Rules of Thumb for Assessing Equilibrium Partitioning of Organic Compounds: Successes and Pitfalls

Kai-Uwe Goss* and René P. Schwarzenbach
Swiss Federal Institute for Environmental Science and Technology (EAWAG) and Swiss Federal Institute of Technology (ETH), CH-8600 Dübendorf, Switzerland; **goss@eawag.ch*

Chemists often face situations where they need to understand the equilibrium partitioning of an organic compound from one phase into a different phase. Such situations include chromatographic processes, extraction processes, preparation of solutions, clean up of water or air by a sorbent, and transport of pollutants in the environment. However, the estimation of partition constants is by no means trivial. The partition constants[1] of a given compound between different media can vary by many orders of magnitude. For example, octane has a hexane/air partition constant of 9300 whereas its ethanediol/air partition constant is only 13 *(1)*. Consequently, many different concepts have been developed to estimate partition constants, $K_{i\,1,2}$, which describe the partition equilibrium of a compound, *i*, between two phases, 1 and 2

$$K_{i\;1,2} \equiv \frac{\text{concentration of } i \text{ in phase 1 } (c_{i\,1})}{\text{concentration of } i \text{ in phase 2 } (c_{i\,2})}$$

The theory behind these concepts ranges from quantum chemical calculations to purely empirical approaches *(2–6)*. However, in practical situations in the lab one often only needs a rough estimate of the relative order of partition constants. Hence, simple rules of thumb have evolved to judge the plausibility of the observed retention order in chromatography, or to select an appropriate solvent or sorbent for an extraction process or sample cleanup. These intuitive estimations are often based on a characterization of compounds and phases in terms of volatility, polarity, or hydrophobicity. The following are some examples of frequently used rules of thumb for the partitioning equilibrium of organic compounds:

> *Volatility:* Partitioning between air and a bulk phase correlates with the vapor pressure of the compounds.
>
> *Polarity:* Like dissolves like. Polar molecules prefer polar phases and nonpolar molecules prefer nonpolar phases.
>
> *Hydrophobicity:* There are repulsive forces between hydrophobic molecules and polar molecules (e.g., water). Hydrophobic interactions are attractive interactions that only occur between nonpolar molecules but not between nonpolar and polar molecules.

Table 1. Air/Olive Oil Partition Constants for Compounds with Similar Saturation Vapor Pressure

Compound	p_L* / Pa	$K_{\text{olive oil,air}}$[a]
N,N-Dimethylaniline	107	110,000
Hexanol	110	14,000
Triethylamine	7,700	1,200
Ethanol	7,870	140

[a] Olive oil/air partition data is from ref *7*, extrapolated to 25 °C.

Misleading Concepts

In several surveys we have found that approximately 90% of our graduate students and also more advanced scientists intuitively agree with the statements above. But are the statements really correct? Let us take a closer look at these rules of thumb.[2]

Volatility

There is no general relationship between the saturated liquid vapor pressure, p_L* (i.e., the partitioning between air and the pure liquid phase of a compound), and its partitioning between air and other liquid phases. A correlation can only be expected if both partition equilibria are governed by the same type of interactions. The examples in Table 1 show pairs of compounds with similar p_L* that differ by almost an order of magnitude in their partitioning between olive oil and air as a result of differences in their capability to form hydrogen bonds with like and unlike molecules.

Polarity

The correct rule, like dissolves like, is often misunderstood as like only dissolves like. The latter would imply that polar molecules prefer polar phases and nonpolar molecules prefer nonpolar phases. This is not correct as the following examples show (data from ref *1*):

- Nonpolar octane has a similar preference to partition into the polar diethyl ether phase as into the nonpolar decane phase: $K_{\text{diethyl ether, air}}$ = 8,300; $K_{\text{decane, air}}$ = 6,700
- Polar 1,4-dioxane has the same preference for the nonpolar CCl_4 phase as for the more polar ethyl acetate phase and it strongly prefers the $CHCl_3$ phase: $K_{\text{CCl4, air}}$ = 4,300; $K_{\text{ethyl acetate, air}}$ = 4,700; $K_{\text{CHCl3, air}}$ = 27,400

The problem results from the classification of a compound as polar relative to the presence of a permanent dipole moment. However, dipole–dipole interaction energies in bulk phases are very small compared to other intermolecular interactions *(8, 9)*, and therefore, have no substantial influence on the bulk phase partitioning of uncharged organic compounds. Instead, it is the hydrogen-bond polarity that governs much of the partitioning behavior that we are interested in. It does not suffice to distinguish between hydrogen-bonding molecules (i.e., polar) and non-hydrogen-bonding molecules (i.e., apolar); hydrogen bonds are specific and do not simply occur between all polar molecules but only between hydrogen donors (HD) and hydrogen acceptors (HA). For example, the polar 1,4-dioxane, a HA, cannot form hydrogen bonds with ethyl acetate, also a HA, and therefore does not like the polar ethyl acetate any better than the apolar CCl_4. However, 1,4-dioxane has a 6 times higher partition constant into $CHCl_3$, which is a HD. Hence, a simple classification of compounds as polar or apolar is not sufficient for understanding partitioning.

Hydrophobicity

Hydrophobicity, which means water fearing, is often misunderstood as being the result of repulsive forces between hydrophobic molecules and water molecules, or specific attractive interactions between nonpolar molecules. These statements would indeed explain the low water solubility of hydrophobic compounds but they are falsified by the following experimental findings:

- Adsorption of the hydrophobic compound toluene from air to a water surface is twice as high as to a nonpolar Teflon surface.
- Small water droplets adhere to a hydrophobic wax or Teflon surface even against gravity. They will not fall off if the surface is turned upside down as they should if the intermolecular interactions between water and the surface were repulsive.
- Hexane spreads on a water surface to form a monomolecular film. If there were repulsive forces between hexane and water or if the interactions of hexane with itself were more attractive than those with water, then hexane would minimize its contact area with the water surface and form a droplet with a high contact angle.

The hydrophobic effect comes from the high amount of free energy that is required to form a cavity in the bulk water phase. It renders the partitioning of nonpolar molecules into an aqueous phase unfavorable because these molecules cannot compensate the neighboring water molecules for their disturbed H-bonding structure. Hence, hydrophobicity is the result of water–water interactions being more attractive than the also attractive interactions between water and hydrophobic molecules. Obviously, the incorrect idea of repulsive interactions between water molecules and hydrophobic molecules would still provide an explanation of the low water solubility of hydrophobic compounds. However, this explanation completely fails if applied to surfaces or interfaces. Phenomena that occur at a water surface do not exhibit a hydrophobic effect because they do not require any cavity formation. The examples given above, which may seem counter intuitive, are, in fact, quite easy to understand:

- Water exhibits stronger van der Waals interactions than Teflon *(10)* with any interaction partner so that even nonpolar molecules adsorb stronger to a water surface than to a Teflon surface.
- Water adheres by van der Waals interactions to any surface even nonpolar ones like wax.
- Hexane has stronger van der Waals interactions with water than with itself so that the formation of a hexane-monolayer on a water surface is favored over the formation of a droplet.

In the following section we will present a simple concept of partitioning that helps to avoid misconceptions like those presented above. Based on this concept we will then develop a set of simple rules that allows the rough estimation of relative partition constants of a given compound into various bulk phases.

A Simple Model for Evaluation of Bulk Phase Partitioning

The transfer of a molecule, *i*, from one phase to another requires that interactions between molecules are given up while new interactions become possible. The more total Gibbs free energy, $\Delta G_{i\,1,2}$, that is released from a system by the transfer of one molecule, *i*, from phase 1 to phase 2, the higher the equilibrium concentration of *i* in phase 2 will be *(11)*. In more quantitative terms, the logarithm of the partition constant is proportional to the change in Gibbs free energy of the partition process^3

$$\ln K_{i\,1,2} \propto -\Delta G_{i\,1,2} \qquad (1)$$

The Cavity Approach

For reasons of simplicity it is helpful to first consider the transfer of *i* from air (a bulk phase without any interactions) into a bulk phase 2. This transfer always involves two steps (see Figure 1). The first step is the creation of a cavity in phase 2 for the molecule, which requires Gibbs free energy ($G_{i\text{ cavity }2} > 0$). The free energy required depends on the interaction free energy between the like molecules of phase 2, that is, the cohesive free energy in phase 2 ($G_{\text{cohesion }2}$ is a negative term), and on the size of the required cavity (Size_i) dependent on the size of the molecule *i*

$$G_{i\text{ cavity }2} = -\text{Size}_i\, G_{\text{cohesion }2} \qquad (2)$$

A higher cavity energy leads to a smaller equilibrium concentration in the respective phase. In the second step *i* is placed inside the cavity. Free energy is released in this step ($G_{i\,2} < 0$) by the interactions between *i* and its new neighbors, the molecules of phase 2 . This interaction energy favors the partitioning into the respective phase.

Obviously, the energy balance for the transfer of a molecule between two liquid phases 1 and 2 requires that these two steps are considered for both phases, however, with opposite signs. Hence, equilibrium partitioning between two phases (with low mutual solubility) can be understood as a combination of the single bulk phase/air partition constants

$$\ln K_{i\,1,2} = \ln K_{i\,1,\text{air}} - \ln K_{i\,2,\text{air}} \qquad (3)$$

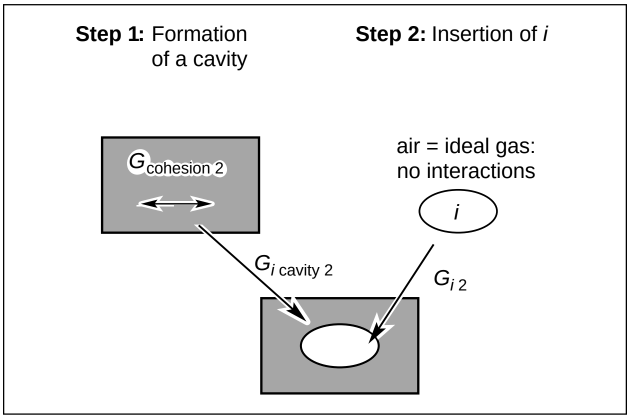


Figure 1. Schematic model of the interactions involved when a molecule partitions from air into a bulk phase 2.

or using partition constants in their nonlogarithmic form

$$\frac{C_{i\,1}}{C_{i\,2}} = \left(\frac{C_{i\,1}}{C_{i\,\mathrm{air}}}\right) \Bigg/ \left(\frac{C_{i\,2}}{C_{i\,air}}\right) \qquad (4)$$

The Interaction Energies

The energies involved in the partition process come from the free energy of interaction of neighboring molecules. For nonionic organic compounds there are only two types of interactions that are important: van der Waals interactions and H-bonds (more generally: weak Lewis acid–base or electron donor–acceptor interactions). Consequently, we can rewrite eq 1 for the partitioning between air and a bulk phase 2 as

$$\begin{aligned} \ln K_{i\,2,\mathrm{air}} &\propto -G_{i\,\mathrm{cavity}\,2} - G_{i\,2} \\ &= \mathrm{Size}_i \left(G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2} + G^{\mathrm{H}}_{\mathrm{cohesion}\,2}\right) - G^{\mathrm{vdW}}_{i\,2} - G^{\mathrm{H}}_{i\,2} \end{aligned} \qquad (5)$$

Fortunately, only little needs to be known about these intermolecular interactions in order to use eq 5 for a good qualitative understanding of partitioning.

van der Waals interactions are attractive and occur between any kind of molecules, even between water and alkanes; they comprise dispersive (induced dipole–induced dipole) interactions, Debye (dipole–induced dipole) interactions, and Keesom (dipole–dipole) interactions. The dispersive interactions dominate the overall van der Waals interactions in condensed phases and at interfaces *(8, 9)*. For a simple but convenient visualization of van der Waals forces, all molecules may be thought of as being covered with a “glue”. The strength of the attraction between molecules *i* and their interaction partners depends on the stickiness of the glue and the contact area (Figure 2).

The interesting aspect about van der Waals interactions is that the stickiness of the glue is similar for organic compounds and water. Hence, a given molecule exhibits similar van der Waals interactions in any organic phase and in the water phase (energy varies by less than a factor of 1.7). As a consequence, van der Waals interactions cannot be the cause of huge differences in the partitioning constants of a given compound *i* in various phases.

H-bond interactions also are attractive but can only occur between a H-donor (e-acceptor) and a H-acceptor (e-donor). Hence, these interactions are specific. While van der Waals interactions were visualized as resulting from sticky surfaces, H-bond interactions have an analogy to Velcro: only the complementary parts of the Velcro will stick together (Figure 3).

To understand the partition behavior of an organic compound it is crucial to recognize that a simple distinction between nonpolar and polar molecules is not sufficient. Instead it is important to distinguish between molecules that are apolar (no H-donor or H-acceptor), monopolar (H-donor or H-acceptor), or bipolar (H-donor and H-acceptor).[4] A classification of a compound can usually be accomplished when its structural formula is known (see Table 2). For example, compounds that contain an −OH or −NH group are H-donors; compounds containing oxygen or nitrogen are strong H-acceptors; and compounds with π electrons are weak H-acceptors.

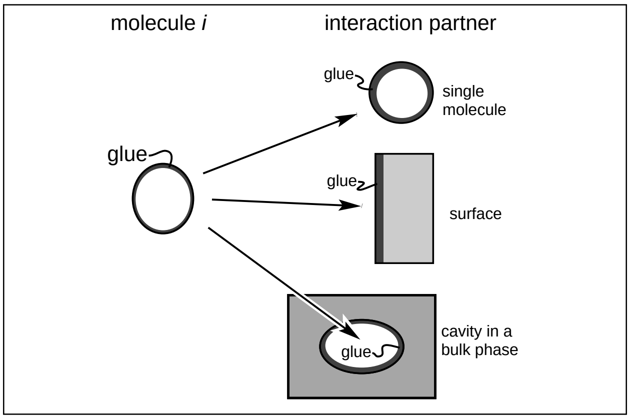


Figure 2. A simple analogy to illustrate van der Waals interactions: attraction occurs between all kinds of molecules and the strength, $G^{\mathrm{vdW}}_{i\,2}$, depends mainly on the contact area.

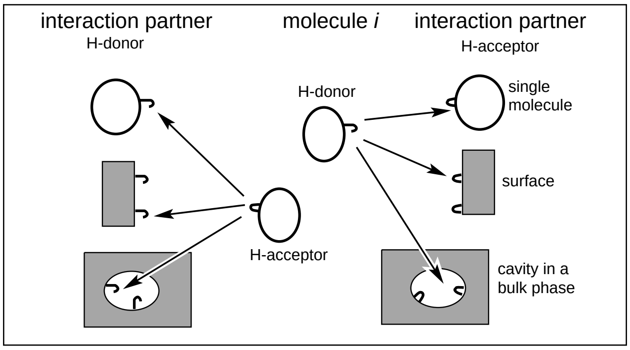


Figure 3. A simple analogy to illustrate H-bond interactions: attraction only occurs between molecules with complementary functionalities (symbolized by hooks, H-donors, and eyes, H-acceptors). The strength, $G^{\mathrm{H}}_{i\,2}$, depends on the number and strength of hooks and eyes that are connected.

Table 2. Different Abilities of Organic Compounds To Interact by van der Waals and H-Bond Forces

Compound	Interactions	Examples
Apolar	only van der Waals	alkanes, chlorobenzenes,[b] PCBs[b]
Monopolar	van der Waals + H-acceptor (e-donor)	alkenes, alkynes, alkylaromatic compounds, ethers, ketones, esters, aldehydes
Monopolar	van der Waals + H-acceptor (e-donor)	$CHCl_3$, CH_2Cl_2
Bipolar[a]	van-der-Waals + H-donor + H-acceptor	R–NH_2, R_2–NH, R–COOH, R–OH

[a] Intramolecular H-bonds like in 2-nitrophenol strongly reduce the ability of the compound to form H-bonds with neighboring molecules.

[b] Compounds whose π-electron density is diminished by electron withdrawing substituents are apolar.

General Rules for the Partitioning of an Uncharged Organic Compound into Different Bulk Phases

The few aspects discussed about intermolecular interactions and the cavity concept enable the derivation of some very useful qualitative rules for the relative partitioning of a molecule from a given reference phase into various bulk phases.[5] An overview of the cases and the respective rules that can be derived are shown in Figure 4. In the following sections, these cases and rules are discussed in detail. Partition constants supporting the rules are given in Table 3.

Partitioning Processes Involving Only van der Waals Interactions

When only van der Waals interaction are involved, the cohesive free energy in the bulk phases and the interactions between *i* and phase 2 will only depend on the van der Waals forces and eq 5 simplifies to

$$\ln K_{i\,2,\,\mathrm{air}} \propto \mathrm{Size}_i\, G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2} - G^{\mathrm{vdW}}_{i\,2} \qquad (6)$$

It was pointed out above that the van der Waals stickiness of a bulk phase does not vary very much for different liquid organic phases or even water. Hence, the variance of $G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2}$ and $G^{\mathrm{vdW}}_{i\,2}$ for a given solute *i* and different phases 2 is rather small and the following rule of thumb is obtained:

> I. Partition constants of a given organic compound *i* from any reference phase into phases in which only van der Waals interactions occur (in the cavity formation as well as in the interactions between the molecule *i* and the phases) do not differ very much from each other (usually by less than a factor of three).

There are three different situations where solely van der Waals interactions occur in a partition process and where the partition constants therefore are similar: any kind of compound partitioning into different apolar phases (partitioning into hexane and CCl_4 is similar for each compound, as seen in Table 3); any apolar compound partitioning into monopolar and apolar phases (octane shows similar partitioning into *n*-hexane, CCl_4, diethyl ether, and $CHCl_3$, as seen in Table 3); any monopolar compound partitioning into monopolar phases of the same polarity and into apolar phases (partitioning of pyridine into diethyl ether is comparable to its partitioning into *n*-hexane and CCl_4, as seen in Table 3).

Partitioning Processes Involving van der Waals and H-bond Interactions

In those cases where H-bonds are involved in the partitioning process, the partition constants of a given compound in different solvents may vary by much more than a factor of three. One can distinguish three subcases where H-bond interactions occur in partitioning.

In the first scenario H-bonds only affect the cavity energy but not the interactions between the compound *i* and phase 2 (i.e., apolar compounds partitioning into bipolar phases). In this case eq 5 simplifies to

$$\ln K_{i\,2,\mathrm{air}} \propto \mathrm{Size}_i \left(G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2} + G^{\mathrm{H}}_{\mathrm{cohesion}\,2} \right) - G^{\mathrm{vdW}}_{i\,2} \qquad (7)$$

The cavity energy required for the partitioning into bipolar solvents may be much larger than that in apolar and monopolar solvents because H-bonds must be overcome in addition to the always present van der Waals interactions to form a cavity. Hence, rule IIa follows:

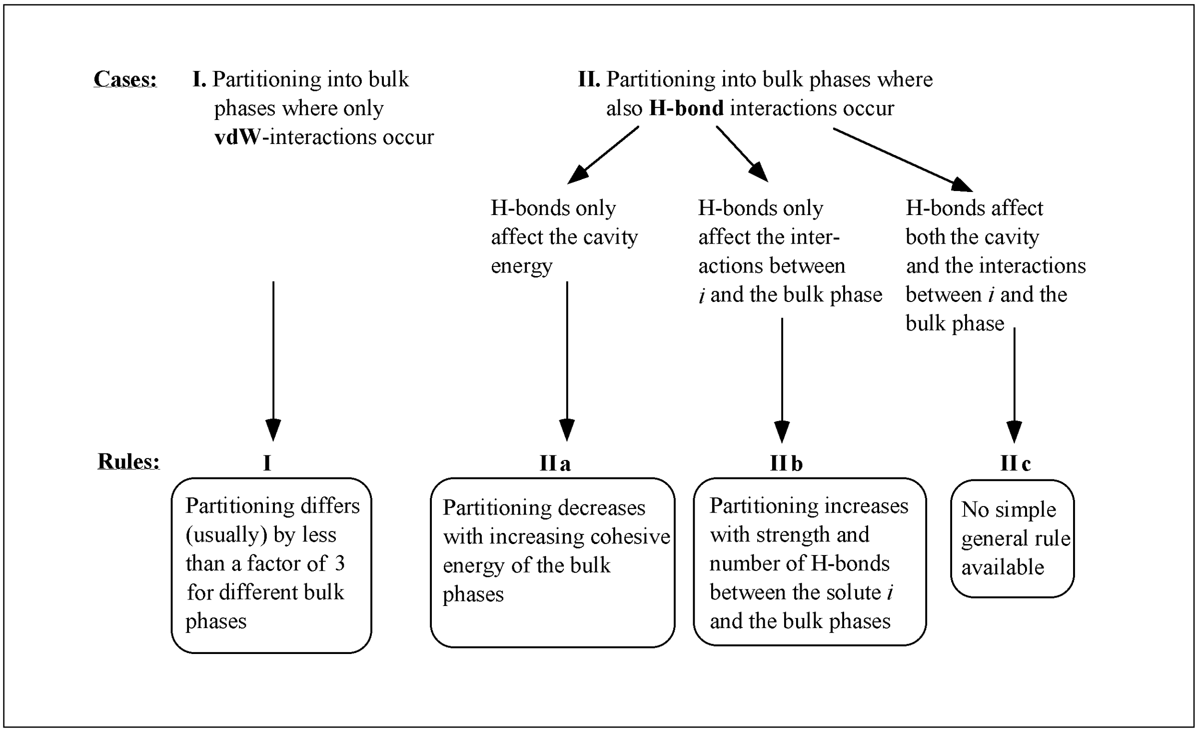
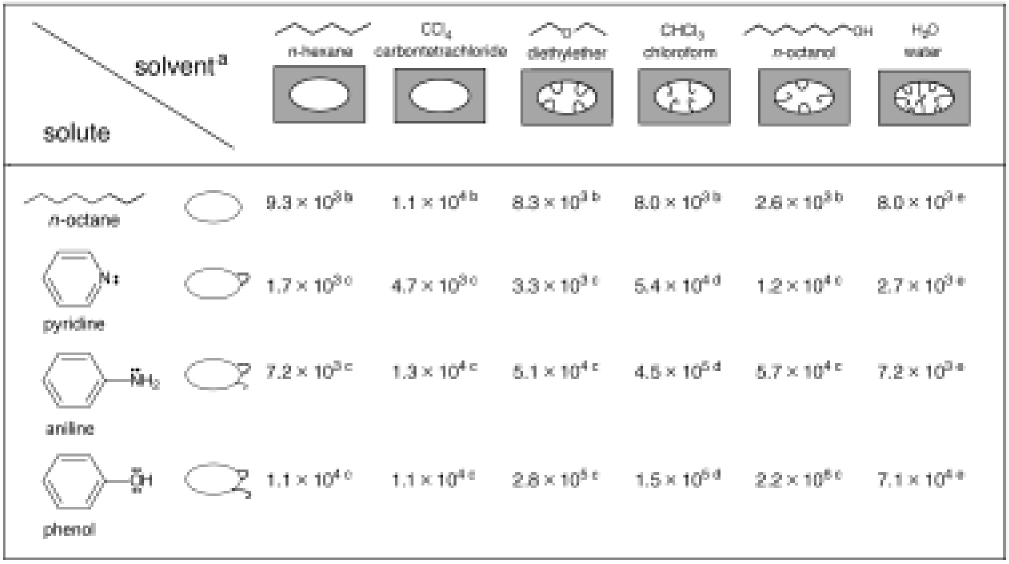


Figure 4. Rules for the relative partitioning of a given organic molecule, *i*, from a reference phase into different bulk phases.

Table 3. Partition Coefficients, $K_{i\,\mathrm{solvent,\,air}}$, of Uncharged Organic Solute Molecules between the Solvent and Air Phases at 25 °C



solute \ solvent[a]		n-hexane	CCl_4 carbontetrachloride	diethylether	$CHCl_3$ chloroform	n-octanol	H_2O water
n-octane		9.3×10^{3} [b]	1.1×10^{4} [b]	8.3×10^{3} [b]	8.0×10^{3} [b]	2.6×10^{3} [b]	8.0×10^{3} [e]
pyridine		1.7×10^{3} [c]	4.7×10^{3} [c]	3.3×10^{3} [c]	5.4×10^{4} [d]	1.2×10^{4} [c]	2.7×10^{3} [e]
aniline		7.2×10^{3} [c]	1.3×10^{4} [c]	5.1×10^{4} [c]	4.5×10^{5} [d]	5.7×10^{4} [c]	7.2×10^{3} [e]
phenol		1.1×10^{4} [c]	1.1×10^{4} [c]	2.8×10^{5} [c]	1.5×10^{5} [d]	2.2×10^{5} [c]	7.1×10^{4} [e]

[a] Diagrams refer to H-donor or H-acceptor characteristics as defined in Figure 3. [b] Data from ref *1.* [c] Partition coefficient calculated from solvent/air *(14)* and water/air partition constants. [d] Partition coefficient from ref *15.* [e] Partition coefficient from ref *2.*

> IIa. The partitioning of a given apolar compound into different bipolar solvents decreases with increasing cohesive energies of the solvents. The cohesive energy increases with strength and the number density of the H-bonds in the solvent (number density is equal to the number of H-bonds that must be interrupted for a cavity of a given size).

The partitioning of octane between air and various solvents can serve as an example (see data in Table 3 and an extended data set Table 4).

In the second case, in which H-bonds only affect the interactions between the compound *i* and the molecules of the bulk phase, eq 5 simplifies to

$$\ln K_{i\,2,\mathrm{air}} \propto \mathrm{Size}_i\, G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2} - G^{\mathrm{vdW}}_{i\,2} - G^{\mathrm{H}}_{i\,2} \qquad (8)$$

These cases occur for monopolar compounds *i* in monopolar phases of opposite polarity and for bipolar compounds in monopolar phases. In these cases the variance in partitioning depends mainly on the variance in $G^{\mathrm{H}}_{i\,2}$ because the variance in $G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2}$ and $G^{\mathrm{vdW}}_{i\,2}$ is rather small (see rule I). Hence, rule IIb applies:

> IIb. Partitioning of a compound to a monopolar phase will increase with the strength and number of H-bonds that it can form with this phase.

In Table 3 the following examples can be found for this rule: pyridine shows a much higher partitioning into chloroform than into hexane, CCl_4, or diethyl ether; aniline and phenol exhibit a higher partitioning into diethyl ether and chloroform than into hexane or CCl_4.

In the third case, in which H-bonds affect the cavity energy and the interactions between the compound *i* and the phase, eq 5 is unchanged

$$\ln K_{i\,2,\mathrm{air}} \propto \mathrm{Size}_i \left(G^{\mathrm{vdW}}_{\mathrm{cohesion}\,2} + G^{\mathrm{H}}_{\mathrm{cohesion}\,2}\right) - G^{\mathrm{vdW}}_{i\,2} - G^{\mathrm{H}}_{i\,2} \qquad (9)$$

These cases occur for monopolar and bipolar compounds in bipolar phases. No general rule can be derived for these cases because the H-bonds have two opposite effects: they increase the cavity energy, which lowers the partitioning into the respective phase, and they increase interactions between *i* and the phase, which favors the partitioning of *i* into the phase. Examples from Table 3 are the partitioning of pyridine, aniline, and phenol into octanol and water.

Example from Solvent/Solvent Partitioning

All the rules given above were illustrated with examples from air/bulk phase partitioning. However, the rules can also easily be applied to partitioning between two liquid phases. The problem of extracting pyridine from a water sample may serve as an example. For an effective extraction we need to find a solvent with a high solvent/water partition constant for pyridine. The overall free energy of this transfer, $\Delta G_{\mathrm{pyridine\ solvent,water}}$, depends on the free energy of taking pyridine out of the water phase, $\Delta G_{\mathrm{pyridine\ air,water}}$, and inserting it into the solvent, $\Delta G_{\mathrm{pyridine\ solvent,air}}$,

$$\Delta G_{\mathrm{pyridine\ solvent,water}} = \Delta G_{\mathrm{pyridine\ solvent,air}} + \Delta G_{\mathrm{pyridine\ air,water}} \qquad (10)$$

or equivalently,

$$\frac{c_{\mathrm{pyridine\ in\ solvent}}}{c_{\mathrm{pyridine\ in\ water}}} = \left(\frac{c_{\mathrm{pyridine\ in\ solvent}}}{c_{\mathrm{pyridine\ in\ air}}}\right)\left(\frac{c_{\mathrm{pyridine\ in\ air}}}{c_{\mathrm{pyridine\ in\ water}}}\right) \qquad (11)$$

The selection of a suitable solvent only depends on $\Delta G_{\mathrm{pyridine\ solvent,air}}$, because $\Delta G_{\mathrm{pyridine\ air,water}}$ is constant no matter which solvent is used.[6] The discussion above has revealed that a monopolar, H-donor solvent like chloroform would be best suited for the extraction of a H-acceptor molecule like pyridine. The experimental water/solvent partition constants of pyridine confirm this estimation (data from ref *13*): $K_{\mathrm{hexane,water}} = 0.62$, $K_{\mathrm{CCl_4,water}} = 1.7$, $K_{\mathrm{diethyl\ ether,water}} = 1.2$, $K_{\mathrm{CHCl_3,water}} = 23$.

Table 4. Partition Coefficients of *n*-Octane between Air and Solvents with Different Cohesive Energies at 25 °C

Solvent	Formula	$K_{solvent/air}$
Heptane	C_7H_{16}	8600
Diethyl ether	$H_5C_2{-}O{-}C_2H_5$	8300
Butanol	$C_4H_9{-}OH$	2300
1,3-Propanediol	$C_3H_6{-}(OH)_2$	80
1,2-Ethanediol	$C_2H_4{-}(OH)_2$	13
Water	$H{-}O{-}H$	0.008

NOTE: Based on data from ref *1* and *3;* see Note 7.

The Solubility of Nonionic Liquid Organic Compounds

The above rules can also be applied to the dissolution equilibrium, that is, the partition equilibrium between the pure liquid phase of a compound *i* and a solvent phase. One then finds that a high mutual solubility can be expected for apolar and monopolar phases in any combination. In contrast the mutual solubility of bipolar and apolar compounds is rather low.

Conclusion

The examples given above have demonstrated that one should be very careful when using terms like hydrophobicity, hydrophobic interactions, polarity, and volatility in the interpretation or qualitative prediction of partition data. These terms are ill-defined and can easily lead to inappropriate interpretations or wrong predictions.

In order to achieve a correct understanding of the partition behavior of organic compounds it is imperative to: distinguish between van der Waals and H-bond interactions; consider both cavity formation and interactions between compound and phase(s) for any bulk phase partitioning; and distinguish between bulk phase partitioning and the adsorption to surfaces or interfaces.

The rules given above should enable a good qualitative understanding of bulk phase partitioning. Analogous rules can be derived for adsorption to surfaces, but this lies beyond the scope of this paper.

Finally, note that the same concepts used here for a qualitative understanding of partitioning can also be applied to the quantitative description of partition constants if quantitative parameters for the van der Waals and H-donor or H-acceptor interactions are used (e.g., ref *3, 15–17*).

Acknowledgments

We thank Torsten Schmidt, Martin Elsner, and Andreas Kappler for reviewing the manuscript.

Notes

1. The dimensionless partition constants used in this article are on a mass-per-volume basis.

2. The experimental partition constants used in the following for illustration of the principles of partitioning are mainly solvent/air and not solvent/solvent partition data. The reasons are twofold: much more experimental data are available for solvent/air than for solvent/solvent partitioning, and solvent/air partitioning is a simpler system since only interactions in the solvent phase have to be considered (air can be regarded as an ideal gas without interactions). The principles that apply to solvent/solvent partitioning, however, are the same and solvent/solvent partition constants can easily be derived from the respective solvent/air constants for solvents with low mutual solubility

3. The term, $\Delta G_{i\,1,2}$, is negative if free energy is released.

4. The terms apolar, monopolar, and bipolar, first introduced by van Oss *(12),* describe the hydrogen-bond polarity of compounds and must not be confused with other measures of polarity like the permanent dipole moment.

5. Qualitative rules for the partitioning of various compounds between given bulk phases and for the adsorption on surfaces will be discussed in a separate paper.

6. Here we neglect the effect of the mutual partitioning of water and the solvent.

7. The strength of a single H-bond between −OH groups is the same in different alcohols but the number density of H-bonds and thus the cohesive energy per volume increases from 1-butanol to 1,2-ethanediol.

Literature Cited

1. Park, J. H.; Hussam, A.; Couasnon, P.; Fritz, D.; Carr, P. W. *Anal. Chem.* 1987, *59,* 1970–1976.
2. Ruelle, P.; Kesselring, U. W. *Chemosphere* 1997, *34,* 275–298.
3. Abraham, M. H.; Andonian-Haftvan, J.; Whiting, G. S.; Leo, A.; Taft, R. S. *J. Chem. Soc., Perkin Trans. 2* 1994, 1777–1791.
4. Huuskonen, J.; Salo, M.; Taskinen, J. *J. Chem. Inf. Comput. Sci.* 1998, *38,* 450–456.
5. Lyman, W. J.; Reehl, W. F.; Rosenblatt, D. H. *Handbook of Chemical Property Estimation Methods;* American Chemical Society: New York, 1990.
6. Aydt, E. M.; Hentschke, R. *Ber. Bunsen-Ges. Phys. Chem.* 1997, *101,* 79–83.
7. Abraham, M. H.; Grellier, P. L.; McGill, R. A. *J. Chem. Soc., Perkin Trans. 2* 1987, 797–803.
8. Fowkes, F. M. In *Acid-Base Interactions—Relevance to Adhesion Science and Technology*; Mittal, K. L., Anderson, H. R., Eds.; VSP: Utrecht, 1991; pp 93–115.
9. Good, R. J.; Chaudhury, M. K. In *Fundamentals of Adhesion*; Lee, L.-H., Ed.; Plenum Press: New York, 1991; pp 137–172.
10. Fowkes, F. M. *Ind. Eng. Chem.* 1964, *56,* 40–52.
11. Ben-Naim, A. *J. Phys. Chem.* 1978, *82,* 792–803.
12. van Oss, C. J. *Interfacial Forces in Aqueous Media*; Marcel Dekker: New York, 1994.
13. Hansch, C.; Leo, A. *Substituent Constants for Correlation Analysis in Chemistry and Biology*; John Wiley & Sons: New York, 1979.
14. Abraham, M. H.; Platts, J. A.; Hersey, A.; Leo, A. J.; Taft, R. W. *J. Pharm. Sci. 1999, 88,* 670–679.
15. Abraham, M. H.; Poole, C. F.; Poole, S. K. *J. Chromatogr., A* 1999, *842,* 79–114.
16. Abraham, M. H.; Chadha, H. S.; Dixon, J. P. *J. Phys. Org. Chem.* 1994, *7,* 712–716.
17. Goss, K.-U.; Schwarzenbach, R. P. *Environ. Sci. Technol.* 2001, *35,* 1–9.