of k are listed in two ways: a column heading $10^3\,k$ requires that every listing in that column be multiplied by 10^{-3} to retrieve k; a column listing of 5(−11) means 5×10^{-11}. Throughout this review, we have used the SI unit of joules for energy and entropy; conversion of joules to calories requires division by 4.184. Units for A are the same as for k; units for $\Delta S^{\ddagger}$ are J/mol; units for E and $\Delta H^{\ddagger}$ are kJ/mol. To simplify presentation of data in the tables, we have fixed the value of R at 0.01914 kJ/mol K, which includes a conversion factor of 2.303 for base 10 log units.

1.6. Acknowledgments

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2. Physical and Chemical Properties of Freshwater Aquatic Systems

The freshwater systems of North America's streams, rivers, lakes, and groundwater comprise several hundred thousand square miles of area draining 7 million square miles of diverse environmental regions. Since one objective of this review is to provide reliable estimates of lifetimes for organic compounds in freshwater aquatic systems, some review of the range of physical and chemical properties of freshwater is desirable.

2.1. Characteristics of U.S. Freshwater Streams and Rivers

Detailed information on most freshwater systems in the United States, including temperature range, pH, and mineral content, is available from the United States Geological Survey [1]. These data have been summarized for 111 streams and rivers in the United States that account for over 95% of the water volume; table 2.1 summarizes mean concentration values for 11 inorganic constituents found in natural waters at pH 7.5 and 287 K.

Table 2.1 Analysis for U.S. rivers and streams

(mean values in mg/ℓ)

Ca^{2+}	Mg^{2+}	Na^{+}	K^{+}	Fe^{2+}	HCO_3^-	SO_4^{2-}	Cl^-	NO_3^-	PO_4^{3-}	CO_3^{2-}
36	9	25	2	0.06	127	36	7.7	1.4	0.65	0.10

pH 7.5, Temperature 287K

The following values represent average conditions in most freshwater systems in nonwinter months: temperature 293 K, pH 7.0, and ionic strength 0.00.

Review of the hydrolysis literature for organic chemicals shows that many rate constants have been evaluated at 298 K or higher; only a few data have been obtained at lower temperatures. Although higher than either the mean or the average temperature of the rivers and streams, we have chosen 298 K as the environmental reference temperature for this review in order to use as much primary kinetic data as possible for estimates of persistence without temperature extrapolations.

Estimates of persistence (half-lives) summarized in Section 5, will be 50 percent longer at 293 K and 130 percent longer at 287 K, based on an average energy of activation for hydrolysis of 65 kJ/mol.

Reference for Section 2

[1] Quality of Surface Waters of the United States, Parts 2-16, U.S. Geological Survey Water Supply Papers 2092-2099, 1972-73.

3. Hydrolysis Kinetics

3.1. Rate Laws

Hydrolysis refers to reaction of a compound in water with net exchange of some group X with OH at the reaction center:

$$RX+H_2O\rightarrow ROH+HX.$$

The detailed mechanism may involve a protonated or anionic intermediate or a carbonium ion, or any combination of these intermediates. But whatever the mechanism, the rate law for hydrolysis of substrate RX usually can be put in the form

$$-\frac{d(RX)}{dt}=k_h[RX]=k_B[OH^-][RX]+k_A[H^+][RX] +k_N'[H_2O][RX], \quad (1)$$

where k_B, k_A, and k_N' are the second-order rate constants for acid and base catalyzed and neutral processes, respectively.

In water $k_N'[H_2O]$ is a constant (k_N). The pseudo-first-order rate constant, k_h, is the observed or estimated rate constant for hydrolysis at constant pH. Equation (1) assumes that the individual rate processes for the acid, base, and neutral hydrolyses are each first order in substrate. With only a few exceptions, this is the case, and

$$k_h = k_B[OH^-] + k_A[H^+] + k_N. \qquad (2)$$

From the autoprotolysis water equilibrium, eq (3), eq (2) may be rewritten as eq (4).

$$[H^+][OH^-] = K_w \qquad (3)$$

$$k_h = \frac{k_B K_w}{[H^+]} + k_A[H^+] + k_N \qquad (4)$$

3.2. Effect of pH

From eq (4) it is evident how pH affects the overall rate: at high or low pH (high OH^- or H^+) one of the first two terms is usually dominant, while at pH 7 the last term can often be most important. However, the detailed relationship of pH and rate depends on the specific values of k_B, k_A, and k_N. At any fixed pH, the overall rate process is pseudo first order, and the half-life of the substrate is independent of its concentration:

$$t_{1/2} = 0.693/k_h. \qquad (5)$$

Equation (4) is conveniently expressed graphically as three equations—one each for the acid, base, and neutral hydrolysis reactions—in which log k_h is plotted against pH. The curves obtained are especially useful for estimating the effect of acid or base on the rate of hydrolysis. Figure 1 depicts a typical log k_h vs pH plot for compounds which undergo acid, water, and base promoted hydrolysis. It is obvious from eq (4) that the upper curve in figure 1 is a composite of three straight lines—(a) $\log k_h = \log (k_B K_w) + \text{pH}$; (b) $\log k_h = \log k_N$, and (c) $\log k_h = \log k_A - \text{pH}$—with slopes +1, 0, and −1, respectively. The lower curve results when $k_N = 0$.

Most log k_h vs pH curves are found to have one or two intercepts corresponding to pH values where two kinds of rate processes contribute equally to the overall process. Thus in figure 1 the intercept I_{AN} corresponds to a value of pH where $k_A[H^+] = k_N$; similarly, I_{NB} corresponds to $k_B[OH^-] = k_N$. In cases where k_A, k_B, or $k_N = 0$, only one intercept is observed. Values of pH corresponding to I may be calculated readily from the values of k_A, k_N and k_B:

$$I_{AN} = -\log (k_N/k_A) \qquad (6)$$

$$I_{NB} = -\log (k_B K_w/k_N) \qquad (7)$$

$$I_{AB} = -[\log (k_B K_w/k_A)]/2. \qquad (8)$$

Use of intercepts in tabulating data on rates of hydrolysis as a function of pH greatly simplifies the task of estimating the effect of pH on the rate constant k_h for a specific compound. For example, I_{NB} for alkyl halides lie above pH 11. Obviously, the base-catalyzed process for hydrolysis of alkyl halides is never of concern in estimating persistence in aquatic systems. Values of I_{AN}, I_{NB}, and I_{AB} are tabulated for specific compounds in section 5.

3.3. Effect of Temperature

Section 5 reports estimated half-lives at 298 K based on rate data listed in section 4 at several different temperatures and with temperature coefficients for the rate constants.

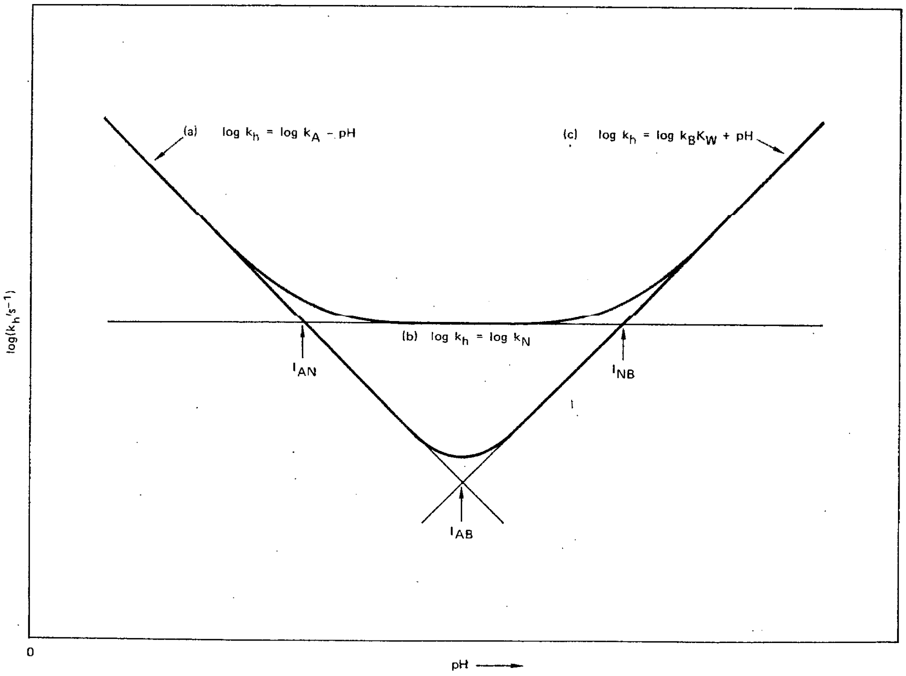

FIGURE 1. pH dependence of k_h for hydrolysis by acid, water, and base promoted processes.

The effect of temperature on the rate constant for a specific hydrolysis process can be expressed in several ways, all of which are variants of the general relation

$$\log (k/\mathrm{s}^{-1}) = -A/T + B \log T + C. \qquad (9)$$

There is no uniform practice for expressing values of *A*, *B*, or *C*; different investigators have used different versions of eq (9), usually in more familiar Arrhenius or absolute rate theory format

$$\log (k/\mathrm{s}^{-1}) = \log A - E/RT \qquad \text{Arrhenius} \quad (10)$$

$$\log (k/\mathrm{s}^{-1}) = \log (kT_m/h) - \Delta H^{\ddagger}/RT + \Delta S^{\ddagger}/R \qquad \text{Absolute Rate Theory} \quad (11)$$

Equation (9) can also be used to describe more accurately the effect of temperature on the rate constant and activation parameters by inclusion of a heat capacity contribution in the constants *B* and *C*. When the heat capacity term, $\Delta C_p^{\ddagger}$, is known or can be estimated, $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$ are then calculable at a specific temperature, T_2, from $\Delta H_o^{\ddagger}$ and $\Delta S_o^{\ddagger}$ or $\Delta H_{T_1}^{\ddagger}$ and $\Delta S_{T_1}^{\ddagger}$

$$\Delta H_{T_2}^{\ddagger} = \Delta H_o^{\ddagger} + \Delta C_p^{\ddagger} T_2 \text{ or } \Delta H_{T_2}^{\ddagger} = \Delta H_{T_1}^{\ddagger} - \Delta C_p (T_1 - T_2)$$

$$\Delta S_{T_2}^{\ddagger} = \Delta S_o^{\ddagger} + 2.303 \ \Delta C_p^{\ddagger} \ (\log \ T_2)$$

$$\text{or } \Delta S_{T_2} = \Delta S_{T_1} + 2.303 \ \Delta C_p \log \ (T_2/T_1)$$

A few authors have expressed rate parameters for a reaction in the mixed terms of *E* and $\Delta S^{\ddagger}$. Although such a practice is not immediately useful within the individual Arrhenius or absolute rate theory treatments, the rationale appears to be that $\Delta S^{\ddagger}$ and *E* are more easily conceptualized for diagnosis of the reaction mechanism.

Recasting of data to fit a uniform temperature dependent relationship might assist in mechanistic interpretations; however, the probable loss of precision for some data argues against such a practice. Therefore this review uses the original temperature dependence relationships developed for each compound to extrapolate rate constants to the chosen environmental temperature of 298 K.

Each table of data in section 4 lists the appropriate temperature coefficients together with values of coefficients in appropriate units. For a few compounds where no temperature coefficients were determined and the reported data are not at 298 K, we have extrapolated to 298 K using a stated temperature coefficient, based on a similar reaction with known temperature dependence relationships. In a very few examples where data are reported within ±5 K of 298 K, we have used the data without further correction.

3.4. Ionic Strength and Buffer Effects on Hydrolysis

The ionic strength of most natural freshwater is quite low, usually much less than 0.01 M in total cation and anion concentrations (see table 2.1). This is fortunate since ionic strength effects on hydrolysis reactions are difficult to predict: they can lead either to rate acceleration or retardation, depending on the substrate, the specific salts, and their concentration. Salt effects of this kind are associated with changes in activity coefficients of ionic or polar species or transition states with significant charge separation; no bond making or breaking is involved in such interactions.

Table 3.1 which lists some effects of different non-nucleophilic salts on the hydrolysis of several kinds of compounds, shows that in most cases rather massive amounts of these salts cause only 30-40% changes in rate constant. Concentrations necessary to effect these rate changes are found only in ocean or brackish waters or in highly buffered laboratory experiments.

Table 3.1. Salt effects in hydrolysis reactions.

Compound	Salt, M	Solvent, T/K	k/k_o[a]	Reference
t-BuBr	LiBr, 0.1065	90/10 acetone-water, 323	1.44	1
γ-Butyrolactone	NaCl, 0.51	0.10M HCl, 298	1.08	2
	$NaClO_4$, 0.40	0.10M $HClO_4$, 298	0.95	
CH_3CO_2Et	LiCl, 2.00	0.15M $HClO_4$, 298	1.35	3
	$NaClO_4$, 2.00	0.15M $HClO_4$, 298	1.00	
CH_3CO_2Me	NaCl, 0.100	0.01M NaOH, 308	0.96	4

[a] k/k_o is the ratio of rates with (k) and without (k_o) salt.

Another effect of nucleophilic added salts is to accelerate the rate of hydrolysis by a general acid or general base promoted process (where H^+ or OH^- promoted reactions are specific acid or base processes). Some anions can effect displacement of leaving groups more rapidly than water (k_N) and in doing so catalyze hydrolysis via the sequence

$$R\text{-}X + A^- \rightarrow RA + X^-,$$
$$RA + H_2O \rightarrow ROH + HA,$$
$$(HA + X \rightleftarrows A^- + HX).$$

The use of nucleophilic anions such as phosphate or acetate at 1.0 M to 0.01 M to buffer hydrolysis reactions is a common and useful practice. Unfortunately, the general acid or base catalyzed term $k_G[A^{\ddagger}]$ added to eq (2) to give

$$k_h = k_B[OH^-] + k_A[H^+] + k_N + k_G[A^{\ddagger}],$$

can often be as large as the specific acid or base terms because both $[H^+]$ and $[OH^-]$ are relatively low in the buffered

Table 3.2. Nucleophilicity constants (n) for displacement reactions [5].

Anion	n[a]
Cl	3.04
SO_4^{2-}	2.5
NO_3	1.03
HCO_3^-	3.8
HPO_4^{2-}	3.8
OAc^-	2.72
H_2O	0.00

[a] From the Swain-Scott relation $\log(k_G/k_N) = ns$, where s is a substrate-dependent constant which varies from 0.75 to 1.5; s for MeBr is 1.00 [5]. k_G and k_N refer to specific rate constants for anion and water, respectively.

region from pH 5 to 9. As a result, the value of k_h may increase significantly. Table 3.2 lists values for k_G relative to k_N for specific ions in displacement reactions. However, the fact that primary and secondary salt effects also may be important makes it difficult to predict the overall direction or magnitude of such rate changes.

To ensure that hydrolysis measurements reported in buffered systems are not subject to these salt effects, the following procedures are recommended:

(1) Measure rates of hydrolysis at concentrations of substate less than 10^{-4} M. In most cases this procedure eliminates or minimizes the need for buffers and avoids their possible effects on rate.

(2) Check the effect of different buffers on the rate of hydrolysis; when possible use borate or acetate instead of phosphate. With low concentrations of chemicals buffers may be used at 0.01 M concentrations to hold pH constant without introducing significant salt or buffer effects.

(3) Compare the initial rates of hydrolysis in the absence and presence of buffers.

3.5. Effect of Solvent Composition

With a few exceptions, rate data reported here refer to water solvent. Many investigators have used mixed water-organic solvents to work at conveniently high concentrations of substrate, most of which are relatively insoluble in pure water. Although extrapolation of rate data from mixed solvents to water can be done with moderate success using schemes like the Winstein-Grunwald relation [6], combined extrapolations of temperature and solvent composition together with the questionable meaning of pH in mixed solvents introduce sufficient error in the final estimate to make such effort of questionable value for purposes of this review. For the most part, these data were not included.

The effect of solvent composition is most pronounced in those reactions in which charge separation is well developed in the transition state, as, for example, in solvolysis of t-butyl chloride. In ethanol-water mixtures this hydrolysis is purely solvolytic—no effect of acid or base is noted. Table 3.3 summarizes values of k_N, which increases by a factor of near 10^4 on going from 90% ethanol to water.

Table 3.3. Effect of solvent/composition on the rate of hydrolysis of t-BuCl in ethanol/water at 298 K [6].

Percent ethanol	90	70	50	40	0
k_N/s^{-1}	1.71(-6)	9.14(-6)	4.03(-6)	1.26(-4)	1.0(-2)[a]

[a] Estimated from data in table 4.1.

3.6. Effects of Metal Ion Catalysis

A number of alkaline earth and heavy metal ions catalyze hydrolysis of a variety of organic esters [7,8]. The principal function of the metal ion appears to be to increase the effective concentration of HO^- ion at pH levels where its concentration would otherwise be negligible [8,10]. This explanation is by no means universally applicable, especially to enzyme-like processes where prior complexation seems a necessary prelude to bond cleavage. Nonetheless, the kinetic features of several kinds of esters hydrolyses are similar, and they follow the same rate law

$$k_h = k_M K_A [M]_T / (K_A + [H^+]) + k_N + k_B,$$

where k_h, k_N, and k_B have their usual significance, k_M is the metal ion catalysis constant, M is the total metal ion concentration, and K_A is the equilibrium constant for dissociation of the hydrated metal ion complex $M(H_2O)_x^{2+}$:

$$M(H_2O)_x^{2+} \rightleftarrows M(H_2O)_{x-1}(OH)^{+} + H^{+}.$$

Study of the hydrolysis of the nerve agent isopropyl methyl phosphonofluoridate showed that Cu, Mn, and Mg ions were effective catalysts for hydrolysis and at 10^{-2} M concentration increased the values of k_h by factors of 7 to 32, respectively [9]. The magnitude of the catalytic constants suggests that at 1 ppm ($\sim 10^{-6}$ M) of each of these cations, their contribution would be negligible.

Similarly, the catalyzed hydrolysis of propionic anhydride by Co and other metal ions [10] may be shown to be slow relative to the water reaction at pH 7 when the Co concentration is less than 10^{-5} M, a value far in excess of the amounts of Co ever found in the environment.

Cu ion is also an effective catalyst for hydrolysis of many dialkyl thiophosphorolate esters but is completely ineffective when complexed to soil organics [11]. When complexed to montmorillonites, however, Cu ion was still effective. A rough calculation from the data indicates that at concentrations of Cu (total) found in most freshwaters, about or less than 1 ppm [12], the contribution of the catalyzed reaction to the observed rate is negligible, even if all of the Cu is assumed to be in an effective form, an unlikely situation.

There is one report by Ketelaar et al. some 20 years ago [13] that does suggest that Cu is effective in catalyzing hydrolysis of parathion at Cu concentrations below its limits of detection; only by completely removing all traces of metal ions with EDTA was it possible to measure the uncatalyzed process in glass vessels. The effect of EDTA was small, only a factor of two, but it may mean that some reported rate constants for hydrolysis of other phosphorus esters have in fact a small metal-ion catalyzed component.

In general, we believe that metal ions do not play a significant role in most hydrolysis reactions that occur in the environment; at the concentrations of metal ions normally found, the catalyzed rates are too small to contribute importantly compared to the acid, base, or water reactions; moreover, most metal ions are probably tightly complexed to natural organics and are not available as catalysts. Obviously, this area deserves much more careful study before the question can be decided fully. As a start, heavy metal concentrations in water should be reported both as soluble and insoluble metal based on

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4. Rates of Hydrolysis: Literature Review

4.1. Alkyl, Allyl, and Benzyl Halides and Polyhaloalkanes

$$RX + H_2O \longrightarrow ROH + HX$$

Aliphatic or allylic halides hydrolyze in water by neutral and base catalyzed reactions to give alcohols. In water β-elimination to give olefins and hydrogen halides is rare except at high temperature. No acid catalyzed processes are reported. Comprehensive kinetic data are available for the neutral hydrolysis of simple alkyl halides, but only scattered rate

Table 4.1. Alkyl halides: rate constants k_N, with temperature coefficients at 323.2K for $\log k_N = -\Delta H^{\ddagger}/RT + \Delta S^{\ddagger}/R + \log T + 10.32$.[a,b]

Alkyl halide	T/K	k_N/s^{-1}	$\Delta H^{\ddagger}$/kJ/mol	$10^3\Delta S^{\ddagger}$/kJ/mol K	$-10^3\Delta C_p$/kJ/mol K	Reference[c]
MeF	373	4.42(-6)[d,e]	107.1	-51.1	288	4,3
MeCl	363.1	5.64±0.2(-5)[f,g]	105.7	-37.2	218	5,3
MeBr	343.7	1.065±0.004(-4)	101.0	-27.8	195	5,3
MeI	353.2	8.19±0.06(-5)	108.6	-16.5	237	6,3
EtCl	373.2	1.148(-4)[e,h]	104.4	-21.9	209	6,3
EtBr	371.8	1.395±0.004(-3)[j]	101.5	-29.7	209	6,3
EtI	371.8	8.78±0.07(-4)	107.1	-15.1	209	6,3
i-PrCl	371.8	1.00±0.011(-3)	104.4	-22.0	161	6,3
i-PrBr	323.2	1.129±0.003(-4)	101.9	-5.98	236	5,3
i-PrI	353.2	2.75±0.004(-3)	106.8	7.78	247	6,3
n-PrBr	353.2	1.614(-4)	97.4	-42.3	209	6,3
i-BuBr	353.2	1.406(-5)[k]	111.2	-11.0	209	6,3
neo-Pentyl Br	363.4	5.08±0.04(-5)[l]	134.7	-5.73	209	6,3
t-BuF	363.4	5.18±0.02(-4)	94.6	-11.4	209	6,3
t-BuCl	287.2	6.36±0.03(-3)	91.8	34.8	188	5,3

[a] R = 0.01914 kJ/mol K.

[b] To calculate k_N at temperature T_2 where $T_1 = 323.2$K, $\Delta H^{\ddagger}_{T_2} = \Delta H^{\ddagger}_{T_1} + \Delta C_p(T_2 - T_1)$ and $\Delta S^{\ddagger}_{T_2} = \Delta S^{\ddagger}_{T_1} - \Delta C_p\ (2.303) \log\left(\frac{T_1}{T_2}\right)$

[c] First reference is for kinetic data, second reference is for temperature coefficient data.

[d] Calculated value at 373K is 1.2(-5).

[e] A better temperature dependence fit is given by $\log (k_N/s^{-1}) = 109.435 - 33.729 \log T - 10467/T$ [4]; k_N calculated at 373K using this expression is 4.29(-6).

[f] Calculated value at 363K is 4.6(-5).

[g] A better temperature dependence fit is given by $\log (k_N/s^{-1}) = 85.336 - 25.118 \log T - 9192.2/T$ [17]; k_N calculated at 363.2K using this expression is 5.44(-5).

[h] Calculated value at 373.2K is 1.1(-3).

[j] Calculated value at 371.8K is 9.5(-4).

[k] Calculated value at 353.2K is 6.4(-5).

[l] Calculated value at 363.4K is 1.4(-7).

constants and a few temperature coefficients are available for other halides. Polyhalomethanes hydrolyze by neutral and base catalyzed processes and, in the case of trihalomethanes, by proton transfer followed by loss of halide to give dihalocarbene:

$$HCX_3+OH^- \longrightarrow {}^-CX_3+H_2O$$

$${}^-CX_3 \longrightarrow CX_2+X^-.$$

Carbon tetrachloride is unique among halides in hydrolyzing by a rate process second order in carbon tetrachloride. The mechanistic significance of this unusual rate dependence is not clear.

It should be noted that certain classes of organic halogen compounds probably do not hydrolyze at rates fast enough to be detected, even at elevated temperatures. Vinyl chloride, for example, is not hydrolyzed significantly in water even at 373 K after 10 days [1]. Perhalogenated compounds—e.g., perchlorobicycloalkanes such as mirex—are resistant to strong base for short times and in boiling water for several months at 393 K [2]. Under environmental conditions, these compounds must degrade, if they do, by processes other than hydrolysis. Rate constants for hydrolysis by neutral and base promoted reactions for alkyl, allyl and benzyl halides and polyhaloalkanes are listed in tables 4.1 through 4.6.

Table 4.2. Methyl halides: rate constants k_B at 373K and temperature coefficients for log k_B = log A - (E/RT)[a] [7].

Methyl halide	$k_B/M^{-1}s^{-1}$	E_A/(kJ/mol)	log $(A/M^{-1}s^{-1})$
MeF	8.98(-4)	90.4	9.607
MeCl	2.42(-2)	101.7	12.614
MeBr	3.52(-1)	96.3	13.017
MeI	1.24(-1)	92.9	12.093

[a] R = 0.01914 kJ/mol K.

Table 4.3. Allyl halides: rate constants k_N and temperature coefficients for log k_N = - A/T -- B log T + C [8].

Allyl halide RX	$10^6k_N/s^{-1}$	T/K	A/s^{-1}	B/s^{-1}	C/s^{-1}
CH_2CHCH_2Cl[a,d,e]	4.294 ± 0.017	308.16	8345.3	24.189	81.916
CH_2CHCH_2Br[b]	16.98 ± 0.05	298.16	8794.7	28.727	95.813
CH_2CHCH_2I[c]	4.103 ± 0.012	298.13	8304.5	22.560	78.289

[a] Measured over temperature range 308-358K.

[b] Measured over temperature range 298-388K.

[c] Measured over temperature range 298-348K.

[d] These data are in good agreement with those of Shvets and Belyaev [9] in the temperature range 323-363K.

[e] log k_B = 8.92 - 75.0/RT), R = 0.01914 kJ/mol K.

Table 4.4. Polyhalomethanes: rate constants k_N and k_B.

Polyhalomethane	T/K	$10^8k_N/s^{-1}$	$10^4k_B/M^{-1}s^{-1}$	Reference
CH_2Cl_2	∿ 298[a]	1.39 ± 0.04	--	11
	373.16	57.7 ± 0.04	1.61	10
$CHCl_3$	∿ 298[a]	1.62 ± 0.27	--	11
	373.16	7.285 ± 0.0146	--	7
	307.96	--	3.082 ± 0.093	7
	298.9[b]	--	0.602 ± 0.023	12
CCl_4[c]	373.16	1.208 x 10^{-3}	< 10^{-5}[d]	7
$BrCHCl_2$[b]	298.9	--	16.0 ± 0.6	12
Br_2CHCl[b]	298.9	--	8.01 ± 0.22	12
$HCBr_3$[b]	298.9	--	3.20 ± 0.12	12
$ICHCl_2$[b]	308.9	--	39.9	12
$BrCHFCl$[b]	287.3	--	2150	12

[a] Measured in ∿ 8 x 10^{-6}M in sealed tubes without temperature control. The value of k_N is suspect.

[b] Measured in 66.67% (v/v) dioxane/water; the values for k_B for $CHCl_3$ in water and in dioxane/water agree quite closely. Values for other compounds are presumed to be close to those in water.

[c] Hydrolysis of CCl_4 is reported to be second order in CCl_4.

[d] Reference 13 .

Table 4.5. Polyhalomethanes: temperature coefficients for k_N and k_B in form log k = log A - (E/RT)[a] or log k = C - B log T - A/T.

Polyhalomethane	Temperature Coefficient ($k_B/M^{-1}s^{-1}$ or k_N/s^{-1})	Reference
CH_2Cl_2	log k_N = 98.441-29.66 log T - 10597.3/T	10
	log k_B = 11.567-109.86/RT	10
$CHCl_3$	log k_N = 149.6905-46.28 log T-14108/T	7
	log k_B = 15.90-114.2/RT[b]	12
	log k_B = 212.80249-67.18 log T-15132/T	7
CCl_4	log($k_N/M^{-1}s^-s^{-1}$) = 11.0189 - 23646/2.3 RT	7
$BrCHCl_2$	log k_B = 16.22-108.0/RT[b]	12
Br_2CHCl	log k_B = 16.44-111.3/RT[b]	12
$HCBr_3$	log k_B = 15.33-107.5/RT[b]	12
$ICHCl_2$	log k_B = 17.33-116.3/RT[b]	12
$BrCHFCl$	log k_B = 19.44-110.5/RT[b]	12

[a] R = 0.01914 kJ/mol K.

[b] Solvent 66.7% dioxane in water.

Table 4.6. Benzyl halides: rate constant k_N

$C_6H_5CH_2X$	T/K	$10^5 k_N/s^{-1}$	Reference
$C_6H_5CH_2Cl$	293.19	0.703[a]	8
$\log (k_N/s^{-1}) = 7306.02/T - 19.309 \log T + 67.404$			
$p\text{-}MeC_6H_4CH_2Cl$	303.0	86.2	6
$C_6H_5CHCl_2$	303.0	297[b]	6
$C_6H_5CCl_3$	278.1	387[c]	6
$C_6H_5CH_2Br$	303.0	27.5	6
$p\text{-}MeC_6H_4CH_2Br$	293.0	140.6	6
$o\text{-}MeC_6H_5CH_2Br$	293.0	56.0	6

[a] Rate reported to be independent of pH below 13 [14].

[b] Rate reported to be independent of pH below 13 [15].

[c] Rate reported to be independent of pH below 13 [16].

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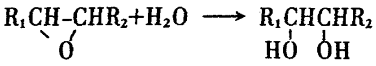
4.2. Epoxides



$$\underset{\diagdown\ O\ \diagup}{R_1CH\text{-}CHR_2}+H_2O \longrightarrow R_1\underset{HO}{CH}\underset{OH}{CHR_2}$$

Epoxides hydrolyze to give diols in acid, neutral, or base catalyzed reactions. Published rate constants for hydrolysis in water by one or more of these processes are available for a few simple aliphatic epoxides. Since these compounds are quite reactive, most rate measurements were made near 298 K. In general, there is little effect of structure in terminal epoxides on the neutral or base catalyzed processes, both of which proceed mainly by nucleophilic displacement at the terminal carbon; for these epoxides 10^7k_N=5.7±3.4 and 10^4k_B=0.85±0.2. However, acid catalyzed hydrolyses are very sensitive to substitution of an incipient carbonium ion. These same epoxides have values of k_A that increase by a factor of 6 on addition of one methyl group and 100 on addition of a second methyl group. The presence of halo or hydroxy substituents beta to the secondary epoxy-carbon slows down the rates of hydrolysis significantly. Aromatic or conjugated epoxides are much more reactive than aliphatic epoxides toward acid or water, often by factors of 10^3 to 10^5. Rate constants for acid, base and water promoted hydrolyses of aliphatic and dihydroaromatic epoxides are listed in table 4.7.

Table 4.7. Epoxides: rate constant k_A, k_B, and k_N and temperature coefficients $\log k = \log A - (E/RT)$ or $\log k = \log(ek/h) + \log T + (\Delta S^{\ddagger}/R) - (\Delta H^{\ddagger}/RT)^{a,b}$

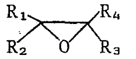
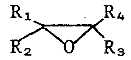
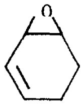
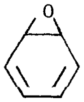
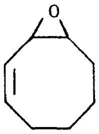
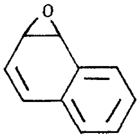
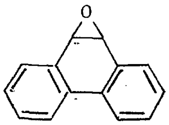
				T/K	$10^3k_A/M^{-1}s^{-1}$	$10^7k_N/s^{-1}$	$10^4k_B/M^{-1}s^{-1}$	Reference
R_1	R_2	R_3	R_4					
H	H	H	H	293	5.34	3.61	--	1
				293.2	--	4.2	0.65	2
				298	9.3	6.75	--	3
				298	10	--	1.0	4
				303.2	16.9	--	--	5
					$\log k_A = 10.753 + \log T - 0.0255/R - 79.5/RT$			5
					$\log k_N = 7.726 - 79.5/RT$			2
					$\log k_B = 9.312 - 75.3/RT$			2
Me	H	H	H	273	3.50	--	--	6
				297.7	46.5	--	--	7
				298	60	∿ 5.5	∿ 1	8
				298	--	--	0.87	4
				303.2	74.5	--	--	5
					$\log k_A = 10.753 + \log T - 0.0180/R - 77.9/RT$			5
					$\log k_A = 12.62 - 79.4/RT$			7
Me	Me	H	H	273	330	--	--	6
				298	6800	11.1	0.87	8
				298	--	--	0.76	4
				298.2	4600	--	--	5
				298.9	7350 ± 240	--	--	9
				301.9	8580	--	--	7
					$\log k_A = 10.753 + \log T - 0.0167/R - 67.4/RT$			5
					$\log k_A = 14.94 - 80.8/RT$			7
CH_2OH	H	H	H	273	0.266	--	--	6
				293.0	2.46	2.84	--	1

Table 4.7. (Cont.)

								
R_1	R_2	R_3	R_4	T/K	$10^3k_A/M^{-1}s^{-1}$	$10^7k_N/s^{-1}$	$10^4k_B/M^{-1}s^{-1}$	Reference
CH_2Cl	H	H	H	273	0.048	--	--	6
				293	0.406	9.75	--	1
				297.9	0.804	--	--	7
				298	0.8	--	--	4
					log k_A = 11.48 - 83.1/RT			7
CH_2OH	Me	H	H	273	0.75	--	--	6
				298	11	--	--	4
CH_2Br	H	H	H	273	0.038	--	--	6
CH_2Cl	Me	H	H	273	0.115	--	--	6
Me	H	Me	H	273	7.70	--	--	6
H	Me	Me	H	273	14.8	--	--	6
				298.0	1.64(7)	2.20(3)	14.8	10
				303	3.2(4)	1.40(4)	--	11
				298.0	3550	< 3.60(3)	--	10
				303	4.70(5)	3.00(4)	--	11
				303	1.30(5)	--	--	11

[a] R = 0.01914 kJ/mol K.

[b] log(ek /h) = 10.32.

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4.3. Esters

$$R_1C(O)OR_2+H_2O \rightarrow R_1C(O)OH+R_2OH$$

Esters hydrolyze to give alcohols and carboxylic acids by acid, neutral, and base catalyzed processes. In this review, data on esters have been organized into two sections: 4.3.1 esters of aliphatic acids and 4.3.2 esters of aromatic acids.

4.3.1. Aliphatic Acid Esters

The hydrolysis of esters has been the subject of a recent excellent review [16]. The hydrolysis of esters is acid and base catalyzed, and for some compounds the water promoted reaction is also important. Limited data are available for evaluation of the acid catalyzed reaction in dilute acids (less than ~0.1 M $[H^+]$, pH=1). The data listed in table 4.8 indicate that, in general, the reaction is little affected by steric or electronic influences in the ester structure; for most esters, $10^4k_A=1.0\pm0.5$ at 298 K. The base catalyzed process has been extensively studied. Steric factors significantly retard k_B, with the rate constant two orders of magnitude less for t-butyl acetate than for methyl acetate. Electron-withdrawing groups on either ester moiety increase k_B by making the ester carbonyl more susceptible to nucleophilic attack with subsequent displacement of the alcohol. In some cases, the ester is so activated that the neutral reaction with water is more rapid than the acid catalyzed process at low pH.

Table 4.8. Aliphatic acid esters, RC(O)OR: k_A, k_N, and k_B at 298K and temperature coefficients[a] $\log k = \log A-(E/RT)$ or $\log k = (-\Delta H^{\ddagger}/RT) + (\Delta S^{\ddagger}/R) + \log T + 10.32$.

$R_1C(O)OR_2$					$\Delta H^{\ddagger}$/kJ/mol or ΔE/kJ/mol	$10^3\Delta S^{\ddagger}$/kJ/K mol or log A/s^{-1} or $M^{-1}s^{-1}$	
R_1	R_2	$10^4k_A/M^{-1}s^{-1}$	$10^6k_N/s^{-1}$	$k_B/M^{-1}s^{-1}$			Reference
H	Me	--	--	28.6[b]	$\Delta H^{\ddagger} = 41.0$	$\Delta S^{\ddagger} = -77.0$	4
H	Me	--	--	36.6	E = 40.0	log A = 8.57	5
H	Et	--	--	25.7	E = 37.4	log A = 7.96	5
H	Pr	--	--	22.8	E = 35.6	log A = 7.60	5
H	Bu	--	--	21.8	E = 33.5	log A = 7.19	5
H	Pr	--	--	10.9	E = 34.8	log A = 7.13	5
Me	Me	1.133	--	0.179			7
Me	Et	1.097	0.000153	0.1077			7
		--	--	0.111			12
Me	Pr	1.100	--	0.0655[c]			10
		--	--	0.0970			12
Me	Bu	1.133	--	0.0705[c]			10
		--	--	0.065±0.001[d]	$\Delta H^{\ddagger} = 44.4$	$\Delta S^{\ddagger} = -115.9$	8
Me	i-Pr	0.600	--	0.0262			10
Me	s-Bu	--	--	0.01738			12
Me	cyclo-C_5H_9	--	--	0.026±0.001[e]	$\Delta H^{\ddagger} = 49.8$	$\Delta S^{\ddagger} = -105.4$	8
Me	_t_-Bu	1.26	--	0.00150			10

$R_1C(O)OR_2$					$\Delta H^{\ddagger}$/kJ/mol or ΔE/kJ/mol	$10^3\Delta S^{\ddagger}$/kJ/K mol or log A/s^{-1} or $M^{-1}s^{-1}$	
R_1	R_2	$10^4k_A/M^{-1}s^{-1}$	$10^6k_N/s^{-1}$	$k_B/M^{-1}s^{-1}$			Reference
Me	CH_2CH-	1.355	0.1133	10.33			7
				3.28±0.13[d]	$\Delta H^{\ddagger} = 53.6$	$\Delta S^{\ddagger} = -53.1$	8
Me	$CH_2C(CH_3)$	--	--	3.01±0.15[d]	$\Delta H^{\ddagger} = 48.5$	$\Delta S^{\ddagger} = -88.7$	8
Me	1-(cyclo-C_5H_7)	--	--	0.855±0.038[d]	$\Delta H^{\ddagger} = 47.0$	$\Delta S^{\ddagger} = -86.4$	12
Me	CH_2CHCH_2	--	--	0.209			12
Me	$CH_2CHCH(CH_3)$	--	--	0.0717			12
Me	(Me)(Et)(CH_2CH)C	--	--	0.00400			12
Me	$CHCCH(CH_3)$	--	--	0.365			12
Me	$C_6H_5CH_2$	1.09	--	0.197			10
Me	C_6H_5	--	0.066	--			2
		0.782	--	1.37			10
Me	p-MeC_6H_4	--	0.039	--			2
Me	p-ClC_6H_4-	--	0.063	--			2
Me	p-$NO_2C_6H_4-$	--	0.846	--			2
Me	3,4-$(NO_2)_2C_6H_3$	--	4.60	--			2
Me	2,4-$(NO_2)_2C_6H_3$	--	11.1	94.0			9
Me	2,6-$(NO_2)C_6H_3$	--	14.2	--			2

Table 4.8. (Cont.)

$R_1C(O)OR_2$		$10^4k_A/M^{-1}s^{-1}$	$10^6k_N/s^{-1}$	$k_B/M^{-1}s^{-1}$	$\Delta H^‡$/kJ/mol or ΔE/kJ/mol	$10^3\Delta S^‡$/kJ/K mol or log A/s^{-1} or $M^{-1}s^{-1}$	Reference
R_1	R_2						
$ClCH_2$	Et	--	0.1083	36.7			1
$ClCH_2$	$p-NO_2C_6H_4$	--	3.3	5920			9
Cl_2CH	Me	2.33	15.33	2830			7
Cl_2CH	Et	--	4.91	883			1
$ClCH_2$	C_6H_5	--	1760	12800			9
F_2CH	Et	--	56.7	4500			1
Cl_3C	Me	--	733	--			11
F_3C	Me	--	3280±5	--	$\Delta H^‡ = 44.35 \pm 0.33$	$\Delta S^‡ = 135.1 + 1.3$	13
F_3C	Et	--	2735±12[f] 1253±3[g]	--	$\log k_N = -6439.0409/T - 32.714 \log T + 100.0558$		13
F_3C	i-Pr	--	3494±12	--	$\Delta H^‡ = 43.97 \pm 0.21$	$\Delta S^‡ = 155.2 \pm 2.5$	13
F_3C	t-Bu	--	1320±43[h]	--	$\log k_N = -10612/T - 38.35 \log T + 127.63$		14
(CH_3S-CH_2)	Et	--	--	0.92			15
$(^{-}S-CH_2-)$	Et	--	--	0.0064			15
CH_3SOCH_2-	Et	--	--	4.2			15
$(CH_3SO_2CH_2)$	Et	--	--	12.8			15
$(CH_3)_2\overset{+}{S}CH_2$	Et	--	--	204			15
$2,4-(Cl_2)C_6H_3OCH_2$	$BuOCH_2CH_2$	6.59[j]	--	--	$\Delta H^‡ = 73.6$	$\Delta S^‡ = 89.1$	6
$C_6H_5OCH_2$	Et	--	--	30.2[ℓ]	$\Delta H^‡ = 84.1$	$\Delta S^‡ = 61.9$	6
$C_6H_5OCH_2$	Me	--	--	17.3[ℓ]			6
Et	Et	0.33[c]	--	0.059[c]	--	--	17
			--	0.087	E = 44.85	log A = 6.780	18
n-Pr	Et	0.18[c]	--	0.035[c]	--	--	17
		--	--	0.038	E = 43.18	log A = 6.142	18
i-Pr	Et	--	--	0.023[d]	E = 42.55	log A = 5.93	19
$CH_2=CH$	Et	0.012[c]	--	0.064[c]			17
				0.078	E = 50.00	log A = 7.638	20
trans-$CH_3CH=CH$	Et	0.063[c]	--	0.0091[c]	--	--	
		--	--	0.013	E = 53.22	log A = 7.431	20
cis-$CH_3CH=CH$	Et	--	--	0.015	--	--	20
$H-C\equiv C-$	Et	--	--	4.68	E = 51.30	log A = 9.638	20
$CH_3C\equiv C-$	Et	--	--	0.568	E = 53.16	log A = 9.055	20

[a]R = 0.01914 kJ/mol K; [b]293.3K; [c]293K; [d]293.2K; [e]294.1K; [f]295.2K; [g]283.1K; [h]297.8K; [j]340.2K; [ℓ]301.2K.

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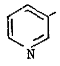
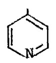
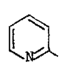
4.3.2. Aromatic Acid Esters

Little work has been carried out in water as the sole solvent for aromatic ester hydrolyses, primarily because of the low solubility of these compounds in water. Consequently, cosolvents with water such as dioxane, ethanol, methanol, and acetone, among others, have been used to expedite laboratory studies. Data listed in table 4.9 indicate that increasing amounts of organic cosolvent decrease rate constants k_A and k_B. Unfortunately, these studies fail to examine pure water solvent but rates do not vary more than an order of magnitude in the solvent compositions used. Whereas the base hydrolysis studies have often been carried out around 298 K, the slower acid hydrolyses were usually carried out at higher temperatures, but some temperature dependence data are available to calculate k_A at lower temperatures. No data are available for water promoted hydrolysis reaction of aromatic acid esters. As with the aliphatic acid esters, k_A is generally insensitive to structure variations, and k_B is increased when electron-withdrawing substituents are present on the aromatic ring.

Table 4.9. Aromatic acid esters; k_A and k_B at 298K and temperature coefficients[a] for acid or base catalyzed reactions.

$R_1-\overset{O}{\overset{\|}{C}}-O-R_2$			$10^3k_B/M^{-1}s^{-1}$	% Organic solvent[b]	log k = log A - 2.303 E_a/RT		
R_1	R_2	$10^4k_A/M^{-1}s^{-1}$	at 298K	in H_2O	E_a	log A	Ref.
C_6H_5	Me	1.68[c]		60 v% M	74.0	6.57	4
			1.86	50 w% M			2
			9.01[j]	56 w% A	59.3	8.38	5
C_6H_5	Et	3.9[d]	30[e]	water			3
			8.7[e]	33 % D			3
C_6H_5	Et	0.845[f]		60 v% E	82.13	7.46	4
			1.20[g]	67.1 w% E	69.3	9.28	1
C_6H_5	Et	0.209[h]		60 v% A	84.7	7.84	4
			2.87[j]	56 w% A	60.9	8.18	5
			2.69	55 w% A			2
C_6H_5	Et		9.72[n]	40 v% D	51.1	6.80	7
			5.08[n]	70 v% D	58.4	7.76	7
C_6H_5	i-Pr		6.2[e]	water			3
p-BrC_6H_4	Et	0.200[ℓ]		60 v% E	73.30	6.29	4
		0.185[ℓ]		60 v% A	84.18	7.72	4
p-BrC_6H_4	Me		5.50	55 w% M			2

$R_1-\overset{O}{\overset{\|}{C}}-O-R_2$			$10^3k_B/M^{-1}s^{-1}$	% Organic solvent[b]	log k = log A - 2.303 E_a/RT		
R_1	R_2	$10^4k_A/M^{-1}s^{-1}$	at 298K	in H_2O	E_a	log A	Ref.
p-$NO_2C_6H_4$	Me	1.65[m]		60 v% M	73.51	6.52	4
			74.1	55 w% M			2
			640[j]	56 w% A	51.4	8.84	5
p-$NO_2C_6H_4$	Et	0.286[ℓ]		60 v% E	73.3	6.29	4
		0.271[ℓ]		60 v% A	82.2	7.57	4
			245	55 w% A			2
			244[j]	56 w% A	51.9	8.55	5
p-$NO_2C_6H_4$	Et		58.2[n]	60 v% D	48.4	8.10	7
			50.2[n]	70 v% D	51.4	8.55	7
C_6H_5	$C_6H_5CH_2$		12.11[n]	50 v% D	52.2	7.07	7
			8.52[n]	70 v% D	55.6	7.51	7
C_6H_5	$C_6H_5CH_2$		9.31[n]	50 v% A	56.7	7.74	7
			5.90[n]	70 v% A	60.8	8.25	7
o-$C_6H_4[C(O)OEt]_2$			14.85/2.47[n,r]	50 v% D	52.0/53.8	7.13/6.67	8
			10.20/2.19[n,r]	70 v% D	57.8/59.2	7.97/7.55	8
			12.72/2.05[n,r]	50 v% A	57.1/59.3	7.95/7.54	8
			8.14/1.75[n,r]	70 v% A	64.6/64.8	9.03/8.30	8
o-$C_6H_4[C(O)OCH_2C_6H_5]_2$			24.2/4.03[n,r]	50 v% D	50.9/53.2	7.16/6.77	8
			17.05/3.60[n,r]	70 v% D	54.6/58.1	7.64/7.57	8
			18.65/3.09[n,r]	50 v% A	56.4/58.4	7.98/7.56	8
			11.81/2.85[n,r]	70 v% A	60.0/62.4	8.65/8.20	8
p-$C_6H_4[C(O)OMe]$			252/16.1[p,r]	60 w% D			6
			197/29.0[p,r]	80 w% D			6
p-$C_6H_4[C(O)OEt]_2$			147/52.9[q,r]	50 v% A	56.2/58.2	8.69/8.60	8
			97.6/46.0[q,r]	70 v% A	62.6/63.5	9.61/9.42	8
	Me		42.6	55 w% M			2
	Et		117	55 w% A			2
	Me		218	55 w% M			2
	Et		631	55 w% A			2
	Me		63.1	55 w% M			2
	Et		200	55 w% A			2
			540	pure water			9

[a] R = 0.01914; [b] Organic solvent legend: M = methanol; A = acetone; D = dioxane; E = ethanol; [c] 373.4K; [d] 372K; [e] 298.3K; [f] 372.8K; [g] 298.2K; [h] 353.5K; [j] 298.1K; [ℓ] 353.4K; [m] 373.3K; [n] 303.2K; [p] 298.4K; [q] 308.2K; [r] $k_{B(1)}/k_{B(2)}$ represents first and second stage of hydrolysis.

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4.4. Amides

$$RC(O)NR_1R_2+H_2O \longrightarrow RC(O)OH+R_1R_2NH$$

Amides hydrolyze by acid and base catalyzed processes to carboxylic acids and amines. Simple amides are about 1000 times less reactive than simple esters in base catalyzed processes but about as reactive in the acid catalyzed process (see tables 4.8 and 4.10). This reactivity pattern is consistent with the relative ease of removal of $-OR$, $-NR_2$, and HN^+R_2 as leaving groups from the tetrahedral intermediate common to most acyl hydrolysis reactions. No neutral hydrolysis rate data are reported; probably as for simple esters the rate constants are too small ($<10^{-9}s^{-1}$) to measure. Structure-reactivity patterns for amides are much like those of esters (see section 4.3).

Table 4.10 summarizes data for 24 aliphatic acid amides and their C- or N-substituted derivatives, all measured at 348 K.

Table 4.10. Amides: k_A, k_B at 348K and temperature coefficient[a] $\log k = (-\Delta H^{\ddagger}/RT) + (\Delta S^{\ddagger}/R) + \log T + 10.32$.

Amide	$10^4 k_A/M^{-1}s^{-1}$	$10^4 k_B/M^{-1}s^{-1}$	$\Delta H^{\ddagger}_A$	$\Delta H^{\ddagger}_B$	$-10^3\Delta S^{\ddagger}_A$	$-10^3\Delta S^{\ddagger}_B$	Reference	
			kJ/mol		k J/mol K		A	B
Acetamide	10.3(8.54)	13.6(18.67)	80.3	55.31	72.8	141.8	1(3)	2(3)
Propionamide	12.0	13.1	75.7	61.7	84.5	124.3	1	2
Valeramide	5.93	5.52	78.2	60.58	83.7	134.3	4	2
Isovaleramide	1.29	1.97	81.7	72.68	87.4	108.4	4	2
t-Butylacetamide	0.193	--	86.2	--	88.7	--	4	--
Dimethylacetamide	6.22(6.06)	6.61	80.3	81.4	76.6	74.1	1(5)	4
Phenylacetamide	5.19	17.7	75.3	49.25	91.6	163.2	4	2
Cyclohexylacetamide	1.24	1.77	87.03	68.87	70.3	120.1	4	2
Trimethylacetamide	2.26	2.57	83.3	71.1	77.0	110.9	5	2
Diethylacetamide	0.176	--	91.8	--	73.6	--	5	--
Methoxyacetamide	8.98	8.56	79.1	56.2	77.0	121.3	4	--
Chloroacetamide	12.1	1400	78.2	--	78.7	--	4	1
Bromoacetamide	4.79	--	77.4	--	80.8	--	1(4)	--
Dichloroacetamide	--	18400(3000)	--	--	--	--	--	1
Trichloroacetamide	--	135000(9400)	--	--	--	--	--	1
Cyclohexylcarboxamide	3.96	4.24	84.6	53.6	68.2	157.3	5	2
Butane-2-carboxamide	1.51	1.65	87.0	64.2	69.0	133.9	5	2
N,N-R_1R_2 Acetamide								
R_1 R_2								
H methyl	0.582(0.425)	3.58	87.0(86.2)	69.4	77.8(90.4)	121.8	1(3)	3
H ethyl	0.233	1.80	92.5	67.4	77.8	132.6	3	3
H isopropyl	0.090	0.367	--	73.2	--	123.0	3	3
methyl methyl	0.654(0.377)	5.18	81.6(85.8)	63.2	92.0(92.9)	131.4	1(3)	3
ethyl ethyl	0.0227	0.1167	--	75.3	--	130.1	3	3
methyl ethyl	0.102	0.983	--	67.8	--	128.0	3	3

[a] R = 0.01914 kJ/mol K.

[b] k_A and k_B calculated from temperature coefficient data are good to within ± 5% of measured value listed unless otherwise noted.

[c] Calculated value 1.08(-4).

[d] Calculated value 5.93(-4).

[e] Calculated value 5.96(-4).

[f] Calculated value 8.80(-4).

[g] Calculated value 1.34(-4).

[h] Calculated value 1.03(-3).

[j] At 25°C, estimated from log k_B versus 1/T plots given in reference 6 (no primary data given).

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4.5. Carbamates

$$\mathrm{ROC(O)NR_1R_2+H_2O \longrightarrow ROH+CO_2+HNR_1R_2}$$

Carbamates hydrolyze by acid, base, or neutral processes to give alcohols, amines, and CO_2, probably via the intermediate carbamic acid. Carbamates are moderately reactive at 298 K, where most data are reported, exhibiting a range of rate constants (k_B) that spans 8 powers of 10. The striking increase in reactivity toward OH^- for carbamates having an -RNH structure is caused by a change in mechanism from one involving the usual addition of OH^- to the carbonyl to one involving proton transfer from NH followed by loss of RO^- to form an intermediate isocyanate:

$$\mathrm{ROC(O)NHR+OH \longrightarrow ROC(O)NR+H_2O}$$

$$\mathrm{ROC(O)NR \longrightarrow RNCO+RO^-}$$

$$\mathrm{RNCO+H_2O \longrightarrow RNH_2+CO_2.}$$

Table 4.11 summarizes data for 20 aliphatic and aromatic carbamates.

ble 4.11. Carbamates $R_1OC(O)NR_2R_3$: k_A, k_N, k_B at 298K and temperature coefficients[a] $\log k = -\Delta H^{\ddagger}/RT + \Delta S^{\ddagger}/R + \log T + 10.32$ or $\log k = A - E/RT$.

$R_1OC(O)NR_2R_3$			k_A, k_N or k_B [b,c]			
R_1O	R_2-N-R_3		($M^{-1}s^{-1}$ or s^{-1})	$\Delta H^{\ddagger}$/kJ/mol	$-10^3\Delta S^{\ddagger}$/kJ/mol K	Reference
Me	C_6H_5	H	5.5(-5)	--	--	5
Et	C_6H_5	H	3.3(-5)[d]	66.5	106.7	2
Et	C_6H_5	Me	5.0(-6)[e]	54.0	164.4	2
C_6H_5	C_6H_5	H	4.7(-1)[f]	69.45	-21	2,4
C_6H_5	C_6H_5	Me	4.2(-5)[g]	62.8	117.2	2
p-MeO	C_6H_5	H	2.52(1)	--	--	4
m-Cl	C_6H_5	H	1.83(3)	--	--	4
p-NO_2	C_6H_5	H	2.71(5)[h]	--	--	4
p-NO_2	C_6H_5	Me	7.98(-4)	--	--	5
p-NO_2	Me	H	k_N = < 2.52(-5)	--	--	3
			k_B = 3.0(-3)	--	--	3
1-$C_{10}H_9$	Me	H	k_A = 1.4(-7)	$\log k_A = 2.48 - 53.01/RT$[k]		1
			k_N = 3.69(-5)[j]	$\log k_N = 4.82 - 62.59/RT$[k]		1
			k_B = 3.4[ℓ]	$\log k_B = 11.95 - 70.11/RT$		6
1-$C_{10}H_9$	Me	Me	k_A = 1.1(-9)	$\log k_A = 8.00 - 96.94/RT$[k]		1
			k_N = 2.30(-8)[h]	$\log k_N = 7.48 - 102.22/RT$[k]		1
			k_B = 4.55(-5)[h]	$\log k_B = 9.37 - 92.51/RT$[k]		1
$Et_2NCH_2CH_2$	C_6H_5	H	2.6(-5)	73.2	87.0	2
$Et_2NCH_2CH_2$	$Me_3C_6H_2$	H	9.2(-7)	103.3	13.8	2
m-$Me_3\overset{+}{N}C_6H_4$	Me	H	6.7(-1)	76.6	- 47.7	2
m-$Me_3\overset{+}{N}C_6H_4$	Me	Me	2.8(-4)	59.4	113.0	2
$ClCH_2CH_2$	C_6H_5	H	1.59(-3)	--	--	5
Cl_2CHCH_2	C_6H_5	H	5.00(-2)	--	--	5
CCl_3CH_2	C_6H_5	H	3.16(-1)	--	--	5
CF_3CH_2	C_6H_5	H	1.00(-1)	--	--	5

[a] R = 0.01914 kJ/mol K; [b] Value given is for k_B unless indicated otherwise; [c] Standard deviation for most values of k_B is ± 2%; [d] Other workers [5] report 3.2(-5); k_B calculated from temperature coefficients is 3.7(-5); [e] Other workers [5] report 3.98(-6); k_B calculated from temperature coefficients is 5.5(-6); [f] Other workers [4] report 5.42(1); k_B calculated from temperature coefficients is 5.2(1); [g] Other workers [5] report 1.41(-4); k_B calculated from temperature coefficients is 4.6(-5); [h] Measured at pH 6.5, no contribution from k_N; [j] At 353K; [k] Calculated from data in reference 1; [ℓ] At 296K; other workers [1,7] report 50 and 5.7, respectively, at 298K; on basis of agreement of data in reference (6,7) the data in table is preferred.

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4.6. Phosphoric and Phosphonic Acid Esters

$$ROP(X)R_1R_2+H_2O \longrightarrow HOP(X)R_1R_2+ROH$$

Pentavalent phosphorus forms a great variety of ester and amide derivatives, many of which have significant biological activity. The cholinesterase-inhibiting properties of many of these compounds have led to their extensive application as pesticides and to preparation of many candidate chemical warfare agents.

The chemistry of organic phophorus compounds has been reviewed extensively [1,2,3]. Hydrolysis of pentavalent phosphorus compounds proceeds by nucleophilic displacement at phosphorus or carbon, with formation of leaving groups such as RO^-, R_2N^-, X^-, RS^-, or $(RO)_2PO_2^-$. Although almost all phosphorus ester hydrolyses exhibit promotion by OH^-, only a few esters are hydrolyzed by acid or water promoted reaction.

Mechanisms of hydrolysis of phosphorus esters have also been examined extensively [1-6]. Unlike carboxylic acid derivatives, the phosphorus esters do not form pentacovalent intermediates with water or OH^-. Thus, reactions involving P-O cleavage must involve synchronous bond making and breaking between phosphorus and the entering and leaving groups.

Structure-reactivity relationships among phosphorus esters are more complex than for simple carboxylic esters, but some generalizations are possible within specific classes. More electron-withdrawing substituents accelerate hydrolysis; bulky substituents retard hydrolysis; among substituents, those groups having the more acidic conjugate acids are better leaving groups and their esters cleave more rapidly.

Hydrolysis may involve C-O rather than P-O bond cleavage:

$$H_2O+R_2CH\text{-}OPR_3 \longrightarrow {}^{+}H_2OCHR_2+{}^{-}OPR_3.$$

C-O cleavage is dominant for many acid and water promoted hydrolyses of esters; P-O cleavage is associated predominantly with base hydrolysis [4-6]. Thus, as pH is lowered, the mechanism, stereochemistry, and products change, usually with a decrease in overall rate of hydrolysis.

As expected for an Sn1 or Sn2 process, displacement at carbon is most sensitive to steric constraints at carbon, to charge-stabilizing substituents, and to solvent. Thus, in hydrolysis of dialkyl alkylphosphonates (tables 4.12 and 4.13), isopropyl is more readily hydrolyzed than methyl (by 12×) in the acid process, but the hydrolysis of isopropyl is much slower (by 1000×) in the base process. Similarly, solvent effects on hydrolysis are more marked for the acid or neutral process than for the base process (see table 4.14).

Competition between P-O and C-O bond cleavage in neutral and acidic water has been worked out in detail for methyl phosphates by Bunton and his coworkers [4-5]. Table 4.15 separates the rate constants for the two processes with the neutral, monoanion, and monocation species.

The pH dependence and kinetics of hydrolysis of phosphorus esters can assume complex forms owing to the formation of intermediate phosphorus acids that can catalyze hydrolysis of the parent ester, as, for example, with $(RO)_2P(O)F$ (table 4.16). In other cases the species that hydrolyzes most rapidly is formed in a preequillibrium with HO^- or H^+, and the rate law exhibits a pH maximum, usually below pH 4 [6]. For these reasons, special attention must be given to control of pH and the effect of pH on rate should be clearly defined.

Unfortunately, many studies of hydrolysis of phosphrous esters were done in mixed solvents to facilitate solubility, and with limited or no study of the effect of pH. As a result, data summarized in tables 4.12-4.17 are incomplete in many ways, with some listings for hydrolysis in mixed solvents. Despite these shortcomings, where it is possible to compare rate constants measured in water by different investigators, the agreement is surprisingly good: for example, four determinations of the value of k_N for $(EtO)_3PO$ reported over a period of 75 years agreed within ±80-100% (table 4.14). Also, the precision within a single investigation is often very high <±10% on k. But however precise, if k_A or k_N is measured in 50/50 dioxane/H_2O rather than pure water, it is likely that k will be 20 times too small for accurate estimates of environmental persistence.

Table 4.12. Phosphonic acid esters: dialkylphosphonates, $R_1P(O)(OR_2)_2$: k_A and temperature coefficients $\log k_A = \log A - (E/RT)^a$ [7].

$R_1P(O)(OR_2)_2$		T/K	$10^5k_A/M^{-1}s^{-1}$	Temperature Coefficient $\log k_A =$
Me	Me	372.8	1.14	10.7 - 111.6/RT
Me	Et	382	2.05	10.9 - 111.6/RT
Me	i-Pr	366.6	12.86	11.35 - 107/RT
Me	t-$BuCH_2$	376	0.944	----
Me	C_6H_5	383	0.50	----
Me	Et, $NP^{b,c}$	383.4	1.53	----
Et	i-Pr	365.5	3.4	10.58 - 108.7/RT
n-Pr	i-Pr	369	4.9	10.72 - 108.7/RT
n-Bu	Et	383	1.05	----
n-Bu	i-Pr	369	6.47	----
t-Bu	Me	393	2.25	----
t-Bu	i-Pr	388.2	32.8	----
C_6H_5	Et	374	1.05	10.6 - 111.5/RT

[a] R = 0.01914 kJ/mol K.

[b] p-Nitrophenyl.

[c] Hydrolysis product is ethanol.

Table 4.13. Phosphonic acid esters: dialkylphosphonates, $R_1P(O)(OR)_2$ k_B and temperature coefficient $\log k_B = \log A - (E/RT)$.[a]

$R_1P(O)(OR)_2$		T/K	$10^4 k_B/M^{-1}s^{-1}$	Temperature Coefficient, $\log k_B$ =
Me	Me	322.8	146	7.30-56.5/RT
Me	Et	333	26.0	6.60-58.6/RT
Me	i-Pr	353	0.153	4.43-62.3/RT
Me	t-$BuCH_2$	361	0.958	
Me	Et, NP[b,c]	298	400	7.40-50.2/RT
Et	Et	342.5	17.8	6.31-59.4/RT
Et	i-Pr	377.9	1.06	4.45-67.8/RT
n-Pr	n-Pr	368	3.58	
n-Pr	i-Pr	382.9	0.516	3.88-66.5/RT
C_6H_5	Et[d]	332.8	44.7	

[a] R = 0.01914 kJ/mol K.

[b] p-Nitrophenyl.

[c] Hydrolysis product is p nitrophenol.

[d] Other workers [8] report 48(-4) at 353K.

Table 4.15. Methylphosphates $(MeO)_{3-n}P(O)(OH)_n$: rate constants for C-O and P-O fission in anion, neutral, and conjugate acid forms at 373K [4,5].

Ester	$10^6 k/s^{-1}$ or $M^{-1}s^{-1}$					
	Anion		Neutral		Acid	
	C-O	P-O	C-O	P-O	C-O	P-O
$(MeO)_3PO$	-	-	36.5	-	-	-
$(MeO)_2P(O)OH$	< 1(-4)	0.001	3.3	0.9	0.91	0.11
$(MeO)P(O)(OH)_2$	-	8.23	0.5	-	2.00	1.08

Table 4.14. Phosphoric acid esters $(RO)_3PO$: k_N, k_B, and temperature coefficients $\log k = \log A - (E/RT)$[a]

$(RO)_3PO$	T/K	k_N/s^{-1}	T/K	$k_B/M^{-1}s^{-1}$	Reference
$MeOP(O)(OH)_2$	373	5(-7) 8.2(-6)[b]	--	--	6
$(MeO)_2P(O)OH$	373	3.3(-6)	--	--	5
$(MeO)_3PO$	318	1.6(-7)	298	1.3(-4)	4,1
	373	3.7(-5)	308	3.3(-4)	4
$\log k_N = 8.9 - 95.0/RT$; $\log k_B = 8.1 - 67.8/RT$					4
$(MeO)_2P(O)OEt$	375	2.5(-5)	375	4.5(-2)	9
$(EtO)_3PO$	353	1.4(-6)	298	8.2(-6)[c]	8,11
	375	1.3(-5)	374	4.49(-2)	9
	374	8.35(-6)	--	--	10
	352	1.27(-6)	--	--	10
$(EtS)_3PO$[d]	355	∿ 7 (-7)	298	1.2(-2)	12
$\log k_N = 9.25 - 102.1/RT$; $\log k_B = 7.94 - 62.7/RT$					8,16
$(C_6H_5O)_3PO$	373	3(-9)[d]	298	1.06(-2)	4
$\log k_B = 6.11 - 45.9/RT$[e]					4
$(EtO)_2P(O)NP$[f]	--	--	298	6.7(-3)[g]	13
$(EtO)(EtS)P(O)NP$	310	1.3(-5)	--	--	13
$(EtO)P(O)(NP)_2$	298	3.33(-6)[g]	298	5.25(-1)[g]	13
$\log k_N = 4.95 - 59.8/RT$; $\log k_B = 10.18 - 59.8/RT$					13
$(NP)_3PO$[g]	298	1.02(-3)[g]	298	3.43(1)[g]	13
$\log k_N = 0.062 - 17.1/RT$; $\log k_B = 4.56 - 17.1/RT$					13

[a] R = 0.01914 kJ/mol K

[b] Value for the monoanion.

[c] Measured in dioxane/water (1:1 v/v); only a small solvent effect is expected from the rate of hydrolysis of $(MeO)_3PO$ [1].

[d] Measured in dioxane/water (3:2 v/v).

[e] Original reference miscalculated A and E.

[f] NP = p-nitrophenyl.

[g] Measured in acetone/water (1:1 v/v).

Table 4.16. Thiophosphoric acid esters $R_1R_2R_3PS$: k_N, k_B and temperature coefficients $\log k = \log A - (E/RT)$[a]

Ester	T/K	$k_B/M^{-1}s^{-1}$	Temperature Coefficient	Reference
$(MeO)_2P(S)NP$[b]	313	$k_N/s^{-1} = 6.97 \pm 0.44(-7)$	$\log k_N = 10.3-98.3/RT$	14
	293	5.95(-3)	$\log k_B = 13.66-89.1/RT$	15
$(MeO)P(S)(NP)_2$	298	2(-2)		13
$(MeO)_2P(S)(SCHCH_2[CO_2Et]_2)$	298	5.16[c], 5.5[g,h]	$\log k_B = 18.3-100/RT$[f]	16,22
$(EtO)C_6H_5P(S)(NP)$	298	1.52(-1)[d]	$\log k_B = 12.01-73.20/RT$	16
$(NP)_3PS$	298	2.1(-1)		15
$(i\text{-}PrO)_2P(S)NP$	298	5.5(-5)		15
$(EtO)_2P(S)NP$	298	2.17(-4)[e]	$\log k_B = 10.2-80.3/RT$	16

[a] R = 0.01914 kJ/mol K.

[b] NP = p-nitrophenoxy.

[c] Hydrolyzed in 25% acetone/H_2O.

[d] Hydrolyzed in 20% ethanol/H_2O.

[e] Hydrolyzed in 50% acetone/H_2O.

[f] Recalculated A and E from experimental data.

[g] Measured at 300 K.

[h] $k_A = 4.8 \pm 0.2(-5)$.

Table 4.17. Dialkylhalophosphonates and dialkylhalophosphorates. $R_1R_2P(O)X$: k_N, k_B and temperature coefficients.[a]

$R_1P(O)XR_2$		T/K	k_N/s^{-1}	$k_B/M^{-1}\ s^{-1}$	Reference
			X = F		
Me	Me	298	4(-4)	--	17
Et	Et	298	--	13.9	20
Et	EtO	298	--	3.3(-1)	20
EtO	EtO	--	--	1.8	20
i-PrO	i-PrO	298	1.7(-6)[b]	--	18,19
i-PrNH	i-PrNH	298	3.8(-8)	18.4	21,15
i-PrNH	Me_2N	298	--	8.46	15
Me_2N	Me_2N	301	--	4.1(-3)	15
Et_2N	Et_2N	301	--	2.5(-3)	15
			X = Cl		
Et	Et	273	1.5[c]	--	20
			$\log k_N = 5.9 - 30.5/RT$		
Et	MeO	273	1.6(-3)	--	20
			$\log k_N = 5.7 - 35.1/RT$		
MeO	MeO	273	2.9(-5)	--	20
			$\log k_N = 5.7 - 44.3/RT$		
Me	i-PrO	298	6.9(-4)	--	20

[a] R = 0.01914 kJ/mol K.

[b] k_A = 7.6(-2) at 298K [18].

[c] Measured in 5% (v/v) aqueous acetone.

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4.7. Acylating and Alkylating Agents and Pesticides

Table 4.18 lists rate constants for a variety of acylating and alkylating agents that have little in common except the troublesome property of being mutagenic toward certain yeasts (Ames test) and in some cases being produced and dispersed to the environment in significant quantities. Also listed in this table are rate constants for several important pesticides that are not readily classified into tables 4.1 to 4.17. Some of the data on pesticides have been obtained under environmental conditions including use of natural waters.

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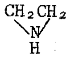
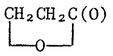
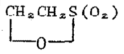
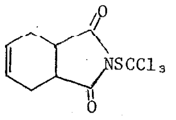
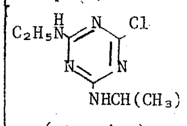
Table 4.18. Acylating and alkylating agents and pesticides: k_A, k_N, k_B and temperature coefficients.

Compound	T/K	$k_A/M^{-1}s^{-1}$	k_N/s^{-1}	$k_B/M^{-1}s^{-1}$	Temperature Coefficient[a]	Reference
	298	5.2(-7)[b]	--	--		1,2
 $CH_2CH_2C(O)$	298	--	3.3(-3)	--	$\log k_N = 10.04 - 81.59/RT$	3
 $CH_2CH_2S(O_2)$	298	-	2.15(-5)	--		4
$(CH_3O)_2SO_2$	298	--	1.66(-4)	1.48(-2)		5
$(ClCH_2)_2O$	293	--	1.8(2)[c]	--	$\log k_N = 9.807^d - 64.78/RT$	6
$C_6H_5N{=}N(CH_3)_2$	310	--	5.5(-5)	--		7
$C_6H_5C(O)Cl$	298	--	4.2(-2)[e]	--		8
$(CH_3)_2NC(O)Cl$	298	--	> 2.5(-3)	--		9
$CH_3OC(O)Cl$	298	--	5.642±0.002(-4)	--		10
$(CH_3OC_6H_4)_2CHCCl_3$ (methoxychlor)	358	--	2.4(-5)	--	E_N = 90.11 kJ/mol K	11
	338	--	4(-6)	--		
	298	--	--	3.64(-4)		12
	298	--	2.97(-8)	--		
 (captan)[f]	300	--	1.87(-5)	5.7(2)		11
 (atrazine)	298	3.9(-5)	7.6(-5)			11

[a] R = 0.01914 kJ/mol K.

[b] Rate constant for hydrolysis of protonated aziridine [1] the pK_a of which is 8 [2].

[c] No acid or base catalyzed process observed.

[d] Calculated from k and E.

[e] In 70:30 (v/v) water-acetone.

[f] Hydrolysis gives imide, sulfur, carbon dioxide and chloride ion.

5. Rates of Hydrolysis: Estimates of Lifetimes Under Environmental Conditions

This section provides estimated rate constants k_A, k_B, k_N, k_h, and half-lives of organic compounds in hydrolysis reactions in water at 298 K and pH 7. Rate constants k_A, k_B, and k_N for individual hydrolysis reactions were calculated at 298 K using the temperature coefficients tabluated in section 4 or were used directly if measured at 298 K; values of k_h were calculated from eq (2), and half-lives were calculated from eq (5):

$$t_{1/2}=0.693/k_h.$$

These tables also list the intercepts, I, corresponding to pH values where two rate processes contribute equally to the observed rate. These values were calculated from the relations in section 3. Inspection of values of I for specific compounds and classes quickly indicates the possible importance of acid or base hydrolysis in the pH range of concern in freshwater systems, namely pH 5–9.

5.1. Alkyl, Allyl, and Benzyl Halides and Polyhaloalkanes

Tables 5.1–5.4 list relevant rate constants for hydrolysis at 298 K and half-lives for halogen-containing organics. Hydrolytic reactivity of these compounds spans 11 powers of 10: benzotrichloride has an estimated half-life of only 3 minutes while chloroform and carbon tetrachloride (at 1 ppm) have half-lives of 3100 and 7000 years, respectively. Under environmental conditions only alkyl fluorides, neopentyl halides, and methyl chloride are likely to be hydrolyzed too slowly for this pathway to be important. All other alkyl, allyl, and benzyl halides are hydrolyzed rapidly enough (30 minutes to 169 days) that hydrolytic degradation should be considered as competitive with microbiological degradation, volatilization, and absorption.

The data suggest the following generalizations about RX:

(1) The rate of hydrolysis is greatest when X is Br and least when X is F. Br is more reactive than Cl by a factor of 5 to 10.

(2) Rate of hydrolysis increases as R goes from primary to secondary to tertiary in the ratio 1:1:1000 for Cl and from primary to secondary in the ratio 1:20 for Br.

(3) Allyl groups enhance the rate of hydrolysis of primary X by a factor of 5 to 100; benzyl groups enhance the rate by a factor of 50.

Table 5.1. Organic halides: hydrolysis rate constants k_N and k_B at 298K and pH 7, and I_{NB}

Organic halide	k_N/s^{-1}	$k_B[OH^-]/s^{-1}$	I_{NB}
MeF	7.44(-10)	5.82(-14)	11.1
MeCl	2.37(-8)	6.18(-13)	11.6
MeBr	4.09(-7)	1.41(-11)	11.5
MeI	7.28(-8)	6.47(-12)	11.0
CH_2Cl_2	3.2(-11)	2.13(-15)	11.1
CH_2CHCH_2Cl	1.16(-7)	6.24(-12)	11.3
$C_6H_5CH_2Cl$	1.28(-5)		> 13
$C_6H_5CHCl_2$	1.56(-3)		> 13
$C_6H_5CCl_3$	6.3(-2)		> 13

Table 5.2. Alkyl halides: k_h and $t_{1/2}$ at 298K[a] and pH 7.

Alkyl halide	k_h/s^{-1}	$t_{1/2}$
MeF[b]	7.44(-10)	30 years
MeCl[c]	2.37(-8)	0.93 years
MeBr	4.09(-7)	20 days
MeI	7.28(-8)	110 days
EtCl	2.10(-7)	38 days
EtBr	2.64(-7)	30 days
EtI	1.62(-7)	49 days
i-PrCl	2.12(-7)	38 days
i-PrBr	3.86(-6)	2.1 days
i-PrI	2.77(-6)	2.9 days
n-PrBr	3.04(-7)	26 days
t-BuF	3.87(-5)	50 days
t-BuCl	3.02(-2)	23 seconds

[a] $k_h = k_N$ since k_B is not important below pH 10 (see Table 5.1).

[b] Calculated from equation in footnote j, table 4.1.

[c] Calculated from equation in footnote h, table 4.1.

Table 5.3. Allyl and benzyl halides: rate constant k_h and half-life $t_{½}$ at 298K and pH 7.

RX	k_h/s^{-1} [a]	$t_{½}$
CH_2CHCH_2Cl	1.157(-7)	69 days
CH_2CHCH_2Br	1.674(-5)	12 h
CH_2CHCH_2I	4.01(-6)	2.0 days
$C_6H_5CH_2Cl$	1.28(-5)	15 h
$p\text{-}CH_3C_6H_4CH_2Cl$	4.5(-4) [b]	0.43 h
$C_6H_5CHCl_2$	1.56(-3) [b]	0.1 h
$C_6H_5CCl_3$	6.3(-2) [b]	19 s
$C_6H_5CH_2Br$	1.45(-4) [b]	1.32 h
$p\text{-}CH_3C_6H_4CH_2Br$	2.67(-3) [b]	4.3 min

[a] Assume $k_h = k_N$; no evidence for base catalyzed reactions. See table 5.1.

[b] Assume that k_h changes by a factor of 1.9 for each 5K change in temperature based on temperature coefficient for benzyl chloride.

Table 5.4. Polyhalomethanes: k_h and $t_{½}$ at 298K[a] and pH 7.

CH_nX_{4-n}	k_h/s^{-1} [a]	$t_{½}$/yr
CH_2Cl_2	3.2(-11)	704
CH_2BrCl	5.0(-10)	44
CH_2Br_2	1.2(-10)	183
$CHCl_3$	6.9(-12)	3500
$CHBrCl_2$	1.6(-10)	137
$CHBr_2Cl$	8.0(-11)	274
$CHBr_3$	3.2(-11)	686
$CHICl_2$	8.0(-11)	275
$CHFICl$	2.2(-8)	1.0
CCl_4	4.8(-7) [b]	7000 (1 ppm) 7 (1000 ppm)

[a] Assume $k_h = k_B[OH]$ except for dihalomethanes.

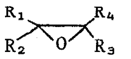
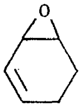
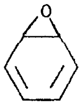
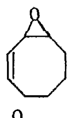
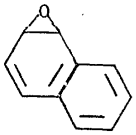
[b] Hydrolysis is second order in CCl_4 ($k_h/M^{-1}s^{-1}$); half-life is instantaneous value at stated concentration.

5.2. Epoxides

Epoxides hydrolyse rapidly at 298 K and pH 7 by a mostly water (neutral) promoted reaction to which the acid catalysed process contributes less than 10 percent of the rate; base catalysis is not detectable at pH < 10.

Table 5.5 summarizes data for 14 epoxides. The half-lives of most of these epoxides are surprisingly similar: 14.6 ± 4.6 days. Major exceptions are conjugated epoxides and aromatic epoxides which have half-lives of a few minutes.

Table 5.5. Epoxides: rate constants k_A, k_N, k_h, and $t_{½}$ at 298K, pH 7 and I_{AN} and I_{NB}.

Epoxide 									
R_1	R_2	R_3	R_4	$k_A/M^{-1}s^{-1}$	k_N/s^{-1}	k_h/s^{-1}	$t_{½}$	I_{AN}	I_{NB}
H	H	H	H	$1(-2)^a$	$6.7(-7)^a$	6.9(-7)	12 days	4.2	11.8
Me	H	H	H	$4.6(-2)^a$	$5.5(-7)^a$	5.5(-7)	14.6 days	4.9	11.7
Me	Me	H	H	7.3^a	$11(-7)^a$	18.3(-7)	4.4 days	6.8	11.8
CH_2OH	H	H	H	$2.5(-3)^c$	$2.84(-7)^c$	2.84(-7)	28 days	3.9	--
CH_2Cl	H	H	H	$8.0(-4)^a$	$9.8(-7)^c$	9.8(-7)	8.2 days^c	2.9	--
CH_2Br	H	H	H	$6.1(-4)^b$	$5(-7)^d$	5(-7)	16 days	3.1	--
CH_2OH	Me	H	H	1.1(-2)	$5(-7)^d$	5(-7)	16 days	4.4	--
CH_2Cl	Me	H	H	$1.84(-3)^b$	$5(-7)^d$	5(-7)	16 days	3.6	--
Me	H	Me	H	0.120^b	$5(-7)^d$	5.12(-7)	15.7 days	5.4	--
H	Me	Me	H	0.240^b	$5(-7)^d$	5.24(-7)	15.3 days	5.7	--
				1.64(4)	2.20(-4)	1.86(-3)	6 min	7.9	13
				$3.2(1)^e$	$1.40(-3)^e$	1.4(-3)	8 min^e	4.4	--
				3.55(-3)	2.20(-4)	2.2(-4)	52 min	4.2	--
				$4.7(2)^e$	$3(-3)^e$	3(-3)	4 min^e	5.2	--

[a] Data at 298K were selected from table 4.7 on the assumption that most recent published data is most accurate.

[b] Calculated from data in table 4.7, assuming that E = 75.3 kJ/mol; k_A at 298K is then 16 times that at 273K, and 1.68 times that at 293K.

[c] At 293K.

[d] k_N in table 4.7 ranged from 3(-7) to 11(-7) at temperatures 293K 298K; we assume k_N = 5(-7) where no k_N available.

[e] At 303K.

5.3. Esters

5.3.1. Aliphatic Acid Esters

Table 5.6 summarizes data for hydrolysis of aliphatic esters at 298 K and pH 7. Available data indicate that the base catalyzed process is dominant for simple esters. Thus, the values of k_B is a reliable guide to their lifetime. Generally, the more rapidly hydrolyzed esters have rate terms k_N and k_B [OH] that are competitive at pH 7; therefore, values of both rate constants are needed to estimate persistence. Simple esters are resistant to hydrolysis; halogenated acid esters are much more reactive and hydrolyze in only a few hours or days.

Table 5.6. Aliphatic acid esters: k_A, k_N, k_B, k_h and $t_{½}$ at 298K and pH 7, and I values.

$R_1C(O)OR_2$									
R_1	R_2	$k_A[H^+]/s^{-1}$[a]	k_N/s^{-1}	$k_B[OH^-]/s^{-1}$[a]	k_h/s^{-1}	$t_{½}$	I_{AN}	I_{AB}	I_{NB}
Me	Et	1.1(-11)	1.5(-10)	1.1(-8)	1.1(-8)	2.0 yr	(5.9)	5.5	(5.2)
Me	i-Pr	6.0(-12)	--	2.6(-9)	2.6(-9)	8.4 yr	--	5.7	--
Me	t-Bu	1.3(-11)	--	1.5(-10)	1.6(-10)	140 yr	--	6.5	--
Me	CH_2CH	1.4(-11)	1.1(-7)	1.0(-6)	1.1(-6)	7.3 days	3.1	(4.6)	6.0
Me	$CHCCH_2$	--	--	7.3(-8)	7.3(-8)	110 days	--	--	--
Me	$C_6H_5CH_2$	1.1(-11)	--	2.0(-8)	2.0(-8)	1.10 yr	--	5.4	--
Me	C_6H_5	7.8(-12)	6.6(-8)	1.4(-7)	2.1(-7)	38 days	3.1	(5.4)	6.7
Me	$2,4-(NO_2)C_6H_3$	--	1.1(-5)	9.4(-6)	2.0(-5)	9.4 h	--	--	7.1
$ClCH_2$	Me	8.5(-12)	2.1(-7)	1.4(-5)	1.4(-5)	14 h	2.6	(3.9)	5.2
Cl_2CH	Me	2.3(-11)	1.5(-5)	2.8(-4)	3.0(-4)	38 min	1.2	(3.5)	5.7
Cl_2CH	C_6H_5	--	1.8(-3)	1.3(-3)	3.1(-3)	3.7 min	--	--	7.1
F_2CH	Et	--	5.7(-5)	4.5(-4)	5.1(-4)	23 min	--	--	--
Cl_3C	Me	--	7.7(-4)	--	$\geq$ 7.7(-4)	$\leq$ 15 min	--	--	--
F_3C	Et	--	3.2(-3)	--	$\geq$ 3.2(-3)	$\leq$ 3.6 min	--	--	--
F_3C	t-Bu	--	1.3(-3)	--	$\geq$ 1.3(-3)	$\leq$ 8.9 min	--	--	--
$R_1C(O)OR_2$									
R_1	R_2	$k_A[H^+]/s^{-1}$[a]	k_N/s^{-1}	$k_B[OH^-]/s^{-1}$[a]	k_h/s^{-1}	$t_{½}$	I_{AN}	I_{AB}	I_{NB}
CH_3SCH_2	Et	--	--	9.2(-8)	9.2(-8)	87 days	--	--	--
$CH_3S(O)CH$	Et	--	--	1.3(-6)	1.3(-6)	6.2 days	--	--	--
$(CH_3)_2\overset{+}{S}CH_2$	Et	--	--	2.0(-5)	2.0(-5)	9.6 h	--	--	--
Et	Et	3.3(-12)	--	8.7(-9)	8.8(-9)	2.5 yr	--	5.3[b]	--
n-Pr	Et	1.8(-12)	--	3.8(-9)	3.8(-9)	5.8 yr	--	4.8[b]	--
i-Pr	Et	--	--	2.3(-9)	2.3(-9)	9.6 yr	--	--	--
$CH_2{=}CH$	Et	1.2(-13)	--	7.8(-9)	6.3(-9)	3.5 yr	--	4.6[b]	--
trans-$CH_3CH{=}CH$	Et	6.3(-14)	--	1.3(-9)	1.3(-9)	17 yr	--	4.8[b]	--
$H-C{\equiv}C-$	Et	--	--	4.68(-7)	4.68(-7)	17 days	--	--	--

[a] $[OH^-] = 10^{-7}M$, $[H^+] = 10^{-7}M$ at pH 7, 298K.

[b] I_{AB} is calculated from k_A at 293 K, and k_B at 298°K (see table 4.8); this affects the value of I_{AB} by less than 3 per cent.

5.3.2. Aromatic Acid Esters

Table 5.7 summarizes constants k_A and k_B, expressed as pseudo-first-order constants k[OH], k_h, $t_{1/2}$, and I_{AB}. Included in these estimates are data from experiments in mixed solvents, from which it is evident that values of k_h are greatly affected by solvent composition (a factor of 20 or more). Estimates of persistence based on mixed-solvent studies are certain to be too high, but for most simple aromatic esters lifetimes are long (over 5 years) and will remain long even in water solvent. The extensive use of these compounds and their widespread occurrence in the environment requires some reliable data on their rates of hydrolysis as a function of pH and temperature in water.

Table 5.7. Aromatic acid esters: k_A, k_B, k_h, $t_{½}$, and I_{AB} at pH 7 and 298K.

$R_1-\overset{O}{\overset{\|}{C}}OR_2$							
R_1	R_2	$k_A[H^+]/s^{-1\,a}$	$k_B[OH^-]/s^{-1\,a}$	k_h/s^{-1}	$t_{½}$/yr	Solvent^b	I_{AB}
C_6H_5	Me	4.0(-14)	1.9(-10)	1.9(-10)	118	∿ 55% M	5.2
C_6H_5	Et	--	3.0(-9)	3.0(-9)	7.3	water	--
		1.2(-14)	1.2(-10)	1.2(-10)	183	∿ 60% E	5.0
		9.9(-15)	2.8(-10)	2.8(-10)	79	∿ 60% A	4.8
		--	7.0(-10)	7.0(-10)	31	40% D	--
		--	3.3(-10)	3.3(-10)	67	70% D	--
C_6H_5	i-Pr	--	6.2(-10)	6.2(-10)	35	water	--
C_6H_5	$C_6H_5CH_2$	--	8.0(-10)	8.0(-10)	27	50 V% D	--
p-NO_2-C_6H_4	Me	4.3(-14)	7.4(-9)	7.4(-9)	3.0	60% M	4.38
		--	6.4(-8)	6.4(-8)	0.34	∿ 56 W% A	--
p-NO_2-C_6H_4	Et	1.4(-14)	2.4(-8)	2.4(-8)	0.92	∿ 60% A	3.88
1-C_5H_4N	Et	--	5.4(-8)	5.4(-8)	0.41	water	--
		--	2.0(-8)	2.0(-8)	1.1	55 W% A	--
o-$C_6H_4[C(O)OEt]_2$		--	1.0(-9)	1.0(-9)	22	50 V% D	--
			1.7(-10)	1.7(-10)			
o-$C_6H_4[C(O)OCH_2C_6H_5]_2$		--	1.7(-9)	1.7(-9)	13	50 V% D	--
			2.8(-10)	2.8(-10)			
p-$C_6H_4[C(O)OMe]_2$		--	2.5(-8)	2.5(-8)	0.88		
			1.6(-9)	1.6(-9)	0.88	60 W% D	--
p-$C_6H_4[C(O)OEt]_2$		--	6.9(-9)	6.9(-9)	3.2	50 W% A	--
			2.5(-9)	2.5(-9)	8.8		

[a] $[OH^-] = 10^{-7}M$, $[H^+] = 10^{-7}M$ at pH 7, 298K.

[b] Organic solvent legend: A = acetone; D = dioxane.

5.4. Amides

With the exception of a few halogenated acetamides, most amides hydrolyze to acids extremely slowly at 298 K and pH 7, with half-lives measured in centuries. Electronegative groups on carbon or nitrogen greatly accelerate base catalyzed hydrolysis, but alkyl groups on nitrogen retard both acid and based catalyzed processes. No neutral (water) process is evident in hydrolysis of amides, and the competition between acid and base catalyzed hydrolyses is important at pH 6–7. Data for some representative amides are summarized in table 5.8.

Table 5.8. Amides: k_A, k_B, k_h, $t_{½}$, and I_{AB} for hydrolysis at pH 7 and 298K.

Amide	$k_A/M^{-1}s^{-1}$	$k_B/M^{-1}s^{-1}$	k_h/s^{-1}	$t_{½}$/yr	I_{AB}
Acetamide	8.36(-6)	4.71(-5)	5.55(-12)	3,950	6.62
Valeramide	5.43(-6)	7.41(-5)	1.96(-12)	11,300	6.79
Isobutyramide	4.63(-6)	2.40(-5)	2.86(-12)	7,700	6.64
Cyclopentanecarboxamide	2.34(-5)	1.67(-5)	4.01(-12)	5,500	7.07
Methoxyacetamide	7.84(-6)	3.95(-4)	4.00(-11)	500	6.15
Chloroacetamide	1.1(-5)	1.5(-1)	1.5(-8)	1.46	4.93
Dichloroacetamide	--	3.0(-1)	3.0(-8)	0.73	--
Trichloroacetamide	--	9.4(-1)	9.4(-8)	0.23	--
Bromoacetamide	--	1.03(-5)	1.03(-12)	21,200	--
N-methylacetamide	3.2(-7)	5.46(-6)	5.76(-13)	38,000	6.38
N-ethylacetamide	9.36(-8)	3.10(-6)	3.10)-13)	70,000	6.23
N,N-methylethylacetamide	5.16(-7)	1.14(-5)	1.19(-12)	18,500	6.33

5.5. Carbamates

Rate constants for hydrolysis of carbamates have been measured at 298 K, and in most cases only k_B is reported; a limited number of values for k_A and k_N suggest that for many carbamates the base catalyzed process will be dominant. Table 4.11 summarizes the values for individual rate constants, which cover the extraordinary range of 10^{12}. Electronegative substituents on oxygen, such as C_6H_5 or p-$NO_2C_6H_4$, greatly accelerate the base catalyzed process, as does the presence of -NH.

Estimates of persistence of carbamates under environmental conditions, based mostly on the assumption that $k_h = k_B \times 10^{-7}/s^{-1}$, are summarized in table 5.9. Half-lives vary from 26 seconds to 240,000 years, a range that cautions against making generalizations concerning structure reactivity other than to note that purely aliphatic carbamates probably are resistant to hydrolysis under these conditions.

Table 5.9. Carbamates $R_1OC(O)NR_2R_3$: k_h, $t_{½}$, I_{AN}, and I_{NB} at pH 7 and 298K.

$R_1OC(O)NR_2R_3$			k_h/s^{-1} [a]	$t_{½}$	I
R_1O	R_2-N-R_3				
Me	C_6H_5	H	5.5(-12)	4,000 yr	--
Et	C_6H_5	Me	5.0(-13)	44,000 yr	--
C_6H_5	C_6H_5	H	5.42(-6)	1.5 days	--
C_6H_5	C_6H_5	Me	4.2(-12)	5,200 yr	--
p-$MeOC_6H_4$	C_6H_5	H	2.5(-6)	3.2 days	--
M-ClC_6H_4	C_6H_5	H	1.8(-4)	1.1 h	--
p-$NO_2C_6H_4$	C_6H_5	H	2.7(-2)	26 s	I_{NB} or $I_{AB} < 6.5$
p-$NO_2C_6H_4$	C_6H_5	Me	8.0(-11)	2,700 yr	--
1-$C_{10}H_9$	Me	H	9.4(-7)	8.5 days	$I_{AN} = 0.80$ $I_{NB} = 6.25$
1-$C_{10}H_9$	Me	Me	1.8(-11)	1,200 yr	$I_{AN} = 1.79$ $I_{NB} = 9.95$
$Et_2NCH_2CH_2$	C_6H_5	H	2.6(-12)	8,400 yr	--
$Et_2NCH_2CH_2$	$Me_3C_6H_3$	H	9.2(-14)	240,000 yr	--
$Me_3NC_6H_4$	Me	H	6.7(-6)	1.2 days	--
$Me_3NC_6H_4$	Me	Me	2.8(-11)	784 days	--
$ClCH_2CH_2$	C_6H_5	H	1.6(-10)	140 yr	--
Cl_2CHCH_2	C_6H_5	H	5.0(-9)	4.4 yr	--
CCl_3CH_2	C_6H_5	H	3.2(-8)	252 days	--
CF_3CH_2	C_6H_5	H	1.0(-8)	2.2 yr	--

[a] In the absence of data for k_A or k_N values of k_h are assumed equal to $k_B \times 10^{-7}/s^{-1}$.

5.6. Phosphoric and Phosphoric Acid Esters

5.6.1. Dialkyl Alkylphosphonates

None of the phosphonates reported in table 4.12 is hydrolyzed at significant rates at 298 K and pH 7. The dominant process is base catalyzed hydrolysis, which is extremely slow with all alkyl-substituted esters. Electron-withdrawing substituents such as p-nitrophenoxy increase the rate of the base catalyzed process by a factor of about 100, but even $MeP(O)(OEt)OC_6H_4NO_2$ has an estimated half-life of 5.5 years. Table 5.10 summarizes data for representative phosphonate esters.

Table 5.10. Phosphonic acid esters, dialkyl phosphonates, $R_1P(O)(OR_2)_2$: k_A, k_B, k_h, $t_{1/2}$, and I_{AB} at 298K and pH 7.

$R_1P(O)(OR_2)_2$		$k_A/M^{-1}s^{-1}$	$k_B/M^{-1}s^{-1}$	k_h/s^{-1}	$t_{1/2}/yr$	I_{AB}
Me	Me	∿ 1.36(-9)	2.5(-3)	2.5(-10)	88	2.8
Me	Et	1.7(-9)	2.2(-4)	2.2(-11)	990	3.0
Me	i-Pr	6.4(-9)	3.2(-7)	3.3(-14)	663,000	3.5
Me	Et, NP[a]	∿ 1.2(-7)[b]	4.0(-2)	4.0(-9)	5.5	2.9
Et	i-Pr	3.2(-9)	3.7(-8)	4.0(-15)	5.5(6)	3.6
C_6H_5	Et	1.1(-9)	> 5(-4)[c]	> 5(-11)	440	2.9

[a] p-Nitrophenyl.

[b] Assume log k_A ∿ 10.6 − (100/RT).

[c] Assume log k_B ∿ 6.6 − (56/RT).

5.6.2. Phosphoric Acid and Thiophosphoric Acid Esters

Hydrolysis rate data for phosphoric acid esters are incomplete in many ways; those esters listed in table 5.11 have been extensively investigated, and it is evident that in almost every case neutral hydrolysis is the dominant process at pH 7.

Lifetimes for phosphate esters are shorter than for phosphonates by a factor of 100 or more. Table 5.11 summarizes data for 11 phosphoric and thiophosphoric acid esters.

Table 5.11. Phosphoric acid and thiophosphoric acid esters: $(RO)_3PO$ and $(RO)_3PS$: k_N, k_B and k_h at 298K and pH 7.

$(RO)_3PO(S)$	k_N/s^{-1}	$k_B/M^{-1}s^{-1}$	k_h/s^{-1}	$t_{1/2}$
$(MeO)_3PO$	1.8(-8)	1.3(-4)	1.8(-8)	1.22 yr
$(EtO)_3PO$	4(-9)[a]	8.2(-6)	4(-9)	5.5 yr
$(EtS)_3PO$	1.4(-9)[b]	1.2(-2)	2.6(-9)	8.5 yr
$(C_6H_5O)_3PO$	2.7(-11)[b]	1.7(-2)	1.7(-9)	1.3 yr
$(EtO)P(O)(NP)_2$	3.3(-6)	5.3(-1)	3.8(-6)	2 days
$(NP)_3PO$	1.0(-3)	3.43(1)	1.0(-3)	11 min
$(MeO)_2P(S)NP$	1.1(-7)	5.95(-3)	1.1(-7)	72 days
$MeOP(S)SCHCH$ $(CO_2Et)_2$	--	4.3	> 4.3(-7)	< 18 days
$(EtO)_2P(S)NP$	∿ 3(-9)	2.2(-4)	∿ 3(-9)	7 yr

[a] Assume E_N = 136 kJ/mol based on data at 375 and 353K.

[b] Assume E_N = 95 kJ/mol.

5.6.3. Dialkyl Phosphonohalidates

Table 5.12 summarizes the available data on hydrolysis of phosphoric acid halides in water. These compounds hydrolyze via base-catalyzed reactions and have very short half-lives under environmental conditions. Exceptions are noted for alkylamino-substituted phosphonohalidates, which appear to be persistent. More detailed studies in water at moderate pH are needed to verify this conclusion.

Table 5.12. Phosphoric acid halides: dialkylphosphonohalidates and dialkylphosphorohalidates $R_1R_2P(O)X$: k_h and $t_{1/2}$ at 298K and pH 7.

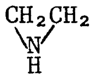
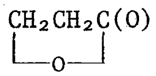
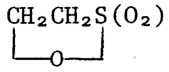
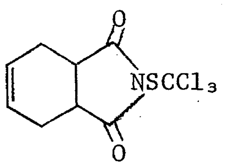
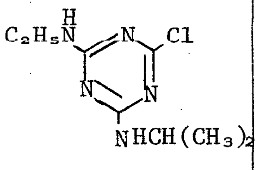
$R_1P(O)XR_2$		k_h/s^{-1}	$t_{1/2}$
X = F			
Me	Me	4(-4)	2.9 min
Et	Et	> 1.4(-6)[a]	< 5 days
Et	EtO	> 3.3(-8)[a]	< 240 days
EtO	EtO	> 1.8(-7)[a]	< 80 days
i-PrO	i-PrO	1.7(-6)	4.7 days
i-PrNH	i-PrNH	1.9(-6)	4.2 days
i-PrNH	Me_2N	8.5(-7)	9.5 days
Me_2N	Me_2N	> 4.7(-10)[a]	< 46 yr
X = Cl			
Et	Et	3.6	0.2 s
Et	MeO	3.5(-1)	2 s
MeO	MeO	8.6(-3)	1.3 min
Me	i-PrO	4.5(-4)	26 min
i-PrO	i-PrO	8.1(-3)	86 s

[a] Based on k_B as a limiting value.

5.7. Acylating and Alkylating Agents and Pesticides

Table 5.13 summarizes data for nine types of acyl and alkyl derivatives and three important pesticides. With the exception of aziridine (ethyleneimine) and methoxychlor, all the compounds have short half-lives in water, the longest being 7 h. Captan was studied both in pure and natural waters and found to have the same rate constant.

Table 5.13. Acylating and alkylating agents and pesticides: k_h and $t_{1/2}$ at 298K and pH 7.

Compound	k_h	$t_{1/2}$
 CH_2CH_2NH	5.2(-8)	154 days
 $CH_2CH_2C(O)O$	3.3(-3)	3.5 min
 $CH_2CH_2S(O_2)O$	2.15(-5)	8.9 min
$(CH_3O)_2SO_2$	1.66(-4)	1.2 min
$(ClCH_2)_2O$	2.8(-2)	25 s
$C_6H_5N{=}N{-}N(CH_3)_2$	$2.75(-5)^a$	7 h
$C_6H_5C(O)Cl$	4.2(-2)	16 s
$(CH_3)_2NCOCl$	> 2.5(-3)	< 4 min
$CH_3OC(O)Cl$	5.64(-4)	20 min
$(CH_3OC_6H_4)_2CHCCl_3$ (methoxychlor)	$5.5(-8)^b$ $3(-8)^c$	147 days 270 days
 (captan)	$6.5(-5)^d$	3 h [d]
 (atrazine)	7.6(-5)	2.5 h

[a] Assumed k_h at 298K was one-half the value at 310K. [b] Extrapolated results of Wolfe et al., [11] section 4. [c] Measured value at pH 7.1, [12], section 4. [d] Value observed at 300K in distilled and natural waters, [11], section 4.