13.2: Types of Solutions and Solubility

A solution may be composed of a solid and a liquid (such as the salt and water that are the primary components of seawater), but it may also be composed of a gas and a liquid, two different liquids, or other combinations (see Table 13.1). In **aqueous solutions**, water is the solvent, and a solid, liquid, or gas is the solute. For example, sugar water and saltwater are both aqueous solutions. Similarly, ethyl alcohol—the alcohol in alcoholic beverages—readily mixes with water to form a solution, and carbon dioxide dissolves in water to form the aqueous solution that we know as club soda.

Table 13.1 **Common Types of Solutions**

Solution Phase	Solute Phase	Solvent Phase	Example
Gaseous solution	Gas	Gas	Air (mainly oxygen and nitrogen)
Liquid solution	Gas	Liquid	Club soda (CO_2 and water)
	Liquid	Liquid	Vodka (ethanol and water)
	Solid	Liquid	Seawater (salt and water)
Solid solution	Solid	Solid	Brass (copper and zinc) and other alloys

You probably know from experience that a particular solvent, such as water, does not dissolve all possible solutes. For example, you cannot clean your greasy hands with just water because the water does not dissolve the grease. However, another solvent, such as paint thinner, can easily dissolve the grease. The grease is *insoluble* in water but *soluble* in the paint thinner. The **solubility** of a substance is the amount of the substance that will dissolve in a given amount of solvent. The solubility of sodium chloride in water at 25 °C is 36 g NaCl per 100 g water, while the solubility of grease in water is nearly zero. The solubility of one substance in another depends both on nature's tendency toward mixing and on the types of intermolecular forces that we discussed in Chapter 11.

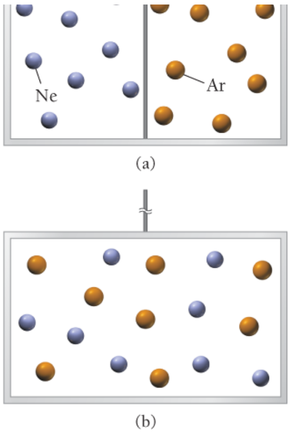
> The general solubilities of a number of ionic compounds are described by the solubility rules in Section 8.4.

Nature's Tendency toward Mixing: Entropy

So far in this book, we have seen that many physical systems tend toward lower *potential energy*. For example, two particles with opposite charges (such as a proton and an electron or a cation and an anion) move toward each other because their potential energy decreases as their separation decreases according to Coulomb's law (see Section 3.3). The formation of a solution, however, *does not necessarily* lower the potential energy of its constituent particles.

The clearest example of this phenomenon is the formation of a homogeneous mixture (a *solution*) of two ideal gases. Suppose that we enclose neon and argon in a container with a removable barrier between them, as shown in Figure 13.2(a). As soon as we remove the barrier, the neon and argon mix together to form a solution, as shown in Figure 13.2(b). *Why*?

Figure 13.2 **Spontaneous Mixing of Two Ideal Gases**



Neon and argon are separated by a barrier.

When the barrier is removed, the gases spontaneously mix to form a uniform solution.

At low pressures and moderate temperatures, both neon and argon behave as ideal gases—they do not interact with each other in any way (that is, there are no significant intermolecular forces between their constituent particles). When the barrier is removed, the two gases mix, but their potential energy remains unchanged. In other words, *we cannot think of the mixing of two ideal gases as lowering their potential energy*. Rather, the tendency to mix is related to a concept called *entropy*.

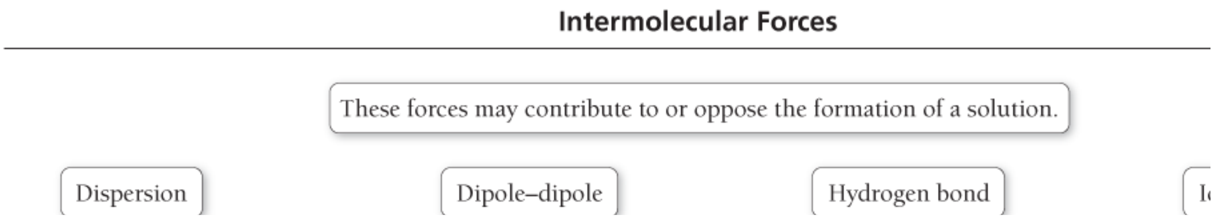
Entropy is a measure of *energy randomization* or *energy dispersal* in a system. Recall that a gas at any temperature above 0 K has kinetic energy due to the motion of its atoms. When neon and argon are confined to their individual compartments, their kinetic energies are also confined to those compartments. When the barrier between the compartments is removed, each gas—along with its kinetic energy—becomes *spread out* or *dispersed* over a larger volume. Thus, the mixture of the two gases has greater energy dispersal, or greater *entropy*, than the separated components.

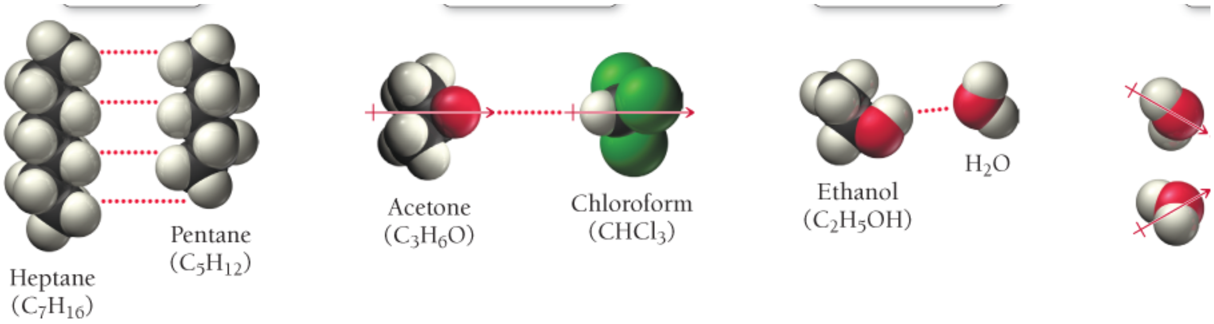
The pervasive tendency for energy to spread out, or disperse, whenever it is not restrained from doing so is the reason that two ideal gases mix. Another common example of the tendency toward energy dispersal is the transfer of thermal energy from hot to cold. If we heat one end of an iron rod, the thermal energy deposited at the end of the rod will spontaneously spread along the entire length of the rod. In similarity to the mixing of two ideal gases—where the kinetic energy of the particles becomes dispersed over a larger volume as the particles become dispersed—the thermal energy in the rod, initially concentrated in relatively fewer particles, becomes dispersed as it is distributed over a larger number of particles. The tendency for energy to disperse is why thermal energy flows from the hot end of the rod to the cold one, and not the other way around. Imagine a metal rod that became spontaneously hotter on one end and ice cold on the other. This does not happen because energy does not spontaneously concentrate itself. In Chapter 18, we will see that the dispersal of energy is actually the fundamental criterion that ultimately determines the spontaneity of any process.

The Effect of Intermolecular Forces

We have just seen that, in the absence of intermolecular forces, two substances spontaneously mix to form a homogeneous solution. We know from Chapter 11, however, that solids and liquids exhibit a number of different types of intermolecular forces, including dispersion forces, dipole–dipole forces, hydrogen bonding, and ion–dipole forces (Figure 13.3). These forces, and the possible localized structures that result from them, may promote the formation of a solution or prevent it, depending on the nature of the forces in the particular combination of solute and solvent.

Figure 13.3 **Intermolecular Forces Involved in Solutions**


Intermolecular Forces
These forces may contribute to or oppose the formation of a solution.
Dispersion
Dipole–dipole
Hydrogen bond



Intermolecular forces exist between (a) the solvent and solute particles, (b) the solvent particles themselves, and (c) the solute particles themselves, as shown in Figure 13.4. Each of the interactions is defined here:

Solvent–solute interactions:	The interactions between a solvent particle and a solute particle.
Solvent–solvent interactions:	The interactions between a solvent particle and another solvent particle.
Solute–solute interactions:	The interactions between a solute particle and another solute particle.

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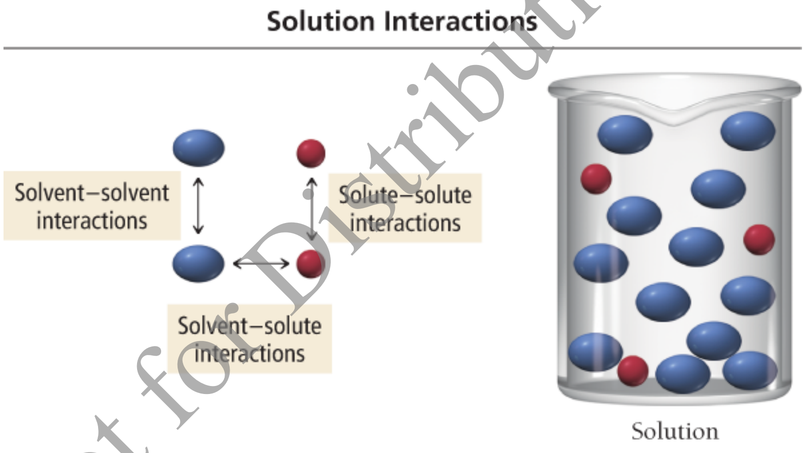
Figure 13.4 Forces in a Solution

The relative strengths of these three interactions determine whether a solution will form.


Solution Interactions
Solvent–solvent interactions
Solute–solute interactions
Solvent–solute interactions
Solution

As shown in Table 13.2, a solution generally forms if the solvent–solute interactions are comparable to, or stronger than, the solvent–solvent interactions and the solute–solute interactions. For example, consider the hydrocarbons pentane (C_5H_{12}) and heptane (C_7H_{16}). The intermolecular forces present within both pentane and heptane are dispersion forces. Similarly, the intermolecular forces present *between* heptane and pentane are also dispersion forces. All three interactions are of similar magnitude, so the two substances are soluble in each other in all proportions—they are said to be **miscible**. The formation of the solution is driven by the tendency toward mixing, or toward greater entropy, that we just discussed.

Table 13.2 Relative Interactions and Solution Formation

Solvent–solute interactions	$>$	Solvent–solvent and solute–solute interactions	Solution generally forms.*
Solvent–solute interactions	$=$	Solvent–solvent and solute–solute interactions	Solution generally forms.*
Solvent–solute interactions	$<$	Solvent–solvent and solute–solute interactions	Solution may or may not form, depending

on relative disparity.

*In some cases, especially solutions involving water, solvent structural changes that occur during solvation result in *decreases in entropy*, which may also prevent solution formation.

If solvent–solute interactions are weaker than solvent–solvent and solute–solute interactions—in other words, if solvent molecules and solute molecules each interact more strongly with molecules of their own kind than with molecules of the other kind—then a solution may still form, depending on the relative disparities between the interactions. If the disparity is small, the tendency to mix results in the formation of a solution even though the process is energetically uphill. If the disparity is large, however, a solution does not form.

In general, we can use the rule of thumb that *like dissolves like* when predicting the formation of solutions. Polar solvents, such as water, tend to dissolve many polar or ionic solutes, and nonpolar solvents, such as hexane, tend to dissolve many nonpolar solutes. Similar kinds of solvents dissolve similar kinds of solutes. Table 13.3 lists some common polar and nonpolar laboratory solvents.

Table 13.3 **Common Laboratory Solvents**

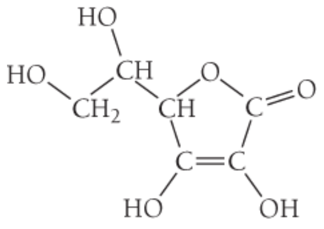
Common Polar Solvents	Common Nonpolar Solvents
Water (H_2O)	Hexane (C_6H_{14})
Acetone (CH_3COCH_3)	Diethyl ether ($CH_3CH_2OCH_2CH_3$)*
Methanol (CH_3OH)	Toluene (C_7H_8)
Ethanol (CH_3CH_2OH)	Carbon tetrachloride (CCl_4)

*Diethyl ether has a small dipole moment and can be considered intermediate between polar and nonpolar.

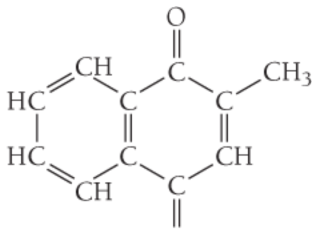
Example 13.1 Solubility

Vitamins are categorized as either fat soluble or water soluble. Water-soluble vitamins dissolve in body fluids and are easily eliminated in the urine, so there is little danger of overconsumption. Fat-soluble vitamins, on the other hand, can accumulate in the body's fatty deposits. Overconsumption of a fat-soluble vitamin can be dangerous. Examine the structure of each vitamin shown here and classify it as either fat soluble or water soluble.

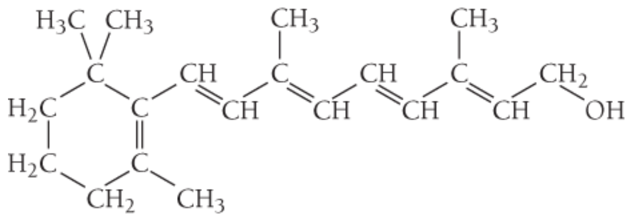
a. Vitamin C



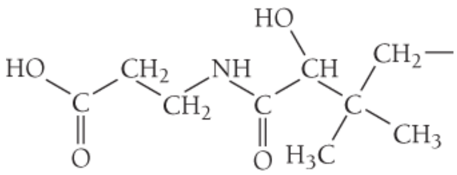
b. Vitamin K_3



c. Vitamin A

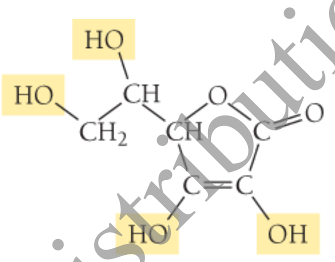


d. Vitamin B_5



SOLUTION

a. The four —OH bonds in vitamin C make it highly polar and allow it to hydrogen-bond with water. Vitamin C is water soluble.



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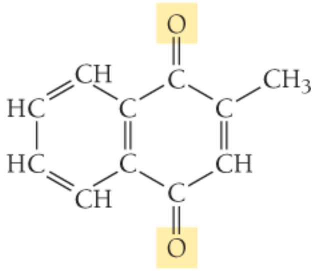
b. The C—C bonds in vitamin K_3 are nonpolar, and the C—H bonds are nearly so. The C═O bonds are polar, but the bond dipoles oppose and largely cancel each other, so the molecule is dominated by the nonpolar bonds. Vitamin K_3 is fat soluble.



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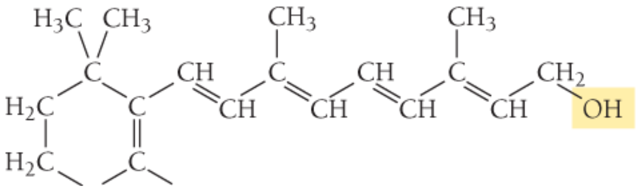
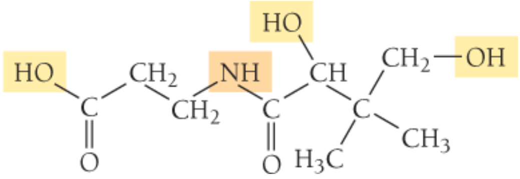
c. The C—C bonds in vitamin A are nonpolar, and the C—H bonds are nearly so. The one polar —OH bond may increase vitamin A's water solubility slightly, but overall it is nonpolar and therefore fat soluble.



d. The three —OH bonds and one —NH bond in vitamin B_5 make it highly polar and allow it to hydrogen-bond with water. Vitamin B_5 is water soluble.



FOR PRACTICE 13.1 Determine whether each compound is soluble in hexane.

a. water (H_2O)
b. propane ($CH_3CH_2CH_3$)
c. ammonia (NH_3)
d. hydrogen chloride (HCl)

Conceptual Connection 13.1 **Solubility**

Interactive