

# TOPIC 14E Self-assembly

## ► Why do you need to know this material?

Aggregates of small and large molecules form the basis of many established and emerging technologies. To see why this is the case, you need to understand their structures and properties.

## ► What is the key idea?

Colloids and micelles form spontaneously by self-assembly of molecules or macromolecules and are held together by molecular interactions.

## ► What do you need to know already?

You need to be familiar with molecular interactions (Topic 14B) and interactions between ions (Topic 5E).

**Self-assembly** is the spontaneous formation of complex structures of molecules or macromolecules that are held together by molecular interactions, such as Coulombic, dispersion, hydrogen bonding, and hydrophobic interactions. Examples of self-assembly include the formation of liquid crystals, and of protein quaternary structures from two or more polypeptide chains (Topic 14C).

## 14E.1 Colloids

A **colloid**, or **disperse phase**, is a dispersion of small particles of one material in another that does not settle out under gravity. In this context, 'small' means that one dimension at least is smaller than about 500 nm (about the wavelength of visible light). Many colloids are suspensions of nanoparticles (particles of diameter up to about 100 nm). In general, colloidal particles are aggregates of numerous atoms or molecules, but are commonly but not universally too small to be seen with an ordinary optical microscope.

### (a) Classification and preparation

The name given to the colloid depends on the two phases involved:

- A **sol** is a dispersion of a solid in a liquid (such as clusters of gold atoms in water) or of a solid in a solid (such as

ruby glass, which is a gold-in-glass sol, and achieves its colour by light scattering).

- An **aerosol** is a dispersion of a liquid in a gas (like fog and many sprays) or a solid in a gas (such as smoke): the particles are often large enough to be seen with a microscope.
- An **emulsion** is a dispersion of a liquid in a liquid (such as milk).
- A **foam** is a dispersion of a gas in a liquid.

A further classification of colloids is as **lyophilic**, or solvent attracting, and **lyophobic**, solvent repelling. If the solvent is water, the terms **hydrophilic** and **hydrophobic**, respectively, are used instead. Lyophobic colloids include the metal sols. Lyophilic colloids generally have some chemical similarity to the solvent, such as  $-OH$  groups able to form hydrogen bonds. A **gel** is a semi-rigid mass of a lyophilic sol.

The preparation of aerosols can be as simple as sneezing (which produces an imperfect aerosol). Laboratory and commercial methods make use of several techniques. Material (e.g. quartz) may be ground in the presence of the dispersion medium. Passing a heavy electric current through a cell may lead to the sputtering (crumbling) of an electrode into colloidal particles. Arcing between electrodes immersed in the support medium also produces a colloid. Chemical precipitation sometimes results in a colloid. A precipitate (e.g. silver iodide) already formed may be dispersed by the addition of a 'peptizing agent' (e.g. potassium iodide). Clays may be peptized by alkalis, the  $OH^-$  ion being the active agent.

Emulsions are normally prepared by shaking the two components together vigorously, although some kind of emulsifying agent usually has to be added to stabilize the product. This emulsifying agent may be a soap (the salt of a long-chain carboxylic acid) or other **surfactant** (surface active) species, or a lyophilic sol that forms a protective film around the dispersed phase. In milk, which is an emulsion of fats in water, the emulsifying agent is casein, a protein containing phosphate groups. It is clear from the formation of cream on the surface of milk that casein is not completely successful in stabilizing milk: the dispersed fats coalesce into oily droplets which float to the surface. This coagulation may be prevented by ensuring that the emulsion is dispersed very finely initially: intense agitation with ultrasonics brings this dispersion about, the product being 'homogenized' milk.

One way to form an aerosol is to tear apart a spray of liquid with a jet of gas. The dispersal is aided if a charge is applied to the liquid, for then electrostatic repulsions help to blast it

apart into droplets. This procedure may also be used to produce emulsions, for the charged liquid phase may be directed into another liquid.

Colloids are often purified by **dialysis**, the process of squeezing the solution through a membrane. The aim is to remove much (but not all, for reasons explained later) of the ionic material that may have accompanied their formation. A membrane (for example, cellulose) is selected that is permeable to solvent and ions, but not to the colloid particles. Dialysis is very slow, and is normally accelerated by applying an electric field and making use of the charges carried by many colloidal particles; the technique is then called **electrodialysis**.

### (b) Structure and stability

Colloids are thermodynamically unstable with respect to the bulk. This instability can be expressed thermodynamically by noting that because the change in Helmholtz energy,  $dA$ , when the surface area of the sample changes by  $d\sigma$  at constant temperature and pressure is  $dA = \gamma d\sigma$ , where  $\gamma$  is the interfacial surface tension (Topic 14C), it follows that  $dA < 0$  if  $d\sigma < 0$ . That is, the contraction of the surface ( $d\sigma < 0$ ) is spontaneous ( $dA < 0$ ). The survival of colloids must therefore be a consequence of the kinetics of collapse: colloids are thermodynamically unstable but kinetically non-labile.

At first sight, even the kinetic argument seems to fail: colloidal particles attract each other over large distances, so there is a long-range force that tends to condense them into a single blob. The reasoning behind this remark is as follows. The energy of attraction between two individual atoms  $i$  and  $j$  separated by a distance  $R_{ij}$ , one in each colloidal particle, varies with their separation as  $1/R_{ij}^6$  (Topic 14B). The sum of all these pairwise interactions, however, decreases only as approximately  $1/R^2$  (the precise variation depending on the shape of the particles and their closeness), where  $R$  is the separation of the centres of the particles. The change in the power from 6 to 2 stems from the fact that at short distances only a few molecules interact but at large distances many individual molecules are at about the same distance from one another, and contribute equally to the sum (Fig. 14E.1), so the total interaction does not fall off as fast as the single molecule–molecule interaction.

Several factors oppose the long-range dispersion attraction. For example, there may be a protective film at the surface of the colloid particles that stabilizes the interface and cannot be penetrated when two particles touch. Thus, the surface atoms of a platinum sol in water react chemically and are turned into  $-\text{Pt}(\text{OH})_3\text{H}_3$ ; this layer encases the particle like a shell. A fat can be emulsified by a soap because the long hydrocarbon tails penetrate the oil droplet but the carboxylate head groups (or other hydrophilic groups in synthetic detergents) surround the surface, form hydrogen bonds with water, and give rise to a shell of negative charge that repels a possible approach from another similarly charged particle.

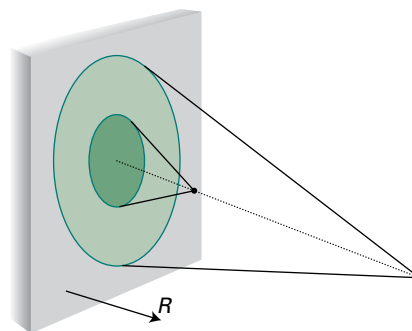


Figure 14E.1 Although the attraction between individual molecules is proportional to  $1/R^6$ , more molecules are within range at large separations (pale region) than at small separation (dark region), so the total interaction energy declines more slowly and is proportional to a lower power of  $1/R$ .

### (c) The electrical double layer

A major source of kinetic non-lability of colloids is the existence of an electric charge on the surfaces of the particles. Ions of opposite charge tend to cluster near each other, and form an ionic atmosphere around the particles, just as for individual ions (Topic 5F).

There are two regions of charge. First, there is a fairly immobile layer of ions that adhere tightly to the surface of the colloidal particle, and which may include water molecules (if that is the support medium). The radius of the sphere that captures this rigid layer is called the **radius of shear** and is the major factor determining the mobility of the particles. The electric potential at the radius of shear relative to its value in the distant, bulk medium is called the **electrokinetic potential**,  $\zeta$  (or the *zeta potential*). Second, the charged unit attracts an oppositely charged atmosphere of mobile ions. The inner shell of charge and the outer ionic atmosphere constitute the **electrical double layer**.

The theory of the stability of lyophobic dispersions was developed by B. Derjaguin and L. Landau and independently by E. Verwey and J.T.G. Overbeek, and is known as the **DLVO theory**. It assumes that there is a balance between the repulsive interaction between the charges of the electrical double layers on neighbouring particles and the attractive interactions arising from van der Waals interactions between the molecules in the colloidal particles. The potential energy arising from the repulsion of double layers on particles of radius  $a$  has the form

$$V_{\text{repulsion}} = + \frac{Aa^2\zeta^2}{R} e^{-s/r_D} \quad (14E.1)$$

where  $A$  is a constant,  $\zeta$  is the zeta potential,  $R$  is the separation of centres,  $s$  is the separation of the surfaces of the two particles ( $s = R - 2a$  for spherical particles of radius  $a$ ), and  $r_D$  is the thickness of the double layer. This expression is valid for

small particles with a thick double layer ( $r_D \gg a$ ). When the double layer is thin ( $r_D \ll a$ ), the expression is replaced by

$$V_{\text{repulsion}} = +\frac{1}{2}Ba^2\zeta^2 \ln(1+e^{-s/r_D}) \quad (14E.2)$$

where  $B$  is another constant. In each case, the thickness of the double layer can be estimated from an expression like that derived for the thickness of the ionic atmosphere in the Debye–Hückel theory (Topic 5F and *A deeper look* 1 on the website for this text) in which there is a competition between the assembling influences of the attraction between opposite charges and the disruptive effect of thermal motion:

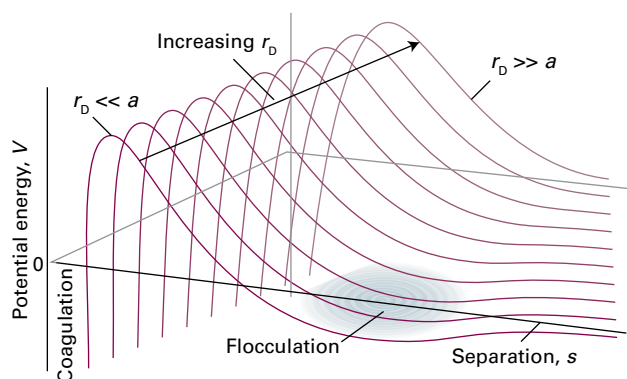
$$r_D = \left( \frac{\epsilon RT}{2\rho F^2 I b^\ominus} \right)^{1/2} \quad \text{Thickness of the electrical double layer} \quad (14E.3)$$

where  $I$  is the ionic strength of the solution (eqn 5F.28,  $I = \frac{1}{2} \sum_i z_i^2 b_i / b^\ominus$  with  $b^\ominus = 1 \text{ mol kg}^{-1}$ ) and  $\rho$  its mass density. As usual,  $F$  is Faraday's constant and  $\epsilon$  is the permittivity,  $\epsilon = \epsilon_r \epsilon_0$ . The potential energy arising from the attractive interaction has the form

$$V_{\text{attraction}} = -\frac{C}{s} \quad (14E.4)$$

where  $C$  is yet another constant. The variation of the total potential energy with separation is shown in Fig. 14E.2.

At high ionic strengths, the ionic atmosphere is dense and the potential shows a secondary minimum at large separations. Aggregation of the particles arising from the stabilizing effect of this secondary minimum is called **flocculation**. The flocculated material can often be redispersed by agitation because the well is so shallow. **Coagulation**, the irreversible aggregation of distinct particles into large particles, occurs when the separation of the particles is so small that they enter the primary minimum of the potential energy curve and van der Waals forces are dominant.



**Figure 14E.2** The variation of the potential energy of interaction with separation of the centres of the two particles and with the ratio of the particle size  $a$  to the thickness of the electrical double layer,  $r_D$ . The regions labelled coagulation and flocculation show the dips in the potential energy curves where these processes occur.

The ionic strength is increased by the addition of ions, particularly those of high charge type, so such ions act as flocculating agents. This increase is the basis of the empirical **Schulze–Hardy rule**, that hydrophobic colloids are flocculated most efficiently by ions of opposite charge type and high charge number. The  $\text{Al}^{3+}$  ions in alum are very effective, and are used to induce the congealing of blood. When river water containing colloidal clay flows into the sea, the salt water induces flocculation and coagulation, and is a major cause of silting in estuaries.

Metal oxide sols tend to be positively charged whereas sulfur and the noble metals tend to be negatively charged. Naturally occurring macromolecules also acquire a charge when dispersed in water, and an important feature of proteins and other natural macromolecules is that their overall charge depends on the pH of the medium. For instance, in acidic environments protons attach to basic groups, and the net charge of the macromolecule is positive; in basic media the net charge is negative as a result of proton loss. At the **isoelectric point** the pH is such that there is no net charge on the macromolecule.

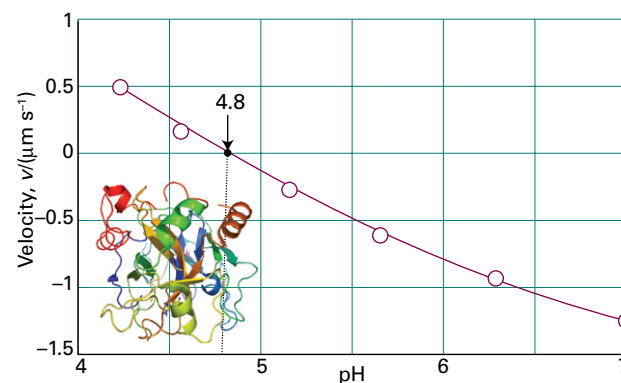
#### Example 14E.1 Determining the isoelectric point of a protein

The velocity with which the protein bovine serum albumin (BSA) moves through water under the influence of an electric field was monitored at several values of pH, and the data are listed below. What is the isoelectric point of the protein?

|                                  |      |      |       |       |       |       |
|----------------------------------|------|------|-------|-------|-------|-------|
| pH                               | 4.20 | 4.56 | 5.20  | 5.65  | 6.30  | 7.00  |
| Velocity/ $(\mu\text{m s}^{-1})$ | 0.50 | 0.18 | -0.25 | -0.65 | -0.90 | -1.25 |

**Collect your thoughts** Plot velocity against pH, then use interpolation to find the pH at which the velocity is zero, which is the pH at which the molecule has zero net charge.

**The solution** The data are plotted in Fig. 14E.3. The velocity passes through zero at  $\text{pH} = 4.8$ ; hence  $\text{pH} = 4.8$  is the isoelectric point.



**Figure 14E.3** The plot of the velocity of a moving macromolecule against pH allows the isoelectric point to be detected as the pH at which the velocity is zero. The data are from *Example 14E.1*.

**Self-test 14E.1** The following data were obtained for another protein:

|                                   |      |       |       |       |       |
|-----------------------------------|------|-------|-------|-------|-------|
| pH                                | 3.5  | 4.5   | 5.0   | 5.5   | 6.0   |
| Velocity/( $\mu\text{m s}^{-1}$ ) | 0.10 | -0.10 | -0.20 | -0.30 | -0.40 |

Estimate the pH of the isoelectric point.

Answer: 4.0

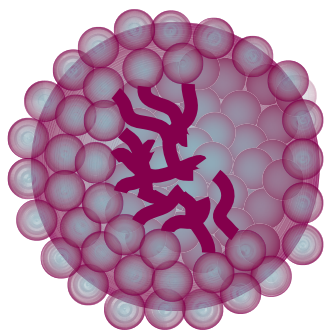
The primary role of the electrical double layer is to confer kinetic non-lability. Colliding colloidal particles break through the double layer and coalesce only if the collision is sufficiently energetic to disrupt the layers of ions and solvating molecules, or if thermal motion has stirred away the surface accumulation of charge. This disruption may occur at high temperatures, which is one reason why sols precipitate when they are heated.

## 14E.2 Micelles and biological membranes

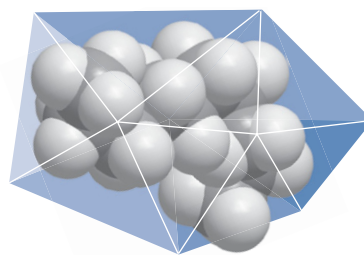
In aqueous solutions surfactant molecules or ions can cluster together as **micelles**, which are colloid-sized clusters of molecules, for their hydrophobic tails tend to congregate, and their hydrophilic head groups provide protection (Fig. 14E.4).

### (a) The hydrophobic interaction

Consider a long-chained alcohol, such as pentan-1-ol ( $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ ). The hydrocarbon chain is hydrophobic and the  $-\text{OH}$  group is hydrophilic. A species with both



**Figure 14E.4** A schematic version of a spherical micelle. The hydrophilic groups are represented by spheres and the hydrophobic hydrocarbon chains are represented by the stalks; these stalks are mobile.



**Figure 14E.5** When a hydrocarbon molecule is surrounded by water, the  $\text{H}_2\text{O}$  molecules form a cage. As a result of this acquisition of structure, the entropy of the water decreases, so the dispersal of the hydrocarbon into the water is accompanied by a local decrease in entropy. However, the aggregation of these individual caged hydrocarbon molecules into a micelle releases many of the caging water molecules back into the bulk and results in an increase in entropy.

hydrophobic and hydrophilic regions is called **amphipathic**.<sup>1</sup> Amphipathic substances do dissolve slightly in water, and an understanding of the process gives insight into the formation of micelles and biological structures in general.

To understand the dissolution process in more detail, imagine a hypothetical initial state in which the alcohol is present in water as individual molecules. Each hydrophobic chain is surrounded by a cage of water molecules (Fig. 14E.5). This order reduces the entropy of the water below its 'pure' value. Now consider the final state, in which the hydrophobic chains have clustered together. Although the clustering contributes to the lowering of the entropy of the system, fewer (but larger) cages are required, and more water molecules are free to move. The net effect of the formation of clusters of hydrophobic chains is therefore a decrease in the organization of water molecules and therefore a *net increase* in entropy of the system. This increase in entropy of the solvent (water) means that the association of hydrophobic groups in an aqueous environment is spontaneous (provided there are no overwhelming enthalpy effects). This spontaneous clustering of hydrophobic groups in the presence of water *gives the appearance* of it being the outcome of an actual intermolecular force and is called the **hydrophobic interaction**.

Some insight into the processes involved can be obtained from studies of the thermodynamics of dissolving (as distinct from micelle formation). The entropy of dissolution of largely hydrophobic molecules in water is positive ( $\Delta_{\text{diss}} S^\ominus > 0$ ) as the molecules disperse and the structure of the water changes to accommodate them. The process is commonly endothermic ( $\Delta_{\text{diss}} H^\ominus > 0$ ), but the Gibbs energy of dissolution ( $\Delta_{\text{diss}} G^\ominus$ ) is typically negative, as is illustrated by the following data (at 298 K):

<sup>1</sup> The *amphi-* part of the name is from the Greek word for 'both', and the *-pathic* part is from the same root (meaning 'feeling') as *sympathetic*.

|  | $\Delta_{\text{diss}}G^\ominus /$<br>(kJ mol <sup>-1</sup> ) | $\Delta_{\text{diss}}H^\ominus /$<br>(kJ mol <sup>-1</sup> ) | $\Delta_{\text{diss}}S^\ominus /$<br>(JK <sup>-1</sup> mol <sup>-1</sup> ) |
|--|--|--|--|
| CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH                 | -10  | +8   | +61  |
| CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH | -13  | +8   | +70  |

In other words, the tendency to dissolve (at least to a small extent) is entropy-driven, with contributions from the dispersion of the solute molecules and the restructuring of the water. Once dissolved, further reorganization of the water occurs to drive the formation of micelles. The experimental values are consistent with a general rule that each additional -CH<sub>2</sub>-group contributes a further -3 kJ mol<sup>-1</sup> to the standard Gibbs energy of dissolution.

A further aspect of this discussion is that it is possible to establish a scale of hydrophobicities. The hydrophobicity of a small molecular group R is reported by defining the **hydrophobicity constant**,  $\pi$ , as

$$\pi = \log \frac{s(\text{RX})}{s(\text{HX})} \quad \text{Hydrophobicity constant [definition]} \quad (14E.5)$$

where  $s(\text{RX})$  is the ratio of the molar solubility of the hydrophobic compound RX in the largely hydrocarbon solvent octan-1-ol to that in water, and  $s(\text{HX})$  is the analogous ratio for the compound HX. A positive value of  $\pi$  indicates that RX is more hydrophobic than RH.

It is found that the  $\pi$  values of most compounds do not depend on the identity of X (which might be OH, NH<sub>2</sub>, and so on). However, measurements suggest that  $\pi$  increases by the same amount each time a CH<sub>2</sub> group is added:

| -R    | -CH <sub>3</sub> | -CH <sub>2</sub> CH <sub>3</sub> | -(CH <sub>2</sub> ) <sub>2</sub> CH <sub>3</sub> | -(CH <sub>2</sub> ) <sub>3</sub> CH <sub>3</sub> | -(CH <sub>2</sub> ) <sub>4</sub> CH <sub>3</sub> |
|-------|------------------|----------------------------------|--|--|--|
| $\pi$ | 0.5              | 1.0                              | 1.5  | 2.0  | 2.5  |

It follows that acyclic saturated hydrocarbons become more hydrophobic as the carbon chain length increases. This trend can be rationalized by noting that  $\Delta_{\text{diss}}G^\ominus$  becomes more negative as the number of carbon atoms in the chain increases, with the data on butan-1-ol and pentan-1-ol (see above) suggesting that the principal effect is due to the entropy.

## (b) Micelle formation

Micelles form only above the **critical micelle concentration** (CMC) and above the **Krafft temperature**. The CMC is detected by noting a pronounced change in physical properties of the solution, particularly the molar conductivity (Fig. 14E.6). There is no abrupt change in some properties at the CMC; rather, there is a transition region corresponding to a range of concentrations around the CMC where physical properties vary smoothly but nonlinearly with the concentra-

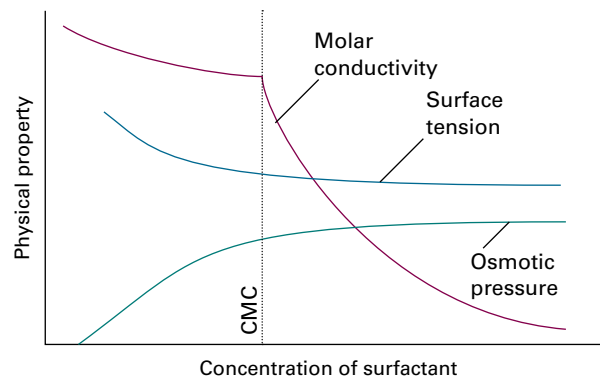
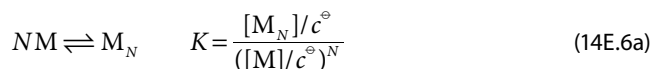


Figure 14E.6 The typical variation of some physical properties of an aqueous solution of sodium dodecyl sulfate close to the critical micelle concentration (CMC).

tion. The hydrocarbon interior of a micelle is like a droplet of oil. Nuclear magnetic resonance shows that the hydrocarbon tails are mobile, but slightly more restricted than in the bulk. Micelles are important in industry and biology on account of their solubilizing function: matter can be transported by water after it has been dissolved in their hydrocarbon interiors. For this reason, micellar systems are used as detergents, for organic synthesis, froth flotation for the treatment of ores, and petroleum recovery.

The self-assembly of a micelle has the characteristics of a cooperative process in which the addition of a surfactant molecule to a cluster that is forming becomes more probable the larger the size of the aggregate, so after a slow start there is a cascade of formation of micelles. If it is supposed that the dominant micelle  $M_N$  consists of  $N$  monomers  $M$ , then the dominant equilibrium to consider is



where it has been assumed, probably dangerously on account of the large sizes of monomers, that the solution is ideal and that activities can be replaced by molar concentrations. The total concentration of surfactant,  $[M]_{\text{total}}$ , is  $[M] + N[M_N]$  because each micelle consists of  $N$  monomer molecules. Therefore (omitting the  $c^\ominus$  for clarity),

$$K = \frac{[M_N]}{([M]_{\text{total}} - N[M_N])^N} \quad (14E.6b)$$

### Brief illustration 14E.1

Equation 14E.6b can be solved numerically for the variation of the fraction of molecules present as micelles with the number of molecules present in a micelle and some results for  $K = 1$  are shown in Fig. 14E.7. For large  $N$ , there is a reasonably sharp transition in the fractions of surfactant molecules that are present in micelles, which corresponds to the existence of a CMC.

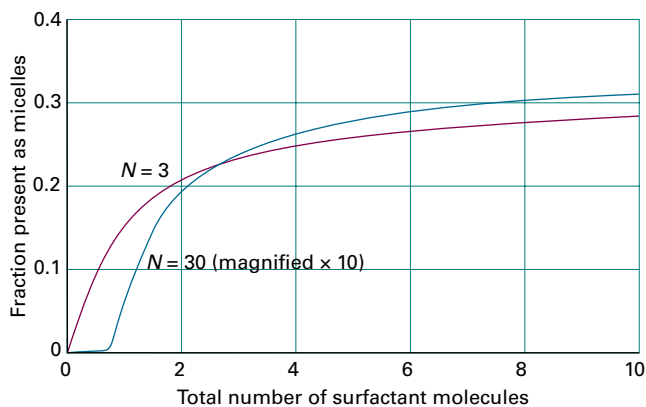


Figure 14E.7 The dependence of the fraction of surfactant molecules present as micelles on the number of molecules in the micelle for  $K = 1$ .

Non-ionic surfactant molecules may cluster together in clumps of 1000 or more, but ionic species tend to be disrupted by the electrostatic repulsions between head groups and are normally limited to groups of less than about 100. However, the disruptive effect depends more on the effective size of the head group than the charge. For example, ionic surfactants such as sodium dodecyl sulfate (SDS) and cetyltrimethylammonium bromide (CTAB) form rods at moderate concentrations whereas sugar surfactants form small, approximately spherical micelles. The micelle population commonly spans a wide range of particle sizes (i.e. it is polydisperse), and the shapes of the individual micelles vary with shape of the constituent surfactant molecules, surfactant concentration, and temperature. A useful predictor of the shape of the micelle is the **surfactant parameter**,  $N_s$ , defined as

$$N_s = \frac{V}{Al} \quad \text{Surfactant parameter [definition]} \quad (14E.7)$$

where  $V$  is the volume of the hydrophobic surfactant tail,  $A$  is the area of the hydrophilic surfactant head group, and  $l$  is the maximum length of the surfactant tail. Table 14E.1 summarizes the dependence of aggregate structure on the surfactant parameter.

In aqueous solutions spherical micelles form, as shown in Fig. 14E.4, with the polar head groups of the surfactant mol-

Table 14E.1 Micelle shape and the surfactant parameter

| $N_s$     | Micelle shape                     |
|-----------|-----------------------------------|
| <0.33     | Spherical                         |
| 0.33–0.50 | Cylindrical rods                  |
| 0.50–1.00 | Vesicles                          |
| 1.00      | Planar bilayers                   |
| >1.00     | Reverse micelles and other shapes |

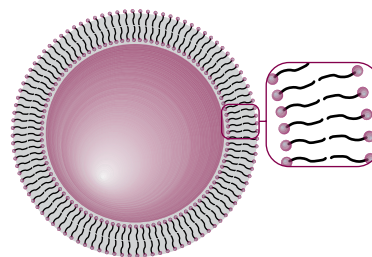


Figure 14E.8 The cross-sectional structure of a spherical liposome.

ecules on the micellar surface and interacting favourably with solvent and ions in solution. Hydrophobic interactions stabilize the aggregation of the hydrophobic surfactant tails in the micellar core. Under certain experimental conditions, a **liposome** may form, with an inward pointing inner surface of molecules surrounded by an outward pointing outer layer (Fig. 14E.8). Liposomes may be used to carry nonpolar drug molecules in blood.

Increasing the ionic strength of the aqueous solution reduces repulsions between surface head groups, and cylindrical micelles can form. These cylinders may stack together in reasonably close-packed (hexagonal) arrays, forming **lyotropic mesomorphs** and, more colloquially, ‘liquid crystalline phases’.

**Reverse micelles** form in nonpolar solvents, with small polar surfactant head groups in a micellar core and more voluminous hydrophobic surfactant tails extending into the organic bulk phase. These spherical aggregates can solubilize water in organic solvents by creating a pool of trapped water molecules in the micellar core. As aggregates arrange at high surfactant concentrations to yield long-range positional order, many other types of structures are possible including cubic and hexagonal shapes.

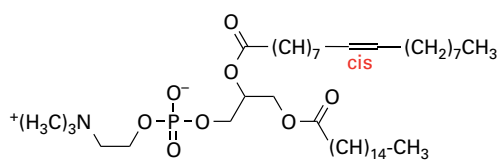
As already noted, micelle formation is driven by hydrophobic interactions. The enthalpy of formation reflects contributions of interactions between micelle chains within the micelles and between the polar head groups and the surrounding medium. Consequently, enthalpies of micelle formation display no readily discernible pattern and may be positive (endothermic) or negative (exothermic). Many non-ionic micelles form endothermically, with  $\Delta H$  of the order of 10 kJ per mole of surfactant molecules. That such micelles do form above the CMC indicates that the entropy change accompanying their formation must then be positive, and measurements suggest a value of about  $+140 \text{ J K}^{-1} \text{ mol}^{-1}$  at room temperature.

### (c) Bilayers, vesicles, and membranes

Some micelles at concentrations well above the CMC form extended parallel sheets two molecules thick, called **planar bilayers**. The individual molecules lie perpendicular to the sheets, with hydrophilic groups on the outside in aqueous so-

lution and on the inside in nonpolar media. When segments of planar bilayers fold back on themselves, **unilamellar vesicles** may form where the spherical hydrophobic bilayer shell separates an inner aqueous compartment from the external aqueous environment.

Bilayers show a close resemblance to biological membranes, and are often a useful model on which to base investigations of biological structures. However, actual membranes are highly sophisticated structures. The basic structural element of a membrane is a phospholipid, such as phosphatidyl choline (1), which contains long hydrocarbon chains (typically in the range  $C_{14}$ – $C_{24}$ ) and a variety of polar groups, such as  $-\text{CH}_2\text{CH}_2\text{N}(\text{CH}_3)_3^+$ . The hydrophobic chains stack together to form an extensive layer about 5 nm across. The lipid molecules form layers instead of micelles because the hydrocarbon chains are too bulky to allow packing into nearly spherical clusters.



1 Phosphatidyl choline

The bilayer is a highly mobile structure. Not only are the hydrocarbon chains ceaselessly twisting and turning in the region between the polar groups, but the phospholipid molecules migrate over the surface. It is better to think of the membrane as a viscous fluid rather than a permanent structure, with a viscosity about 100 times that of water. Typically, a phospholipid molecule in a membrane migrates through about  $1\ \mu\text{m}$  in about 1 min.

All lipid bilayers undergo a transition from a state of high to low chain mobility at a temperature that depends on the structure of the lipid. To visualize the transition, consider what happens to a membrane as its temperature is lowered (Fig. 14E.9). There is sufficient energy available at normal temperatures for limited bond rotation to occur and the flexible chains writhe. However, the membrane is still highly organized in the sense that the bilayer structure does not come apart and the system is best described as a liquid crystal. At lower temperatures, the amplitudes of the writhing motion decrease until a specific temperature is reached at which motion is largely frozen. The membrane then exists as a gel. Biological membranes exist as liquid crystals at physiological temperatures.

Phase transitions in membranes are often observed as ‘melting’ from gel to liquid crystal by calorimetric methods. The

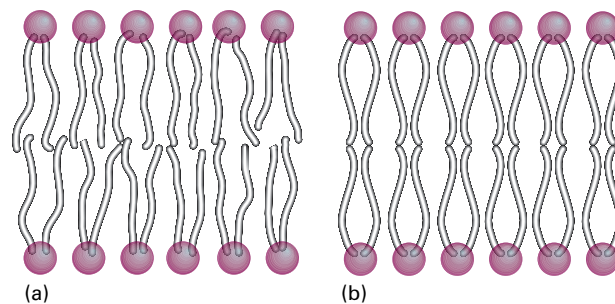
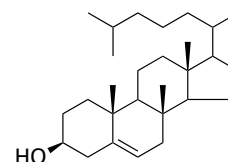


Figure 14E.9 A depiction of the variation with temperature of the flexibility of hydrocarbon chains in a lipid bilayer. (a) At physiological temperature, the bilayer exists as a liquid crystal, in which some order exists but the chains writhe. (b) At a specific temperature, the chains are largely frozen and the bilayer exists as a gel.

data show relations between the structure of the lipid and the melting temperature. Interspersed among the phospholipids of biological membranes are sterols, such as cholesterol (2), which is largely hydrophobic but does contain a hydrophilic  $-\text{OH}$  group. Sterols, which are present in different proportions in different types of cells, prevent the hydrophobic chains of lipids from ‘freezing’ into a gel and, by disrupting the packing of the chains, spread the melting point of the membrane over a range of temperatures.



2 Cholesterol

#### Brief illustration 14E.2

To predict trends in melting temperatures you need to assess the strengths of the interactions between molecules. Longer chains can be expected to be held together more strongly by hydrophobic interactions than shorter chains, so you should expect the melting temperature to increase with the length of the hydrophobic chain of the lipid. On the other hand, any structural elements that prevent alignment of the hydrophobic chains in the gel phase lead to low melting temperatures. Indeed, lipids containing unsaturated chains, those containing some  $\text{C}=\text{C}$  bonds, form membranes with lower melting temperatures than those formed from lipids with fully saturated chains, those consisting of  $\text{C}-\text{C}$  bonds only.

## Checklist of concepts

- 1. A **disperse system** is a dispersion of small particles of one material in another.
- 2. **Colloids** are classified as lyophilic and lyophobic.
- 3. A **surfactant** is a species that accumulates at the interface of two phases or substances.
- 4. Many colloidal particles are thermodynamically unstable but kinetically non-labile.
- 5. The **radius of shear** is the radius of the sphere that captures the rigid layer of charge attached to a colloid particle.
- 6. The **electrokinetic potential** is the electric potential at the radius of shear relative to its value in the distant, bulk medium.
- 7. The inner shell of charge and the outer atmosphere jointly constitute the **electrical double layer**.
- 8. **Flocculation** is the reversible aggregation of colloidal particles.
- 9. **Coagulation** is the irreversible aggregation of colloidal particles.
- 10. The **Schulze–Hardy rule** states that hydrophobic colloids are flocculated most efficiently by ions of opposite charge type and high charge number.
- 11. An **amphipathic** species has both hydrophobic and hydrophilic regions.
- 12. The **hydrophobic interaction** results in the clustering of nonpolar solutes in water.
- 13. A **micelle** is a colloid-sized cluster of molecules that forms at and above the **critical micelle concentration** and above the **Krafft temperature**.
- 14. **Unilamellar vesicles** are micelles that exist as extended parallel sheets.

## Checklist of equations

| Property                                 | Equation   | Comment             | Equation number |
|--|--|---------------------|-----------------|
| Thickness of the electrical double layer | $r_D = (\epsilon RT / 2 \rho F^2 I b^{\ominus})^{1/2}$ | Debye–Hückel theory | 14E.3           |
| Hydrophobicity constant                  | $\pi = \log\{s(\text{RX})/s(\text{HX})\}$              | Definition          | 14E.5           |
| Surfactant parameter                     | $N_s = V/AI$   | Definition          | 14E.7           |



# FOCUS 14 Molecular interactions

## TOPIC 14A The electric properties of molecules

### Discussion questions

**D14A.1** Explain how the permanent dipole moment and the polarizability of a molecule arise.

**D14A.2** Explain why the polarizability of a molecule decreases at high frequencies.

**D14A.3** Describe the experimental procedures available for determining the electric dipole moment of a molecule.

### Exercises

**E14A.1(a)** Which of the following molecules may be polar:  $\text{ClF}_3$ ,  $\text{O}_3$ ,  $\text{H}_2\text{O}_2$ ?

**E14A.1(b)** Which of the following molecules may be polar:  $\text{SO}_3$ ,  $\text{XeF}_4$ ,  $\text{SF}_4$ ?

**E14A.2(a)** Calculate the resultant of two dipole moments of magnitude 1.5 D and 0.80 D that make an angle of  $109.5^\circ$  to each other.

**E14A.2(b)** Calculate the resultant of two dipole moments of magnitude 2.5 D and 0.50 D that make an angle of  $120^\circ$  to each other.

**E14A.3(a)** Calculate the magnitude and direction of the dipole moment of the following arrangement of charges in the  $xy$ -plane:  $3e$  at  $(0, 0)$ ,  $-e$  at  $(0.32 \text{ nm}, 0)$ , and  $-2e$  at an angle of  $20^\circ$  from the  $x$ -axis and a distance of 0.23 nm from the origin.

**E14A.3(b)** Calculate the magnitude and direction of the dipole moment of the following arrangement of charges in the  $xy$ -plane:  $4e$  at  $(0, 0)$ ,  $-2e$  at  $(162 \text{ pm}, 0)$ , and  $-2e$  at an angle of  $30^\circ$  from the  $x$ -axis and a distance of 143 pm from the origin.

**E14A.4(a)** What strength of electric field is required to induce an electric dipole moment of magnitude  $1.0 \mu\text{D}$  in a molecule of polarizability volume  $2.6 \times 10^{-30} \text{ m}^3$  (like  $\text{CO}_2$ )?

**E14A.4(b)** What strength of electric field is required to induce an electric dipole moment of magnitude  $2.5 \mu\text{D}$  in a molecule of polarizability volume  $1.05 \times 10^{-29} \text{ m}^3$  (like  $\text{CCl}_4$ )?

**E14A.5(a)** The molar polarization of fluorobenzene vapour varies linearly with  $T^{-1}$ , and is  $70.62 \text{ cm}^3 \text{ mol}^{-1}$  at  $351.0 \text{ K}$  and  $62.47 \text{ cm}^3 \text{ mol}^{-1}$  at  $423.2 \text{ K}$ . Calculate the polarizability and dipole moment of the molecule.

**E14A.5(b)** The molar polarization of the vapour of a compound was found to vary linearly with  $T^{-1}$ , and is  $75.74 \text{ cm}^3 \text{ mol}^{-1}$  at  $320.0 \text{ K}$  and  $71.43 \text{ cm}^3 \text{ mol}^{-1}$  at  $421.7 \text{ K}$ . Calculate the polarizability and dipole moment of the molecule.

**E14A.6(a)** At  $0^\circ\text{C}$ , the molar polarization of liquid chlorine trifluoride is  $27.18 \text{ cm}^3 \text{ mol}^{-1}$  and its mass density is  $1.89 \text{ g cm}^{-3}$ . Calculate the relative permittivity of the liquid.

**E14A.6(b)** At  $0^\circ\text{C}$ , the molar polarization of a liquid is  $32.16 \text{ cm}^3 \text{ mol}^{-1}$  and its mass density is  $1.92 \text{ g cm}^{-3}$ . Calculate the relative permittivity of the liquid. Take  $M = 85.0 \text{ g mol}^{-1}$ .

**E14A.7(a)** The refractive index of  $\text{CH}_2\text{I}_2$  is 1.732 for 656 nm light. Its mass density at  $20^\circ\text{C}$  is  $3.32 \text{ g cm}^{-3}$ . Calculate the polarizability of the molecule at this wavelength.

**E14A.7(b)** The refractive index of a compound is 1.622 for 643 nm light. Its mass density at  $20^\circ\text{C}$  is  $2.99 \text{ g cm}^{-3}$ . Calculate the polarizability of the molecule at this wavelength. Take  $M = 65.5 \text{ g mol}^{-1}$ .

**E14A.8(a)** The polarizability volume of  $\text{H}_2\text{O}$  at optical frequencies is  $1.5 \times 10^{-24} \text{ cm}^3$ . Estimate the refractive index of water. The experimental value is 1.33.

**E14A.8(b)** The polarizability volume of a liquid of molar mass  $72.3 \text{ g mol}^{-1}$  and mass density  $865 \text{ kg m}^{-3}$  at optical frequencies is  $2.2 \times 10^{-30} \text{ m}^3$ . Estimate the refractive index of the liquid.

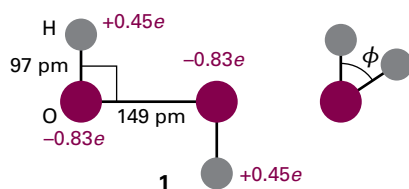
**E14A.9(a)** The dipole moment of chlorobenzene is 1.57 D and its polarizability volume is  $1.23 \times 10^{-23} \text{ cm}^3$ . Estimate its relative permittivity at  $25^\circ\text{C}$ , when its mass density is  $1.173 \text{ g cm}^{-3}$ .

**E14A.9(b)** The dipole moment of bromobenzene is  $5.17 \times 10^{-30} \text{ C m}$  and its polarizability volume is approximately  $1.5 \times 10^{-29} \text{ m}^3$ . Estimate its relative permittivity at  $25^\circ\text{C}$ , when its mass density is  $1491 \text{ kg m}^{-3}$ .

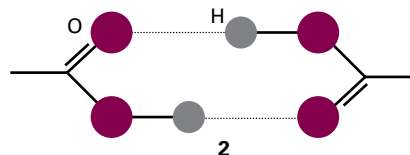
### Problems

**P14A.1** The electric dipole moment of methylbenzene (toluene) is 0.4 D. Estimate the dipole moments of the three isomers of dimethylbenzene (the xylenes). About which answer can you be sure?

**P14A.2** Plot the magnitude of the electric dipole moment of hydrogen peroxide as the H–O–H (azimuthal) angle  $\phi$  changes from 0 to  $2\pi$ . Use the dimensions and partial charges shown in (1).



**P14A.3** Ethanoic (acetic) acid vapour contains a proportion of planar, hydrogen-bonded dimers (2). The apparent dipole moment of molecules in pure gaseous ethanoic acid has a magnitude that increases with increasing temperature. Suggest an interpretation of this observation.



**P14A.4** D.D. Nelson et al. (*Science* 238, 1670 (1987)) examined several weakly bound gas-phase complexes of ammonia in search of examples in which the

\* These problems were supplied by Charles Trapp and Carmen Giunta.

H atoms in  $\text{NH}_3$  formed hydrogen bonds, but found none. For example, they found that the complex of  $\text{NH}_3$  and  $\text{CO}_2$  has the carbon atom nearest the nitrogen (299 pm away): the  $\text{CO}_2$  molecule is at right angles to the C–N 'bond', and the H atoms of  $\text{NH}_3$  are pointing away from the  $\text{CO}_2$ . The magnitude of the permanent dipole moment of this complex is reported as 1.77 D. If the N and C atoms are the centres of the negative and positive charge distributions, respectively, what is the magnitude of those partial charges (as multiples of  $e$ )?

**P14A.5** The polarizability volume of  $\text{NH}_3$  is  $2.22 \times 10^{-30} \text{ m}^3$ ; calculate the contribution to the dipole moment of the molecule induced by an applied electric field of strength  $15.0 \text{ kV m}^{-1}$ .

**P14A.6** The magnitude of the electric field at a distance  $r$  from a point charge  $Q$  is equal to  $Q/4\pi\epsilon_0 r^2$ . How close to a water molecule (of polarizability volume  $1.48 \times 10^{-30} \text{ m}^3$ ) must a proton approach before the dipole moment it induces has a magnitude equal to that of the permanent dipole moment of the molecule (1.85 D)?

**P14A.7** The relative permittivity of trichloromethane (chloroform) was measured over a range of temperatures with the following results:

|                           |      |      |      |      |      |      |      |
|---------------------------|------|------|------|------|------|------|------|
| $\theta/^\circ\text{C}$   | −80  | −70  | −60  | −40  | −20  | 0    | 20   |
| $\epsilon_r$              | 3.1  | 3.1  | 7.0  | 6.5  | 6.0  | 5.5  | 5.0  |
| $\rho/(\text{g cm}^{-3})$ | 1.65 | 1.64 | 1.64 | 1.61 | 1.57 | 1.53 | 1.50 |

The freezing point of trichloromethane is  $-64^\circ\text{C}$ . Account for these results and calculate the dipole moment and polarizability volume of the molecule.

**P14A.8** The relative permittivities of methanol (with a melting point of  $-95^\circ\text{C}$ ) corrected for density variation are given below. What molecular information can be deduced from these values? Take  $\rho = 0.791 \text{ g cm}^{-3}$ .

|                         |      |      |      |      |      |     |     |     |    |    |
|-------------------------|------|------|------|------|------|-----|-----|-----|----|----|
| $\theta/^\circ\text{C}$ | −185 | −170 | −150 | −140 | −110 | −80 | −50 | −20 | 0  | 20 |
| $\epsilon_r$            | 3.2  | 3.6  | 4.0  | 5.1  | 67   | 57  | 49  | 43  | 38 | 34 |

**P14A.9** In his classic book *Polar molecules*, Debye reports some early measurements of the polarizability of ammonia. From the selection below, determine the dipole moment and the polarizability volume of the molecule.

|                                      |       |       |       |       |       |       |
|--------------------------------------|-------|-------|-------|-------|-------|-------|
| $T/\text{K}$                         | 292.2 | 309.0 | 333.0 | 387.0 | 413.0 | 446.0 |
| $P_m/(\text{cm}^3 \text{ mol}^{-1})$ | 57.57 | 55.01 | 51.22 | 44.99 | 42.51 | 39.59 |

The refractive index of ammonia at 273 K and 100 kPa is 1.000 379 (for yellow sodium light). Calculate the molar polarization of the gas at this temperature. Combine the value calculated with the static molar polarization at 292.2 K and deduce from this information alone the molecular dipole moment.

**P14A.10** Values of the molar polarization of gaseous water at 100 kPa as determined from capacitance measurements are given below as a function of temperature.

|                                      |       |       |       |       |       |
|--------------------------------------|-------|-------|-------|-------|-------|
| $T/\text{K}$                         | 384.3 | 420.1 | 444.7 | 484.1 | 521.0 |
| $P_m/(\text{cm}^3 \text{ mol}^{-1})$ | 57.4  | 53.5  | 50.1  | 46.8  | 43.1  |

Calculate the dipole moment of  $\text{H}_2\text{O}$  and its polarizability volume.

**P14A.11** From data in Table 14A.1 calculate the molar polarization, relative permittivity, and refractive index of methanol at  $20^\circ\text{C}$ . Its mass density at that temperature is  $0.7914 \text{ g cm}^{-3}$ .

**P14A.12** Show that, in a gas (for which the refractive index is close to 1), the refractive index depends on the pressure as  $n_r = 1 + \text{constant} \times p$ , and find the constant of proportionality. Go on to show how to deduce the polarizability volume of a molecule from measurements of the refractive index of a gaseous sample.

**P14A.13** Ethanoic (acetic) acid vapour contains a proportion of planar, hydrogen-bonded dimers. The relative permittivity of pure liquid ethanoic acid is 7.14 at 290 K and increases with increasing temperature. Suggest an interpretation of the latter observation. What effect should isothermal dilution have on the relative permittivity of solutions of ethanoic acid in benzene?

## TOPIC 14B Interactions between molecules

### Discussion questions

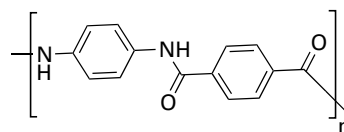
**D14B.1** Identify the terms in and of the following expressions and specify the conditions under which they are valid: (a)  $V = -Q_2\mu_1/4\pi\epsilon_0 r^2$ , (b)  $V = -Q_2\mu_1 \cos \theta/4\pi\epsilon_0 r^2$ , and (c)  $V = \mu_2\mu_1(1 - 3 \cos^2 \theta)/4\pi\epsilon_0 r^3$ .

**D14B.2** Draw examples of the arrangements of electrical charges that correspond to monopoles, dipoles, quadrupoles, and octupoles. Suggest a reason for the different distance dependencies of their electric fields.

**D14B.3** Account for the theoretical conclusion that many attractive interactions between molecules vary with their separation as  $1/r^6$ .

**D14B.4** Describe the formation of a hydrogen bond in terms of (a) electrostatic interactions and (b) molecular orbitals. How would you identify the better model?

**D14B.5** Some polymers have unusual properties. For example, Kevlar (3) is strong enough to be the material of choice for bulletproof vests and is stable at temperatures up to 600 K. What molecular interactions contribute to the formation and thermal stability of this polymer?



3 Kevlar

### Exercises

**E14B.1(a)** Calculate the molar energy required to reverse the direction of an  $\text{H}_2\text{O}$  molecule located 100 pm from a  $\text{Li}^+$  ion. Take the magnitude of the dipole moment of water as 1.85 D.

**E14B.1(b)** Calculate the molar energy required to reverse the direction of an  $\text{HCl}$  molecule located 300 pm from a  $\text{Mg}^{2+}$  ion. Take the magnitude of the dipole moment of  $\text{HCl}$  as 1.08 D.