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PHYSICOCHEMICAL FOUNDATIONS

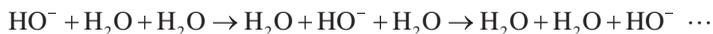
1.1 GENERALITIES

The alchemists' adage, "*Corpora non agunt nisi fluida*," "Substances do not react unless fluid," is not strictly accurate, for crystals can be transformed by processes of nucleation and growth. There is growing interest in "mechanochemical" processes, which are carried out by grinding solid reagents together (and which no doubt involve a degree of local melting). Nevertheless, it is still generally true enough to be worthy of attention. Seltzer tablets, for instance, must be dissolved in water before they react to evolve carbon dioxide. The "fluid" state may be gaseous or liquid, and the reaction may be a homogeneous one occurring throughout a single gas or liquid phase, or a heterogeneous one occurring only at an interface between a solid and a fluid, or at the interface between two immiscible fluids. As the title suggests, this book is concerned mainly with homogeneous reactions, and will emphasize reactions of substances dissolved in liquids of various kinds.

The word "solvent" implies that the component of the solution so described is present in excess; one definition is "the component of a solution that is present in the largest amount." In most of what follows it will be assumed that the solution is dilute. We will not attempt to define how dilute is "dilute," except to note that we will routinely use most physicochemical laws in their simplest available forms, and then require that all solute concentrations be low enough that the laws are valid, at least approximately.

Of all solvents, water is of course the cheapest and closest to hand. Because of this alone it will be the solvent of choice for many applications. In fact, it has dominated our thinking for so long that any other solvent tends to be tagged *nonaqueous*, as if water were in some essential way unique. It is true that it has an unusual combination of properties (see, e.g., Marcus, 1998, pp. 230–232). One property in which it is nearly unique is a consequence of its ability to act both as an acid and as a base.

That is the enhanced apparent mobility of the H_3O^+ and HO^- ions, explained by the Grotthuss mechanism (Cukierman, 2006; de Grotthuss, 1806):



in which protons hop from one molecule or ion to the next following the electric field, without actual motion of the larger ion through the liquid. This property is shared (in part) with very few solvents, including methanol and liquid hydrogen fluoride, but not liquid ammonia, as may be seen from the ionic equivalent conductances (see Table 1.1). It is apparent that in water, both the positive and negative ions are anomalously mobile. In ammonia neither is, in hydrogen fluoride only the negative ion is, and in methanol only the positive ion is.

As aqueous solution of an acid is diluted by addition of a solvent that does not contribute to the hydrogen-bonded network, the Grotthuss mechanism becomes less effective. For an electrolyte that conducts electricity by migration of ordinary ions through the solvent, Walden observed that the product of the limiting equivalent conductance of the electrolyte with the viscosity in different solvent or mixtures of different composition is approximately constant. The limiting equivalent conductance of HCl in several dioxane/water mixtures was measured by Owen and Waters (1938). As can be seen in Figure 1.1, in 82% dioxane the Walden product drops to hardly a quarter of its maximum. The Grotthuss mechanism is largely suppressed.

More and more, however, other solvents are coming into use in the laboratory and in industry. Aside from organic solvents such as alcohols, acetone, and hydrocarbons, which have been in use for many years, industrial processes use such solvents as sulfuric acid, hydrogen fluoride, ammonia, molten sodium hexafluoroaluminate (cryolite), various other “ionic liquids” (Welton, 1999), and liquid metals. Jander and Lafrenz (1970) cite the industrial use of bromine to separate caesium bromide (sol’ y 19.3 g/100 g bromine) from the much less soluble rubidium salt. The list of solvents available for preparative and analytical purposes in the laboratory now is long and growing, and though water will still be the first solvent that comes to mind, there is no reason to stop there.

TABLE 1.1 Limiting Equivalent Conductances of Ions in Amphiprotic Solvents

In H_2O at 25°C		In NH_3 at -33.5°C^a		In HF at 20°C^b		In MeOH at 25°C^c	
H_3O^+	349.8	NH_4^+	131	H_2F^+	102	MeOH_2^+	141.8
HO^-	198.5	NH_2^-	133	HF_2^-	350	MeO^-	53.02
Na^+	50.11	Na^+	130	Na^+	150	Na^+	45.5
K^+	73.52	K^+	168	K^+	150	K^+	53.6

^aKraus and Brey (1913).

^bKilpatrick and Lewis (1956).

^cOgston (1936), Conway (1952, pp. 155, 162).

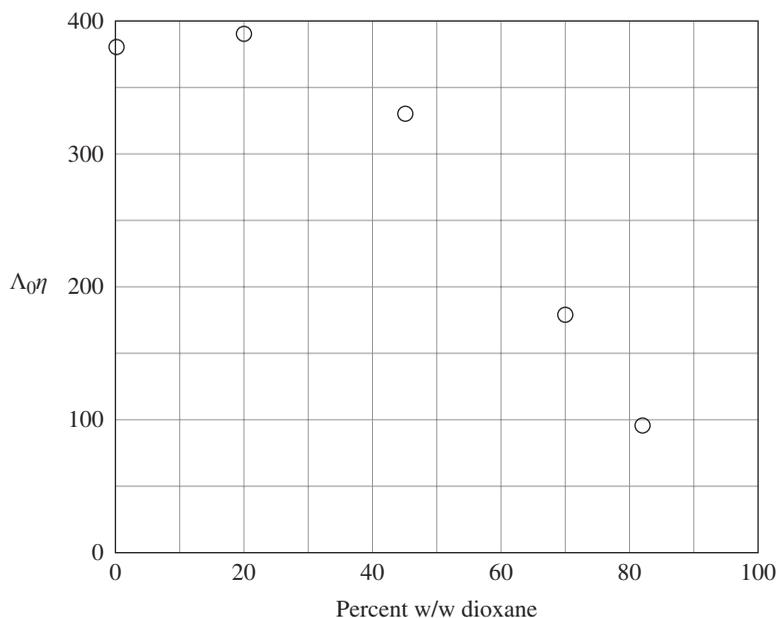
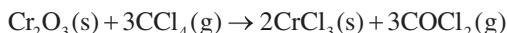


FIGURE 1.1 The Walden product, $\Lambda_0\eta$, for HCl in 1,4-dioxane/water mixtures versus percentage of dioxane at 25°C. Data from Owen and Waters (1938).

After the first observation of the effect of solvent change on reaction rate by Berthelot and Pean de St. Gilles (1862) and the first systematic study, Menshutkin (1887, 1890), the study of solvent effects was for some years largely the work of physical-organic chemists. The pioneer in this growing field was Hammett and Deyrup (1932, and see his book, *Physical Organic Chemistry*, 1970). The study of solvent effects was pursued notably by Hughes and Ingold (1935) and Grunwald and Winstein (1948). One of us (R. A. S.) was privileged to attend Ingold's lectures at Cornell that became the basis of his book (Ingold, 1969), while E. B. can still recall vividly the undergraduate lectures by both Hughes and Ingold on the effect of solvent in nucleophilic substitution: the Hughes-Ingold Rules (Ingold, 1969). Inorganic chemists soon followed. Tobe and Burgess (1999, p. 335) remark that while inorganic substitution reactions of known mechanism were used to probe solvation and the effects of solvent structure, medium effects have been important in understanding the mechanisms of electron transfer.

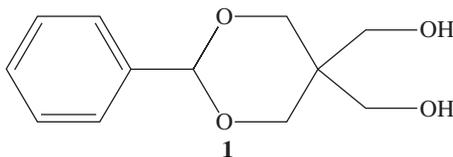
If a solvent is to be chosen for the purpose of preparation of a pure substance by synthesis, clearly the solvent must be one that will not destroy the desired product, or transform it in any undesirable way. Usually it is obvious what must be avoided. For instance, one would not expect to be able to prepare a strictly anhydrous salt using water as the reaction medium. Anhydrous chromium (III) chloride must be prepared by some reaction that involves no water at all, neither in a solvent mixture nor in any of the starting materials, nor as a by-product of reaction. A method that works uses

the reaction at high temperature of chromium (III) oxide with tetrachloromethane (carbon tetrachloride), according to the equation:



Here no solvent is used at all.¹ Some other anhydrous salts may be prepared using such solvents as sulfur dioxide, dry diethyl ether (a familiar example is the Grignard reaction, in which mixed halide–organic salts of magnesium are prepared as intermediates in organic syntheses), hydrogen fluoride, and so on.

A more subtle problem is to maximize the yield of a reaction that could be carried out in any of a number of media. Should a reaction be done in a solvent in which the desired product is most or least soluble, for instance? The answer is not immediately clear. In fact one must say, “It depends....” If the reaction is between ions of two soluble salts, the product will precipitate out of solution if it is insoluble. For example, a reaction mixture containing barium, silver, chloride, and nitrate ions will precipitate insoluble silver chloride if the solvent is water, but in liquid ammonia the precipitate is barium chloride. Another example, from organic chemistry, described by Collard *et al.* (2001) as an experiment suitable for an undergraduate laboratory, is the dehydrative condensation of benzaldehyde with pentaerythritol in aqueous acid to yield the cyclic acetal, 5,5-bis(hydroxymethyl)-2-phenyl-1,3-dioxane, **1**:



At 30°C the product is sufficiently insoluble to appear as a precipitate, so the reaction proceeds in spite of the formation of water as by-product. On the other hand, we will show in Chapter 2 that, in a situation where all the substances involved in a reaction among molecules are more or less soluble, *the most soluble substances will be favored at equilibrium.*

1.2 CLASSIFICATION OF SOLVENTS

Solvents may be classified according to their physical and chemical properties at several levels. The most striking differences among liquids that could be used as solvents are observed between *molecular liquids*, *ionic liquids* (molten salts or salt mixtures, room-temperature ionic liquids), and *metals*. They can be considered as extreme types, and represented as the three vertices of a triangle (Trémillon, 1974) (see Fig. 1.2). Intermediate types or mixtures can then be located along edges or within the triangle. The room-temperature ionic liquids (see later, Section 8.3), which

¹Caution: The reagent tetrachloromethane and the by-product phosgene are toxic and environmentally undesirable.

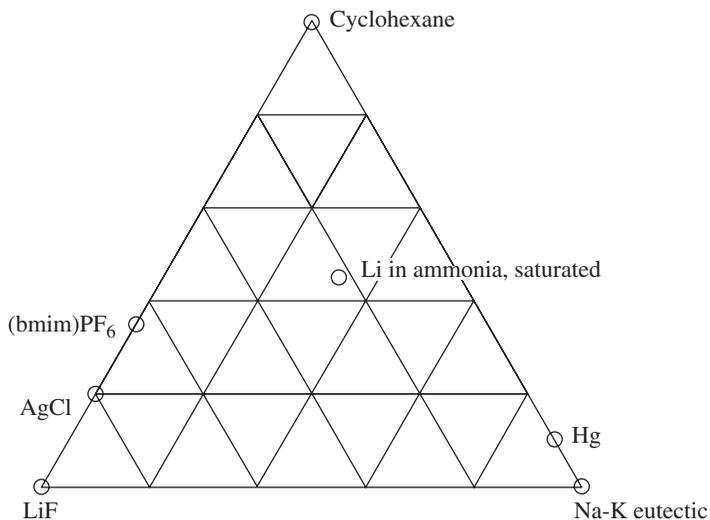


FIGURE 1.2 Ternary diagram for classification of liquids (schematic; location of points is conjectural); [bmim]PF₆ represents a room-temperature ionic liquid (see Section 8.3). After Trémillon (1974).

typically have large organic cations and fairly large anions, lie along the molecular-ionic edge, for instance.

Among the molecular liquids, further division based on physical and chemical properties leads to categories variously described (Barthel and Gores, 1994; Reichardt and Welton, 2011) as **inert** (unreactive, with low or zero dipole moments and low polarizability), **inert-polarizable** (e.g., aromatics, polyhalogenated hydrocarbons), **protogenic** (hydrogen-bonding proton donors, HBD), **protophilic** (hydrogen-bonding proton acceptors, HBA), **amphiprotic** (having both HBD and HBA capabilities), and **dipolar-aprotic** (having no marked HBD or HBA tendencies, but possessing substantial dipole moments). Examples of these classes are listed in Table 1.2. The ability of solvent molecules to act as donors or acceptors of electron pairs, that is, as Lewis bases or acids, complicates the classification. Nitriles, ethers, dialkyl sulfides, and ketones are electron-pair donors (EPD), for example; sulfur dioxide and tetracyanoethene are electron-pair acceptors (EPA). EPD and EPA solvents can be further classified as *soft* or *hard*. (Classifying can be habit-forming.) Pushing the conditions can cause normally inert substances to show weak prototropic properties: dimethyl sulfoxide can lose a proton to form the *dimsyl* ion, CH₃SOCH₂⁻, in very strongly basic media (Olah *et al.*, 1985). An equilibrium concentration of *dimsyl* ion, very small, though sufficient for hydrogen-deuterium isotopic exchange to occur between dimethyl sulfoxide and D₂O, is set up even in very dilute aqueous NaOH (Buncel *et al.*, 1965). Carbon monoxide, not normally considered a Brønsted base, can be protonated in the very strongly acidic medium of HF-SbF₅ (de Rege *et al.*, 1997).

TABLE 1.2 Molecular Solvents

Classes	Examples
Inert	Aliphatic hydrocarbons, fluorocarbons
Inert-polarizable	Benzene (π -EPD), tetrachloromethane, carbon disulfide, tetracyanoethene (π -EPA)
Protogenic (HBD)	Trichloromethane
Protophilic (HBA)	Tertiary amines (EPD)
Amphiprotic	Water, alcohols; ammonia is more protophilic than protogenic, while hydrogen fluoride is the reverse
Dipolar-aprotic	Dimethylformamide, acetonitrile (EPD, weak HBA), dimethyl sulfoxide, hexamethylphosphortriamide

1.3 SOLVENTS IN THE WORKPLACE AND THE ENVIRONMENT

The majority of solvents must be considered as toxic to some degree. Quite aside from those that have specific toxicity, whether through immediate, acute effects, or more insidiously as, for instance, carcinogens the effects of which may take years to manifest, all organic substances that are liquid at ordinary temperatures and are lipophilic (fat-soluble) are somewhat narcotic. The precautions that should be taken depend very much on their individual properties. Inhalation of vapors should always be avoided as much as possible. Many solvents are quickly absorbed through the skin. Use of an efficient fume hood is always advisable. Protective gloves, clothing, masks, and so on, should be available and used as advised by the pertinent literature (in Canada, the Material Safety Data Sheet). The rare solvents that exhibit extreme toxicity, such a liquid HCN or HF, require special precautions. The latter is an example of substances absorbed rapidly through the skin, with resulting severe burns and necrosis. Most common solvents are inflammable to varying degrees.² Those with low boiling points or low flash points (see Table A.1) require special precautions. A few have in addition particularly low ignition temperatures; a notable example is carbon disulfide, the vapor of which can be ignited by a hot surface, without a flame or spark. Transfer of a solvent with low electrical conductivity from a large shipping container to a smaller, ready-use container can be associated with an accumulation of static charge, with the chance that a spark may occur, causing fire. Proper grounding of both containers can prevent this.

Environmental concerns include toxicity to organisms of all sorts, but perhaps more importantly the tendency of each substance to persist and to be transported over long distances. Chemical stability may seem to be a desirable property, but unless a solvent is biodegradable or easily decomposed photochemically by sunlight, it can become a long-lasting contaminant of air, water, or soil, with consequences that we

²In 1978, the Canadian Transportation of Dangerous Goods Code was modified to require that labels on goods that burn easily are to use the word "inflammable" only (Johnstone, 1978; Stairs, 1978b).

probably cannot foresee. Much effort is currently going into the consideration of the long-term effects of industrial chemicals, including solvents, should they escape.

For these reasons, selection of a solvent should always be made with an eye on the effects it might have if it is not kept to minimum quantities and recycled as much as possible. Consideration should also be given to the history of the solvent before it reaches the laboratory. Does its manufacture involve processes that pose a danger to the workers or to the environment? These matters are discussed further in Section 8.6.

1.4 SOME ESSENTIAL THERMODYNAMICS AND KINETICS: TENDENCY AND RATE

How a particular reaction goes or does not go in given circumstances depends on two factors, which may be likened, “psychochemically” speaking, to “wishing” and “being able.”³ The first is the **tendency to proceed**, or the degree to which the reaction is out of equilibrium, and is related to the equilibrium constant and to free energy changes (Gibbs or Helmholtz). It is the subject of **chemical thermodynamics**. The second is the speed or **rate** at which the reaction goes, and is discussed in terms of rate laws, mechanisms, activation energies, and so on. It is the subject of **chemical kinetics**. We will need to examine reactions from both points of view, so the remainder of this chapter will be devoted to reviewing the essentials of these two disciplines, as far as they are relevant to our needs. The reader may wish to consult, for example, Atkins and de Paula (2010), for fuller discussions of relevant thermodynamics and kinetics.

1.5 EQUILIBRIUM CONSIDERATIONS

For a system at constant pressure, which is the usual situation in the laboratory when we are working with solutions in open beakers or flasks, the simplest formulas to describe equilibrium are written in terms of the Gibbs energy G , and the enthalpy H . For a reaction having an equilibrium constant K at the temperature T , one may write:

$$\Delta G_{\text{T}}^0 = -RT \ln K \quad (1.1)$$

$$\Delta H_{\text{T}}^0 = -R \left(\frac{\partial \ln K}{\partial (1/T)} \right)_p \quad (1.2)$$

The equilibrium constant K is of course a function of the activities of the reactants and products, for example, for a reaction: $A + B \rightleftharpoons Y$

³There is a word, very pleasing to us procrastinators, “velleity,” which is defined (Fowler *et al.*, 1976) as “low degree of volition not prompting to action.” See also Nash (1938).

$$K = \frac{a_Y}{a_A \cdot a_B} \quad (1.3)$$

By choice of standard states one may express the activities on different scales. For reactions in the gas phase, it is convenient, and therefore common, to choose a standard state of unit activity on a scale of pressure such that the limit of the value of the dimensionless activity coefficient, $\gamma = a_i/P_i$, as the pressure becomes very low, is unity. The activity on this scale is expressed in pressure units, usually atmospheres or bars, so we may write

$$K = \frac{\gamma_Y P_Y}{\gamma_A P_A \gamma_B P_B} = \Gamma_\gamma K_p \quad (1.4)$$

The activity coefficient quotient Γ_γ is unity for systems involving only ideal gases, and for real gases at low pressure.

For reactions involving only condensed phases, including those occurring in liquid solutions, which are our chief concern, the situation is very different. Three choices of standard state are in common use. For the solvent (i.e., the substance present in largest amount), the standard state almost universally chosen is the pure liquid. This choice is also often made for other liquid substances that are totally or largely miscible with the solvent. The activity scale is then related to the mole fraction, through the *rational activity coefficient* f , which is unity for each pure substance. For other solutes, especially those that are solid when pure, or for ionic species in solution in a nonionic liquid, activity scales are used that are related either to the molar concentration or the molality, depending on experimental convenience. On these scales, the activity coefficients become unity in the limit of low concentration.

If a substance present in solution is to some extent volatile, that is, if it exerts a measurable vapor pressure, its activity in solution can be related to its activity in the gas (vapor) phase. If the solution is ideal, all components obey Raoult's Law, expressed by Equation 1.5, and illustrated by the dashed lines in Figure 1.3.

$$p_i = p_i^0 x_i \quad (1.5)$$

Here p_i is the vapor pressure of the i th substance over the solution, p_i^0 is the vapor pressure it would exert in its standard (pure liquid) state, and x_i is its mole fraction in the solution. We can now define an "absolute" activity (not really absolute, but relative to the gas phase standard state on the pressure scale as earlier) measured by p_i , assuming that the vapor may be treated as an ideal gas or by the *fugacity*⁴ if necessary. We shall always make the "*ideal gas*" assumption, without restating it.

⁴Fugacity f is pressure corrected for nonideality. It is defined so that the Gibbs energy change on isothermal, reversible expansion of a mole of a real gas is $\Delta G = \int V dp = RT \ln(f/f_0)$. For a real gas at low enough pressures, $f = P$. Fugacities can be calculated from the equation of state of the gas if needed. See any physical chemistry textbook, for example, Atkins and de Paula (2010, pp. 129–130). For an only slightly nonideal gas $f = P^2 V_m / RT$, approximately.

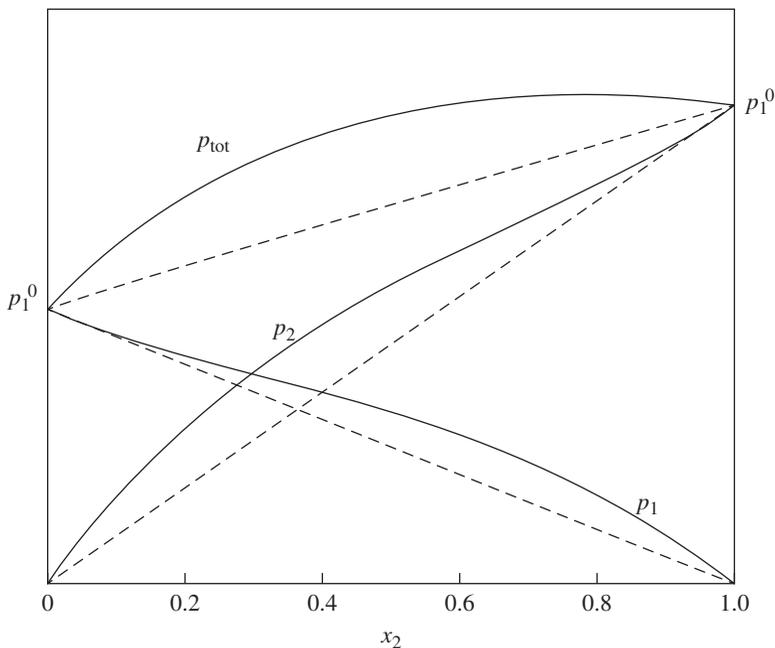
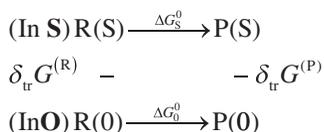


FIGURE 1.3 Vapor pressure over binary solutions. Dashed lines: ideal (Raoult’s Law). Solid curves: positive deviations from Raoult’s Law. Note that where $x_2 \ll 1$, P_1 is close to ideal, and vice versa.

1.6 THERMODYNAMIC TRANSFER FUNCTIONS

The thermodynamic equilibrium constant as defined earlier is independent of the solvent. The practical equilibrium constant is not, because the activity coefficients of the various reactant and product species will change in different ways when the reaction is transferred from one solvent to another. One way of considering these changes is through the use of thermodynamic transfer functions. The standard Gibbs energy of a reaction in a solvent **S**, ΔG_S^0 , may be related to that in a reference solvent **O**, ΔG_O^0 , by considering the change in Gibbs energy on transferring each reactant and product species from the reference solvent to **S**. The reference solvent may be water or the gas phase (no solvent). Other functions (enthalpy, entropy) can be treated in the same fashion as G . A reaction converting reactants **R** to products **P** in the two solvents can be represented in a Born–Haber cycle:



$$\Delta G_S^0 = -\delta_{\text{tr}} G^{(R)} + \Delta G_O^0 + \delta_{\text{tr}} G^{(P)} \quad (1.6)$$

For each participating substance **I**, the term $\delta_{\text{tr}} G^{(I)}$ can be obtained from vapor pressure, solubility, electrical potential, or other measurements that enable the calculation of activity coefficients and hence of standard Gibbs energies, using Equation 1.7.

$$\delta G_{\text{tr}}^{(i)} = G_S^{0(i)} - G_O^{0(i)} \quad (1.7)$$

Since the Gibbs energy and the activity coefficient are related through Equation 1.8, this development could have been carried out in terms of $\ln a$ or $\ln f$.

$$G_S^{0(i)} - G_O^{0(i)} = RT \ln \left(\frac{a_S^{0(i)}}{\alpha_O^{0(i)}} \right) = RT \ln \left(\frac{f_S^i}{f_O^i} \right) \quad (1.8)$$

Because of the analogy between the transition states in kinetics and the products in equilibrium (see later, Section 1.6), similar considerations can be applied to the understanding of solvent effects on reaction rates. This will be illustrated in Chapter 6.

1.7 KINETIC CONSIDERATIONS: COLLISION THEORY

Elementary reactions occurring in the gas phase have been fruitfully discussed in terms derived from the Kinetic–Molecular Theory of Gases. The result is Equation 1.9,

$$\text{Rate} = PZ_0 [A][B] e^{-E_a/PT} \quad (1.9)$$

$$Z_0 = \pi d^2 \left(\frac{8k_B T}{\pi \mu} \right)^{1/2} N^2 \quad (1.10)$$

where Z_0 is the number of collisions per unit time between A and B molecules at unit concentrations given by Equation 1.10, $[A]$ and $[B]$ represent the concentrations of the reacting species, d is the mean diameter of A and B, k_B is the Boltzmann constant, and μ their reduced mass, and E_a is the activation energy. P is the steric or probability factor, that is, the probability that the colliding molecules are in suitable orientations and internal configuration to permit reaction, as illustrated in Figure 1.4. The factors PZ_0 are usually combined to form the *Arrhenius pre-exponential factor*, usually denoted by A . Equations 1.9 and 1.10 have allowed a substantial level of understanding of simple reactions to be achieved, and by combining elementary steps into multistep mechanisms, complex reactions may also be described. This simple Arrhenius treatment is not applicable to reaction in solution, however, so for our purposes another approach is needed.

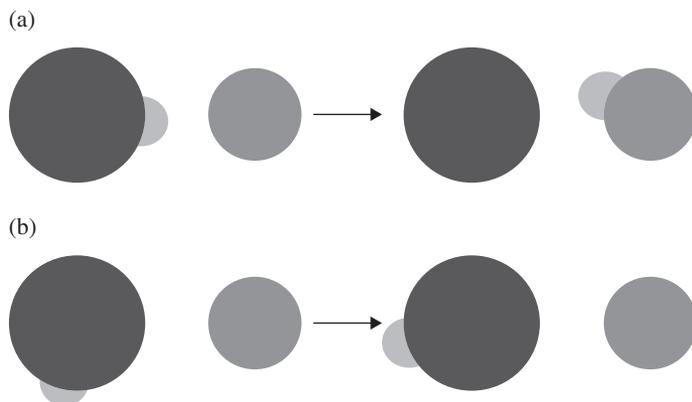


FIGURE 1.4 Successful (a) and unsuccessful (b) transfer of a hydrogen atom from HI to Cl.

1.8 TRANSITION-STATE THEORY

The variously named *transition-state theory* (the preferred name) or *absolute reaction rate theory*, developed by Eyring and associates (Berry *et al.*, 2000, pp. 911–927; Eyring, 1935; Laidler and Meiser, 1995, pp. 382–387) and by Evans and Polanyi (1935), takes a quite different view. The reacting molecules are considered as entering a “transition state,” forming an “activated complex,” which resembles an ordinary molecule in all respects but one, which is that one of its normal modes of vibration is *not* a vibration, because there is no restoring force; rather it will lead to decomposition of the complex, either to form the products of the reaction or to reform the starting molecules. Quantum–mechanical calculations of the energetics and geometry of molecules in configurations that represent transition states can be carried out using such computer programs as GAUSSIAN, SPARTAN, or HYPERCHEM (Levine, 2013). Of the normal modes of vibration of such a transition-state “molecule,” one has a *negative force constant*. What is meant by this is that there is no force restoring the molecule to an equilibrium configuration in the direction of this motion; in fact the force is repulsive, leading to rearrangement or decomposition, to form the products of the reaction, or to reform the starting molecules. Since the force constant is negative, the frequency, which depends on the square root of the force constant, contains the factor $\sqrt{-1}$; that is, it is imaginary. A graph of the energy of the system as a function of the normal coordinates of the atoms (the *potential energy surface*) in the vicinity of the transition state takes the form of a *saddle* or *col*, illustrated in Figure 1.5.

From the saddle point, the energy increases in all the principal directions except along the direction that leads to reaction (forward) or (backward) to reform the starting materials. The course of a simple reaction may be represented as motion along the *reaction coordinate*, which is a combination of atomic coordinates leading from the initial configuration (reactants) through the transition state to the final

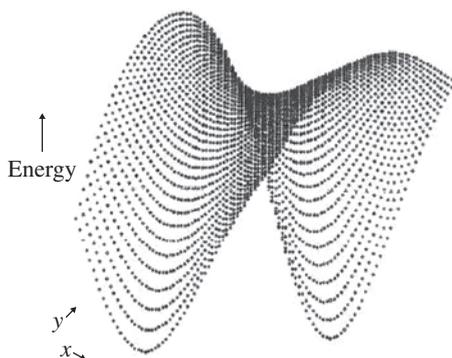


FIGURE 1.5 A portion of a potential-energy surface $E(x,y)$, showing a saddle point.

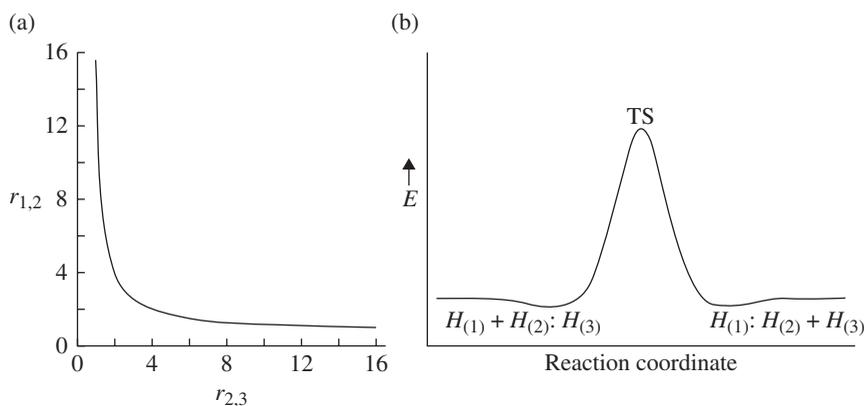


FIGURE 1.6 (a) The reaction pathway of least energy and (b) the profile along the pathway, for the hydrogen atom–molecule exchange reaction (schematic).

configuration (products) along, or nearly along, the path of least energy. Figure 1.6a shows a projection of the path of least energy on the potential energy surface for a very simple reaction, in which a hydrogen atom attacks a hydrogen molecule directly at one end, and one atom is transferred to the attacking atom. The reaction coordinate is measured along the (approximately hyperbolic) pathway. The energy as a function of the reaction coordinate is shown in Figure 1.6b.

In most reactions, especially those taking place in solution, the situation is more complicated. For instance, Figure 1.7 shows a possible form of the energy profile for a reaction in which one ligand in a transition-metal complex ion is replaced by another. In the scheme here, M represents a trivalent metal ion. There may first be formed an outer-sphere complex, perhaps an ion pair (1.11), which then rearranges (1.12) so the arriving and leaving ligands change places (not necessarily with retention of configuration). The leaving ligand, now in the outer sphere, finally leaves (1.13).

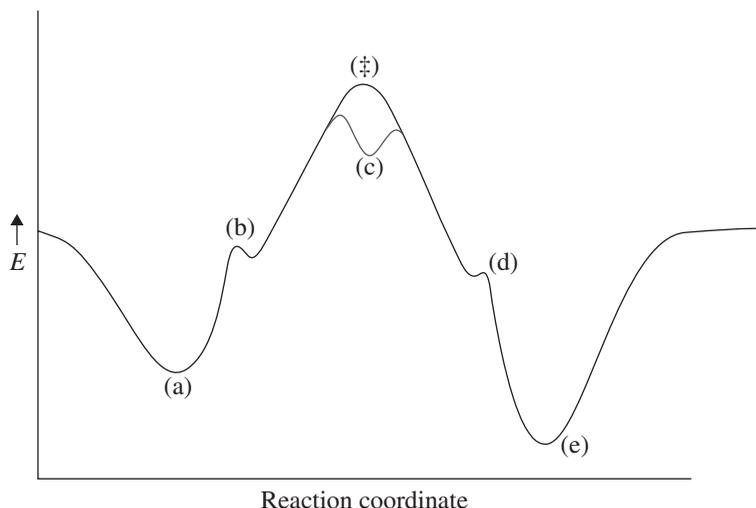
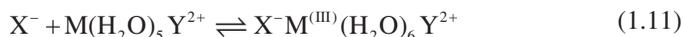
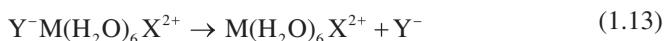


FIGURE 1.7 A possible, more realistic reaction profile for a ligand-exchange reaction, showing reactants (a), precursor (b), and successor (d) complexes, the transition state (\ddagger), the possibility of the formation of a reactive intermediate (c), and products (e). Redrawn after Kettle (1996) by permission of Oxford University Press.



(The seven-coordinate species may be a true activated complex, corresponding to the curve with a single maximum in Figure 1.7, or a transient intermediate, corresponding to the light curve with a dip between two maxima.)



The composition of the activated complex may be deduced from the rate law. In the case of a multistep reaction, if one step is rate controlling, which is usually true, the composition of the activated complex of the rate-controlling step may still be deduced from the rate law for the overall reaction. For example, if a reaction between two substances A and B follows a rate law of the form of Equation 1.14 (over certain ranges of concentrations and temperatures):

$$\text{Rate} = k[A]^n[B]^m \quad (1.14)$$

the activated complex has a composition represented by $A_n B_m$. (There are some subtle aspects of this rule; e.g., see Problem 2.2.) We know nothing, however, of its structure nor of the steps in the reaction, in the absence of other evidence.

Nevertheless, one may write a statement resembling an equilibrium-constant expression, relating the activities or, approximately, the concentrations of the reactants and the activated complex (represented by the “double-dagger” or “Cross of Lorraine” symbol †):

$$K^\ddagger = \frac{[\ddagger]}{[A]^n[B]^m} \quad (1.15)$$

Then, if we assume that the rate of decomposition of the complex is first order, that is, that it reacts to form the products at a rate proportional to its concentration, we obtain:

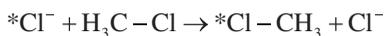
$$\text{Rate} \propto [\ddagger] \propto [A]^n[B]^m \quad (1.16)$$

That is, the ordinary rate constant, k , is proportional to K^\ddagger .

Specifically, one may write (Atkins and de Paula, 2010, p. 846; Laidler and Meiser 1995, p. 741):

$$k = \kappa \frac{k_B T}{h} K^\ddagger \quad (1.17)$$

where k_B is the Boltzmann’s constant, h is the Planck’s constant, and T is the absolute temperature. The group $k_B T/h$ has a value of about 6 ps^{-1} at 298 K. The factor κ is a constant, the *transmission coefficient*, the value of which is close to unity for bimolecular reactions in the gas phase. Abboud *et al.* (1993, p. 75) cite a computer simulation study (Wilson, 1989) of the chloride exchange reaction:



in which the transmission coefficient was calculated to be unity for the reaction in the gas phase, but 0.55 in aqueous solution, apparently owing to confinement of the reacting species within a “cage” of water molecules, so that multiple crossings of the transition barrier can occur.

To the extent that K^\ddagger can be considered an ordinary equilibrium constant, one may then apply the usual thermodynamic relations, that is,

$$\Delta_\ddagger G^0 = -RT \ln K^\ddagger \quad (1.18)$$

$$\Delta_\ddagger H^0 = -R \left(\frac{\partial \ln K^\ddagger}{\partial (1/T)} \right)_p \quad (1.19)$$

$$\Delta_{\ddagger}S^0 = \frac{\Delta_{\ddagger}H^0 - \Delta_{\ddagger}G^0}{T} \quad (1.20)$$

Equation 1.17 may then be written:

$$k = \kappa \frac{k_{\text{B}}T}{h} e^{-\Delta_{\ddagger}G^0/RT} = \kappa \frac{k_{\text{B}}T}{h} e^{\Delta_{\ddagger}S^0/R} e^{-\Delta_{\ddagger}H^0/RT} \quad (1.21)$$

or, in logarithmic form,

$$\ln k = \ln\left(\frac{\kappa k_{\text{B}}}{h}\right) + \frac{\Delta_{\ddagger}S^0}{R} + \ln T - \frac{\Delta_{\ddagger}H^0}{RT} \quad (1.22)$$

and differentiating Equation 1.22 with respect to $(1/T)$,

$$\frac{d(\ln k)}{d(1/T)} = -T - \frac{\Delta_{\ddagger}H^0}{T} = -\frac{\Delta_{\ddagger}H^0 + RT}{R} = -\frac{E_{\text{a}}}{R} \quad (1.23)$$

where E_{a} is the ordinary (Arrhenius) experimental activation energy, which is thus equal to $\Delta_{\ddagger}H^0 + RT$.

The two theoretical approaches, one in terms of molecular collisions and the other in terms of an activated complex, are not opposed, but complementary. A key to the connection between them is the entropy of activation. When both the rate constant and the temperature coefficient of the rate constant are known, $\Delta_{\ddagger}G^0$ and $\Delta_{\ddagger}H^0$ ($=E_{\text{a}} - RT$) can be used with Equation 1.20 to obtain $\Delta_{\ddagger}S^0$. In an ordinary bimolecular reaction with no special steric requirements, the formation of the activated complex means the formation of one rather “loose” molecule from two. A negative entropy change is to be expected, perhaps comparable to that for the combination of two iodine atoms (Atkins and de Paula, 2010, p. 922),



A value much more negative than this implies the loss of much freedom of motion on formation of the complex, and corresponds to a small value of P , the steric factor in the collision theory. On the other hand, less negative or even positive values of $\Delta_{\ddagger}S^0$ occasionally occur, though rarely if ever for bimolecular reactions in the gas phase. They imply that the complex is very loosely bound, or, in solution, that the complex is less tightly solvated than are the reactant species.

Most reactions in the gas phase at low pressures can be treated as if no foreign molecules (i.e., other than reactants, intermediates, or products of the reaction) are present. Thus the presence of an inert gas such as argon is not important. An exception to this rule is any reaction in which two atoms combine to form a stable diatomic molecule. This cannot happen unless some means exists of getting rid of the energy of formation of the bond. A *third body*, which may be any molecule or the container

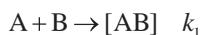
wall, must be present to absorb some of this energy. Its function has been likened (more “psychochemistry?”) to that of a chaperon (Laidler, 1987, p. 183, after G. Porter), present not to prevent union but to ensure that the union is stable and is not formed in an excited state. An extensive literature exists on the efficacy of different molecules as third bodies (Mitchell, 1992; Troe, 1978), and on the influence of the container wall in this and other ways.

1.9 REACTIONS IN SOLUTION

When it comes to reactions in solution, the results of kinetic experiments are difficult to understand, except qualitatively, through the collision theory. The very concept of a collision is hard to define in the liquid phase, in which molecules are not free to travel in straight lines between collisions, but move in constant interaction with neighbors, in a “tipsy reel” (J. H. Hildebrand’s phrase). What happens when two solute molecules come into contact in solution, perhaps to react, or perhaps to diffuse apart unchanged, is sometimes called an “encounter,” rather than a collision. Rabinowitch and Wood (1936) demonstrated this by the use of a model in which a few metal balls rolled about on a level table, making collisions that were detected electrically. When many non-conducting balls were added to the set on the table, so that it became rather crowded, instead of single collisions at long and irregular intervals, collisions happened in groups, while the two metal balls were temporarily trapped in a cage of other balls. Computer modeling in three dimensions, using simulated hard spheres, gave a similar result: collisions in a crowded space between labeled molecules occurred in groups of 10 to nearly 100, depending on the degree of crowding. In the hard-sphere representation, collisions could still be recognized. In a more realistic computer model in which molecular attractions and repulsions are both dependent on distance (which enormously increases the amount of calculation required), an encounter would become a continuous interaction of a complicated kind. During the encounter, something resembling a definite complex, called an *encounter complex*, is present (Langford and Tong, 1977). Eigen and Tamm (1962), in work on ultrasonic effects on solutions of sulfates of divalent metals, interpreted their data as showing that such an encounter complex was formed between the oppositely charged ions, but an encounter complex may exist in the absence of such electrostatic assistance.

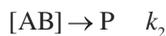
1.10 DIFFUSION-CONTROLLED REACTIONS

Consider a bimolecular reaction in solution as occurring in two steps. In the first step, an encounter complex is formed:



The complex may then either revert to separated reactants or react to form products:





Applying the steady-state assumption to the concentration of the encounter complex:

$$\frac{d[\text{AB}]}{dt} = k_1[\text{A}][\text{B}] - k_{-1}[\text{AB}] - k_2[\text{AB}] = 0 \quad (1.24)$$

it may be shown that the rate of formation of products is given by Equation 1.29:

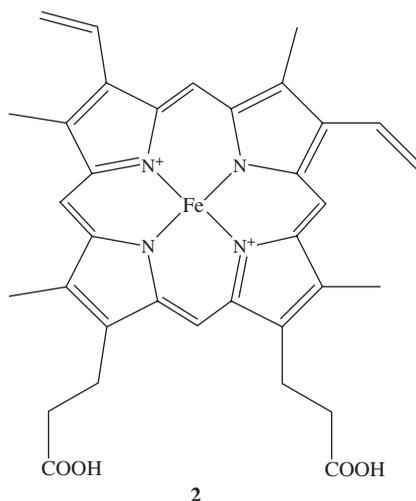
$$\frac{d[\text{P}]}{dt} = k_2[\text{AB}] = k[\text{A}][\text{B}]; \quad k = \frac{k_2 k_1}{k_2 + k_{-1}} \quad (1.25)$$

If the encounter complex reacts to form products much faster than it reverts to reactants, that is, if $k_2 \gg k_{-1}$, then $k = k_1 k_2 / k_2 = k_1$, that is, the rate is controlled by the rate of formation of the encounter complex. Such a reaction is described as *diffusion controlled* or *encounter controlled*.

The magnitude of k_1 is approximately given by Equation 1.26 (Atkins and de Paula, 2010, p. 839–842; Cox, 1994, p. 59):

$$k_1 \approx \frac{8000RT}{3\eta} \quad (1.26)$$

where R is the gas constant and η is the viscosity. A factor of 1000 lets the result be in the conventional units, $\text{l mol}^{-1} \text{s}^{-1}$. A reaction, the rate of which is dependent on bond making or breaking when run in an ordinary solvent, may be diffusion controlled when run in such a highly viscous solvent as glycerol (1,2,3-propanetriol). This has been demonstrated with reactions as diverse as solvent exchange in complexes of Cr^{2+} , Cu^{2+} , and Ni^{2+} (Caldin and Grant, 1973) and reactions of ferroprotoporphyrin IX (**2**) with CO and with O_2 (Caldin and Hasinoff, 1975).



Electron-transfer reactions are a class in which diffusion control may be observed. If an electron-donor species D (reductant) is to react with an electron acceptor A (oxidant), they first form an encounter complex, within which the transfer occurs by tunneling at a rate chiefly determined by the height of the barrier between the donor's HOMO and the acceptor's LUMO, and its width, which is determined by the distance of closest approach. If the transfer rate constant is large, the rate-limiting step will be the formation of the encounter complex by diffusion. This picture is an oversimplification. The theory developed by R. A. Marcus, and independently by others, is described in most physical chemistry texts (e.g., Atkins and de Paula, 2010, pp. 856–861).

1.11 REACTION IN SOLUTION AND THE TRANSITION-STATE THEORY

The most satisfactory way to consider reactions in solution is through the thermodynamic interpretation of the Transition-State Theory, by examining the effects of various properties of the solvent on the activity of each reactant species and on the activated complex, treating the latter almost as “just another molecule.” The solvent can influence the solute molecules by acting on them with “physical” forces (van der Waals forces and electrostatic forces due to the polarity and polarizability of solvent and solute molecules), but also in more obviously “chemical” ways, through the formation of hydrogen bonds or molecular or ionic complexes of various kinds. Changing from an “inert” solvent, one that solvates solutes weakly, to one that exerts stronger forces, may either retard or accelerate a reaction through the change in enthalpy of activation. This depends on whether the latter solvent interacts more strongly with the reactants or with the activated complex. Dewar (1992) discusses an example of a reaction in which the necessity of desolvation of an attacking ion has a profound effect. (Incidentally (p. 160), he describes the hard–soft acid–base distinction as “mythical,” at least as an explanation of the difference between nucleophilic substitutions at carbonyl and saturated carbon atoms.) To take the simplest case, if the reactant in a unimolecular reaction (perhaps an isomerization or an S_N1 substitution) is more strongly solvated, the reaction will be retarded, through the increase in activation enthalpy; if it is the activated complex that is more solvated, the reverse effect will be found (see Fig. 1.8).

The free energy of solvation is a composite of the enthalpy and entropy. Entropy of solvation can also have large effects. Strong solvation usually implies loss of entropy, owing to relative immobilization of solvent molecules. Strong solvation of the reactant, therefore, makes the entropy of activation more positive, thus (from Eq. 1.20) making the Gibbs energy of activation less positive, and the reaction therefore faster. The effect of strong solvation on the entropy of the activated complex, on the other hand, retards the reaction. Thus, the enthalpy and entropy of solvation of either the reactant or the activated complex have opposite effects. Prediction of the overall effect requires that these be disentangled. The required information concerning reactants is in principle available. That for activated complexes is not,

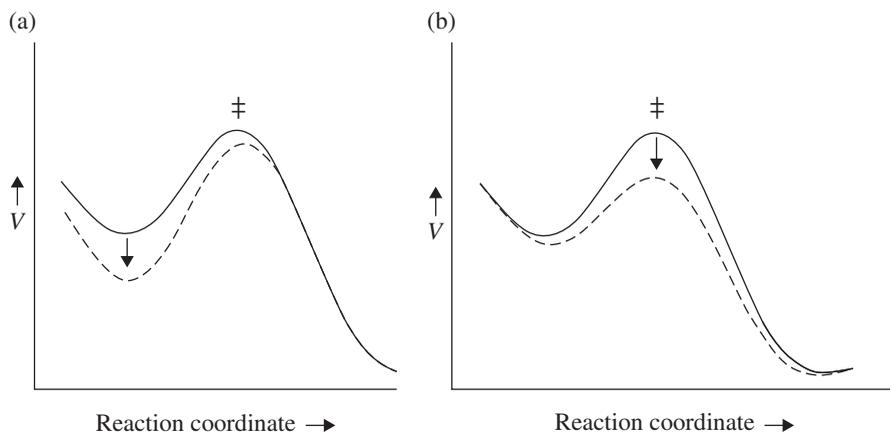


FIGURE 1.8 Effect of solvation on activation energy. Potential energy V versus the reaction coordinate. Solid curves represent the energy profile in the absence of solvation. (a) Solvation of the reactant (increased activation energy). (b) Solvation of the activated complex (reduced activation energy).

though estimates may be made if data on molecules that resemble a postulated activated complex are known. In favorable cases, the dissection of the kinetics into these parts can be done (Blandamer, 1977; Buncl and Symons, 1981; Buncl and Wilson, 1977, 1979; Tobe and Burgess, 1999, p. 346, 363).



The Menshutkin Reaction (1.27) provides an example of a case where in polar solvents the solvation of the activated complex has a major effect on the rate. Hartmann and Schmidt (1969) showed that in a series of 12 solvents of increasing polarity from 1,1,1-trichloroethane ($\epsilon_r = 7.52$, $E_r^N = 0.170$) to nitrobenzene ($\epsilon_r = 34.78$, $E_r^N = 0.324$) the rate increases by a factor of 52, illustrating the acceleration due to solvation of the activated complex. It also illustrates the importance of the entropy of activation. Here the change in the Gibbs energy of activation is made up of reinforcing contributions from enthalpy and entropy changes; the contribution of the latter at 50°C is 2.5 times that of the former.

The overall effect on the reaction rate thus depends on the free energies of the initial and transition states. The various possibilities, in terms of the free energy, are summarized qualitatively in Table 1.3. Reinforcement occurs if the transfer free energies of reactants and transition state have opposite signs. If they have the same sign, partial or complete balancing is expected.

For reactions in solution an additional thermodynamic property that can be helpful is available. The effect of pressure on the equilibrium constant of a reaction yields the volume change of reaction, ΔV , given by Equation 1.28.

TABLE 1.3 Transfer Free Energies of Reactants ($\delta_{\text{r}}G^{\text{R}}$) and Transition States ($\delta_{\text{r}}G^{\text{T}}$) and Solvent Effects on Reaction Rates. Classification of Reaction Types

Case	$\delta_{\text{r}}G^{\text{R}}$	$\delta_{\text{r}}G^{\text{T}}$	Effect on rate ^a	Reaction type
1	–	–	+, 0 or –	Balanced
2	+	–	+	Positively reinforced
3	0	–	+	Positive transition-state control
4	–	0	–	Negative initial-state control
5	+	0	+	Positive initial-state control
6	0	0	0	Solvent independent
7	–	+	+	Positively reinforced
8	+	+	+, 0 or –	Balanced
9	0	+	–	Negative transition-state control

^aThe plus sign refers to rate acceleration, the minus sign to rate retardation, and zero to no effect (Buncel and Wilson, 1979).

$$\Delta V = -RT \left(\frac{\partial(\ln K)}{\partial P} \right)_{\text{T}} \quad (1.28)$$

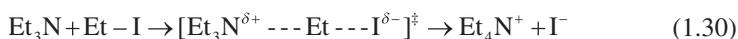
The analogous effect of pressure on the rate constant gives the volume of activation, $\Delta_{\ddagger}V$, through Equation 1.29. Measurement of reaction rates at high pressures, as Tobe and Burgess (1999, p. 10) point out, requires specialized apparatus; nevertheless, a great many volumes of activation are now available: see Isaacs (1984), van Eldik and Hubbard (1997), van Eldik and Meyerstein (2000), van Eldik *et al.* (1989), Blandamer and Burgess (1982), Laidler (1987), Tobe and Burgess (1999), and the large compilations of volumes of reaction and activation in the reviews by Drljaca *et al.* (1998 and references therein).

$$\Delta_{\ddagger}V = -RT \left(\frac{\partial(\ln k)}{\partial P} \right)_{\text{T}} \quad (1.29)$$

It is not likely that $\Delta_{\ddagger}V$ is large, though values outside $\pm 10 \text{ ml mol}^{-1}$ have been obtained, notably for reactions consuming or generating ions in polar solvents. van Eldik and Meyerstein (2000) show that in favorable cases there is a linear correlation between $\Delta_{\ddagger}V$ and ΔV . They present the example of substitutions on $\text{Pd}(\text{H}_2\text{O})_4^{2+}$ in a variety of solvents, where $\Delta_{\ddagger}V \approx \Delta V - 2 \text{ cm}^3 \text{ mol}^{-1}$. Where the activated complex resembles the products, this correlation is not unexpected, but it is by no means universal. Tobe and Burgess (1999) present *volume profiles*, which are schematic graphs of the volume changes along the reaction pathway, showing cases in which a degree of correlation exists (p. 536) and others in which it clearly does not (pp. 11, 301).

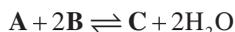
Volumes and entropies of activation for many classes of reactions show parallel trends, and can be interpreted in similar terms. In some cases the volume of activation is more reliable than the entropy of activation, because the latter is obtained by what may be a long extrapolation of the plot of $\ln k$ against $1/T$ to obtain the intercept.

The volume of activation for a reaction in an inert solvent can be a help in the assignment of a mechanism, because dissociative activation may be assumed to result in a positive volume of activation (in the region of 10–15 cm³ mol⁻¹ for each bond assumed to be stretching in the activation process), and associative activation the reverse. In a solvent that interacts strongly with solutes, however, these interactions must also be taken into account. Reactions in which ions are generated, such as Menshutkin reactions (e.g., 1.30), are characterized in solution by large, negative, and solvent-dependent entropies and volumes of activation: $\Delta_{\ddagger}V = -12$ to -58 cm³ mol⁻¹ (Tobe and Burgess, 1999), because solvation of the nascent ions leads to reduced solvent freedom, reducing the entropy, and at the same time electrostriction of the solvent, reducing the volume.



PROBLEMS

- 1.1** Data are tabulated for the equilibrium in aqueous solution between *N,N'*-bis-(hydroxymethyl)-uracil, **A**, and methanol, **B**, to form the diether, **C**, and water:



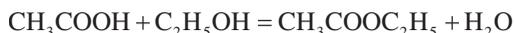
(Reagents were mixed in stoichiometric amounts.)

C_0	0.10	0.25	0.5	1
f at 17°C	0.0125	0.066	0.18	0.35
f at 30°C	0.018	0.091	0.23	0.40

C_0 = initial concentration of **A** in mol l⁻¹.

f = fraction converted at equilibrium.

- Calculate the mean equilibrium constant at each temperature (using the convention that the activity of water, approximately constant, is unity).
 - Calculate ΔG^0 at each temperature, and assuming they are constant, ΔH^0 and ΔS^0 .
 - With the same assumption, calculate the equilibrium constant and the fraction converted at 100°C, $C_0 = 1.0$. (Hint: Try successive approximation, use a calculator with a “Solve” program, or solve graphically.)
- 1.2** Use the values of enthalpy of formation and entropy given here to calculate the equilibrium constant at 25°C for the esterification reaction in the vapor phase:



Does it make any difference to the numerical value whether the constant is expressed in mole fraction, concentration, or pressure units?

- (a) Use the vapor pressures of the pure substances given to calculate the equilibrium constant (in mole fraction terms) for this reaction in a solvent in which all four substances form ideal solutions (a practical impossibility).

Substance (all as gas)	$\Delta_f H^0_{298}$ kJ ⁻¹ mol ⁻¹	S^0_{298} J ⁻¹ K ⁻¹ mol ⁻¹	p^0_{298} mm ⁻¹ Hg
Acetic acid	-434.3	282.7	15.4
Ethanol	-235.37	282	57.2
Ethyl acetate	-437.9	379.6	90.5
Water	-241.8	188.83	23.8

- 1.3 Pure *p*-xylene and water were equilibrated at 25°C. The absorbance A_0 of the aqueous layer measured in a 1 cm cell at $\lambda_{\max}=274$ nm (due to *p*-xylene) was 0.884. A solution of *p*-xylene, mole fraction $x_1=0.686$, and *n*-dodecane, similarly treated, gave absorbance $A=0.749$. Assuming that both Beer's and Henry's laws hold for *p*-xylene in water and that *n*-dodecane is insoluble in water, what was the activity coefficient of *p*-xylene in the solution with *n*-dodecane? (Neglect the small solubility of water in *p*-xylene.)
- 1.4 The rate constants of the reaction between OH radical and H₂ in the gas phase at 25, 45, and 100°C were found to be 3.47×10^3 , 1.01×10^4 , and 1.05×10^5 l mol⁻¹ s⁻¹, respectively. (a) What are the Arrhenius activation energy E_a and the pre-exponential factor PZ at 25°C? (b) Calculate the activation equilibrium constant K^\ddagger and the activation parameters $\Delta^\ddagger H^0$, $\Delta^\ddagger S^0$, and $\Delta^\ddagger G^0$. (c) If the collision diameters of OH and H₂ are 310 and 250 pm, calculate the collision number Z and obtain an estimate of P at 25°C. (Caution: watch the units!)