# CHAPTER 7

## UNDERSTANDING SOLIDS, LIQUIDS, AND GASES

CHEMICAL CONCEPTS IN THIS CHAPTER:

Three states of matter: solids, liquids, and gases
Kinetic molecular theory
Types of intermolecular forces
Heat energy and changes of state
Solutions
Solution concentrations

CONTEMPORARY ISSUES IN THIS CHAPTER:

Oil spill cleanup makes use of solubility considerations, which depend upon intermolecular forces.

# CHAPTER 7. UNDERSTANDING SOLIDS, LIQUIDS, AND GASES

## Three States of Matter: Solids, Liquids, and Gases

Everyday experiences make us familiar with the properties of solids and liquids.We observe that liquids flow and assume the shape of their containers, while "solid" means to us a substance that holds its shape. We notice that some liquids mix together, while others do not. We are familiar with a variety of solids: shiny metals that conduct heat and electricity well, diamonds that are famous for hardness, minerals with beautifully faceted crystal shapes, "dry ice" that disappears in a cloud of mist. Gases are a part of everyday life, too, though, unlike solids and liquids, they are often invisible to us. The air around us is a mixture of gases that is important to our well-being. Yet we "see" these gases only indirectly, as when they are used to inflate a balloon or a tire. The colors of some gas molecules make them visible to us; colored gas molecules like nitrogen oxides produce the brown color of Los Angeles smog. The chemical concepts we have learned about atoms and their properties can give us important insights about the properties of the three states of matter: solid, liquid, and gas.

*Fig. 7-1.* Three photos: a diamond crystal, liquid being poured, Los Angeles smog. Caption: *The three states of matter are solid, liquid, and gas.*

http://www.chem.purdue.edu/gchelp/liquids/character.html

## What Is a Gas? The Kinetic Molecular Theory

What is a gas? When a child's balloon is filled with helium gas, exactly what is inside the balloon? Actually, most of the space inside the balloon is empty. Helium atoms move through the space in the balloon, bouncing off the walls of the balloon and each other. The force of the collisions of the helium atoms against the sides of the balloon causes the balloon to stretch, or expand. The more helium is put into the balloon, the more the walls expand, because more helium atoms result in more collisions with the sides of the balloon.

*Fig. 7-2.* Two-part sketch. On the left, a smaller balloon; on the right, a larger balloon. The balloon on the right contains more molecules (shown as tiny dots) than the balloon on the left. Caption: *The more helium is put into the balloon, the more the walls expand, because more helium atoms result in more collisions with the sides of the balloon.*

The energy of motion of the atoms is called **kinetic energy**. Temperature is a measure of the kinetic energy of a gas. As a gas is heated, its kinetic energy, or energy of motion, increases, and the atoms or molecules of the gas move faster. Faster molecules produce more energetic collisions with the sides of the container. If the sides of the container are expandable, the volume increases as the sides of the container are pushed out by the higher-energy collisions. On a hot day, the helium balloons carried by children at a fair can burst as the atoms of helium gas inside them expand the balloons more and more with increasing temperature.

*Fig. 7-3*. Two-part sketch. On the left, a smaller balloon. On the right, a larger balloon. The gas molecules, represented by tiny dots, appear in exactly the same number on the left and the right. A sun is shining on the right-hand balloon. Caption: *Although the two balloons contain the same number of molecules, the balloon on the right has been warmed to a higher temperature, and the higher-energy molecules colliding with the walls of the balloon with greater force than those in the cooler balloon have caused it to expand.*

What happens if the walls of the container do not easily expand? The force of the molecules against the sides of the container results in **pressure** which increases inside the container when it is heated. To increase the pressure inside automobile tires, for example, more gas molecules are added from an air pump, resulting in more collisions against the tire walls and a higher pressure inside the tire.

*Fig. 7-4.* Photo of a gas-station air pump showing pressure meter. Caption: *As more air molecules are pumped into a tire, the pressure inside the tire increases.*

On a hot day, the pressure inside the tires can increase even if no more air is added, because the higher temperature results in higher kinetic energy of the gas molecules and higher-energy collisions with the walls of the tires. This explanation of the properties of gases as a logical outcome of the fact that they are made up of molecules moving in space is called the **kinetic molecular theory**.

**What Holds a Solid Together? Types of Intermolecular Forces**

**Table 7-1**
**Types of Solids and the Forces That Bind Their Units Together**

| Type of Solid | Examples | Type of Binding Force in the Solid | Special Properties of the Solid |
|---|---|---|---|
| Ionic | NaCl, $CaCl_2$, $NaHCO_3$ | Ionic bonding (very strong attraction between + and - ions) | Very high melting point, brittleness |
| Molecular<br>(a)Dispersion forces (weakest)<br><br>(b)Dipole-dipole forces<br><br>(c)Hydrogen bonding (strongest) | <br>(a)$Cl_2$, $H_2$<br><br>(b)$CO_2$,$H_2S$<br><br>(c)$H_2O$, HF, $NH_3$ | <br>(a)Very weak attractive forces between momentary dipoles<br>(a)Attractive forces between dipoles<br><br>(c) strong attractive forces between H (bonded to O,N, or F) and O,N,or F on another molecule | Relatively weak attractive forces result in low melting points, boiling points; hydrogen bonding is a special case with stronger attractive forces when H is bonded to O, N, or F |
| Covalent | Diamond (C), $SiO_2$ | Covalent bonds extend throughout the solid | Very strong, hard, high-melting solids |
| Metallic | Na, Ca, Cu, Fe | Metallic bonds result from sharing of valence electrons throughout the solid | Normally strong, hard high-melting solids, with good conductivity properties |

Unlike gases, which are made of atoms or molecules moving independently through space, solids are made of ions, atoms or molecules closely packed or joined together. We have already looked at the structure of one solid, sodium chloride, in Chapter 5. Sodium chloride, the substance of ordinary table salt, we learned, is made of positively charged sodium ions and negatively charged chloride ions, alternating in a neatly packed array. The highly ordered crystal structure is held together by the strong attraction of a positive and a negative charge for one another. Only **ionic solids**, those composed of ions, are held together by this strong force, called **ionic bonding**. The simple principle that unlike charges attract one another, however, is central to understanding the forces that hold together most kinds of solids.

What kind of force holds together the substances that are made not of ions, but of molecules? These are called **molecular solids**. As an example, let us examine the solid formed by molecules of carbon dioxide. In Chapter 5 we determined with an electron dot formula that the carbon dioxide molecule was linear in shape (Fig. 7-5a). Its bonds are polar covalent, because the oxygen atoms are more electronegative (have a stronger electron-pulling power) than the carbon atom. As a result, the oxygen atoms, which have a stronger attraction for the negatively charged electrons, have a partial negative charge, indicated by the symbol $\delta$-. The carbon atom is left with a partial positive charge, indicated by the symbol $\delta$+. Because of the symmetrical shape of the molecule, it does not have a positively charged end and a negatively charged end, so it is not a polar molecule.

*Fig. 7-5a.* Sketch of carbon dioxide molecule, showing partial positive charge in the center next to carbon and partial negative charges on either end next to the two oxygens. Caption: *The carbon dioxide molecule has polar covalent bonds. The more electronegative oxygen atoms bear a partial negative charge, and the carbon atom has a partial positive charge.*

The partial positive charge on the carbon atom of the carbon dioxide molecule and the partial negative charge on each of its oxygen atoms are very slight compared to the full charges on ionic solids like sodium chloride. Nevertheless, the carbon dioxide molecules can align themselves so that unlike partial charges are adjacent. The carbon dioxide solid is held together by the attractive forces between unlike charges.

*Fig. 7-5b.* Sketch of three oval-shaped carbon dioxide molecules showing partial positive and partial negative charges. The molecules are aligned so that unlike charges are adjacent. Caption: *Attraction between unlike partial charges causes weak intermolecular attractive forces in carbon dioxide.*

Forces like these which attract molecules to each other are called **intermolecular attractive forces**. The weakness of these attractive forces in carbon dioxide is shown by the fact that carbon dioxide is not a solid, but a gas at room temperature. Only at the very low temperature of -78.5°C do the molecules have a low enough kinetic energy to exist in the solid state. Solid carbon dioxide is called "dry ice" and is often used as a portable cooling agent. It has the unusual property of **subliming**, or going directly from the solid state to the gas state.

*Fig. 7-5c*. Photo of dry ice. Caption: *Solid carbon dioxide, or "dry ice," is sometimes used to keep food frozen in portable containers. Since carbon dioxide freezes only at temperatures below -78.5°C, it is so cold it can damage the skin, and must be handled with gloves.*

The weak forces that hold together the solid form of carbon dioxide are typical of molecular solids. Molecules like chlorine with nonpolar covalent bonding show even weaker bonding between molecules (Fig. 7-6a).

*Fig. 7-6a*. Sketch of the chlorine molecule. Caption: *The chlorine molecule has nonpolar covalent bonds.*

Since the two chlorine molecules have the same electronegativity, there is no permanent partial charge on any part of the atom. As the electrons travel in their orbitals around the atom, however, a momentary polarity of charge can form as the electron density is momentarily greater on one side of the molecule, producing a partial negative charge on that side of the molecule and a partial positive charge on the other side. This momentary charge can help to induce polarities in other molecules, as the partial positive charge attracts electrons on another nearby molecules and the partial negative charge repels electrons on nearby molecules. These partial charges last only for an instant, and the end of the molecule with a partial negative charge at one moment may become a partially positive end the next. At no time are the partial charges large in magnitude. Slight intermolecular attractive forces called **dispersion forces** result from these momentary partial charges (Fig. 7-6b). Though these attractive forces are slight, they are the only intermolecular forces for molecules with nonpolar bonds like $Cl_2$ and $H_2$ and inert gas atoms like He and Ne. Not surprisingly, solids held together only by dispersion forces are so loosely held that they are not seen under normal conditions. Chlorine is in the solid form only at temperatures below 100.98°C. Neon exists as a solid only below -248.67°C, and $H_2$ below 259.14°C!

Fig. 7-6b. Two-part diagram (I and II) of dispersion forces showing otherwise unlabeled oval molecules with partial charges which are reversed for a given molecule on the two sides of the figure. Caption: *Dispersion forces are weak partial charges which are momentary in nature. The charge may vary from one moment (I) to the next (II).*

One type of intermolecular force is unusually strong. Molecules in which hydrogen is bonded to oxygen, nitrogen, or fluorine exhibit a special kind of bonding called **hydrogen bonding**. The reasons for hydrogen bonding are based on the same principles we have discussed for other molecules. As an example, let us look at water, that important compound which covers two-thirds of the earth's surface and makes up two-thirds of our body weight. We have already found the structure of the water molecule using an electron-dot formula in Chapter 5. Its shape is bent, and its bonds are polar covalent (Fig. 7-7).

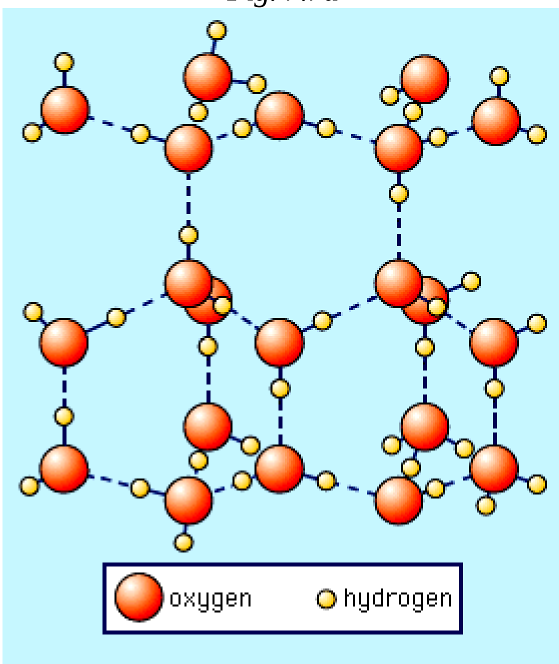
Fig. 7-7. Sketch of space-filling models of several water molecule showing a partial negative charge on the oxygen atom and partial positive charges on the hydrogen atoms. Hydrogen bonds (labeled) are indicated by dotted lines between molecules. Caption: *Hydrogen bonding is responsible for the strong intramolecular bonds between water molecules.*

Oxygen is more electronegative than hydrogen, and so the covalent bond between hydrogen and oxygen is polar, with the oxygen atoms having a greater share of electron density and a partial negative charge. Hydrogen, with part of its electron cloud pulled away by the oxygen atoms, is left with a partial positive charge. Hydrogen, as we learned in Chapter 5, has only one occupied energy level of electrons. When this minimal layer of electron density is diminished by the pull of a highly electronegative atom (like oxygen, nitrogen, or fluorine), the positively charged hydrogen nucleus is left relatively bare. The partial charges on the water molecule, then, are significant ones, greater than the partial charges usually associated with polar bonds in molecules. The resulting intermolecular forces, called hydrogen bonds, are especially strong ones.

What are the consequences of the strength of hydrogen bonds in our everyday lives? For one thing, this would be a very different world if water were a gas and not a liquid at normal temperatures. Yet that is precisely what would happen without hydrogen bonding. Hydrogen sulfide, $H_2S$, has the same electron dot structure as water and the same shape. Because it does not exhibit hydrogen bonding, hydrogen sulfide is a gas at room temperature, known best as the foul-smelling gas that gives rotten eggs their smell.

Because of its hydrogen bonding, solid water, also known, of course, as ice, has an unusual structure in which each water molecule is rigidly attached to three other water molecules. Each hydrogen, with a partial positive charge, is attracted by hydrogen bonding forces to a partially negative oxygen on another water molecule. Each oxygen is attracted by hydrogen bonding forces to a hydrogen on a neighboring water molecule. The bent shape of the water molecule dictates the directions that the hydrogen bonds between molecules must take. The resulting three dimensional structure, represented in Fig. 7-7a, does not fill space efficiently like the sodium chloride lattice structure. Instead, it has the appearance of interlocking hexagons.

*Fig. 7.7a*



The liquid state is less rigidly ordered, with molecules tumbling over one another, though they are still held together by mutual attractive hydrogen bonding forces. Lacking the hexagonal "holes," the liquid state is more dense than the solid state. Thus the solid state of water, being less dense, floats on top of the liquid, a most unusual property among chemical compounds.

Fig. 7-8. Two photos side by side, an iceberg floating in the sea and ice cubes floating in a glass. Caption: *Because solid water is less dense than the liquid state, ice floats on water.*

Because ice floats on water, northern lakes thaw much more easily in the spring than would be the case if ice were far from the sun's heat on the lake bottom, with enormous consequences for the ecosystem. These are but a few of the consequences of the uncommon properties of that most common chemical compound, water, that are a result of hydrogen bonding forces.

Some solids are held together by covalent bonds that extend throughout the solid, as if the entire solid were one giant molecule. Such solids are called **covalently bonded solids, or covalent solids**. Because covalent bonds, which involve shared electrons between atoms, are very strong forces, covalently bonded solids are very tightly held together. Diamond, made of carbon atoms joined together by covalent single bonds, is an example of a covalently bonded solid. Its strength and hardness are legendary, for the entire solid is joined by strong covalent bonds. These bonds are oriented in space in the symmetrical tetrahedral configuration we have seen for other single-bonded carbon atoms in Chapter 5 (Fig. 7-9). Another example of a covalently bonded solid is quartz, in which silicon and oxygen atoms are covalently bonded in a three-dimensional structure with two oxygen atoms for every silicon atom.

Fig. 7-9. Four-part-figure. (a) Sketch of the diamond lattice. Caption: *The solid we call diamond is made of carbon atoms covalently bonded together. Four bonds are formed from each carbon atom in a tetrahedral shape.* (b) Photo of a diamond. Caption: *The covalently bonded diamond solid is the hardest substance known.* (c) Sketch of the silicon dioxide lattice. Caption: *Silicon dioxide, or quartz, is a covalently bonded solid in which silicon and oxygen atoms are covalently bonded in a three-dimensional structure.* (d) Photo(s) of attractive quartz crystals. Caption: *Quartz, or silicon dioxide, is one of the most common minerals in the earth's crust, comprising 12% of its volume. Granite is a common form of silicon dioxide, but it is also found in pure crystalline form. Slight impurities may result in colored stones, like amethyst or tiger's eye.*

Metals are another solid form with a distinctive structure and bonding type. These are shared by all metallic elements, including alkali metals (Group 1 metals), alkaline earth metals (Group 2 metals), transition metals, lanthanides, and actinides. The atoms of these solids, called **metallic solids**, are packed in orderly, space-filling rows. The atoms of metallic elements have one or more valence electrons in an outer electron level that are more loosely held by the central, positively charged nucleus than the other electrons in the atom. These valence electrons are shared by all the metal atoms in the solid, forming a type of bonding called **metallic bonding**. One way to envision metallic bonding is to picture a "sea" of valence electrons surrounding the positively charged metal ions that result when valence electrons are removed from the metal atoms (Fig. 7-10). Since the valence electrons are not bound to a particular atom in the solid, it is possible to pull electrons from one end of a metallic solid while replacing them from the other end. The resulting flow of negatively charged electrons forms an electric current. Thus the nature of metallic bonding explains why metals are such good electrical conductors.

Fig. 7-10. (a) Sketch of two-dimensional array of metal ions (labeled only as +) surrounded by smaller spheres (labeled as $e^-$). Caption: *In metallic bonding, a "sea" of valence electrons is shared by all the metal atoms in the solid. (b) Photo of silver wire or object. Caption: Silver exhibits the shiny surface, malleable shape, and electrical conductivity that result from metallic bonding.*

**Heat Energy and Changes of State**

Now that we have some understanding of the gaseous state and the solid state, we are ready to consider the changes that matter undergoes as it changes from one state to another by absorbing energy in the form of heat. In changing from the closely bonded, rigid solid state to the free-moving molecules or atoms of the gaseous state, most substances pass through an intermediate state, the liquid state. We are familiar with the physical properties of liquids, for example, the fact that unlike solids, which tend to be rigid, liquids are pourable and take the shape of their containers. Liquids are less rigid than solids because they contain enough heat energy to overcome to some extent the bonding forces that hold the substance together. The weaker these intermolecular forces are, the easier it is to overcome them and change from the solid to the liquid form. Yet more heat energy is required to overcome entirely the intermolecular forces, producing the gas state, in which the atoms or molecules which form the substance move independently in space, free of intermolecular attraction.

Fig. 7-11. Three-part figure. (a) Sketch of the ice lattice (already featured in the chapter, here smaller in size). Caption: *The solid form of water is highly ordered in structure.* (b) Sketch of space-filling models of water in the liquid state. Caption: *The liquid state of water is less ordered than the solid state.* (c) Sketch of space-filling models of water separated by space as in the gas phase. Caption: *In the gas state, water molecules move in space, unbound by intermolecular attractive forces.*

*Even better, check out this link!*
http://www.chem.purdue.edu/gchelp/liquids/character.html

A few substances change directly from the solid state to the gas state without passing through a liquid state. We have already encountered an example of this process, called **sublimation**, in carbon dioxide. Solid carbon dioxide, or dry ice, sublimes, forming carbon dioxide gas. Another example of sublimation is iodine, whose dark purple crystals give off purple iodine gas.

Fig. 7-12. Photo of iodine in a glass container, showing iodine vapor. Caption: *The sublimation of solid iodine forms an easily visible iodine gas.*

Changes of state are physical changes, not chemical changes. When carbon dioxide changes from the solid state to the gas state, for example, carbon dioxide molecules are present both before and after the change. They have passed from one physical state to another by absorbing enough heat energy to overcome intermolecular attractions.

To examine more closely what happens when a substance undergoes changes of state, let us examine what happens when water, perhaps the chemical substance of greatest importance in our lives, undergoes these changes. If we begin, for example, to heat a block of ice that is at a temperature of -10°C, the temperature of the ice will begin to rise (Fig. 7-12). The water will remain in the solid state until it reaches its melting point, 0°C. What, then, is happening to the heat energy that is absorbed below the melting point? This energy appears as vibrations of the water molecules in their crystal structure, which become more energetic as the temperature rises. When the temperature rises to the melting point, these vibrations become so energetic that gradually the crystal is shaken apart as the intermolecular attractive forces, which in this case are hydrogen bonding forces, are overcome. The temperature of the water remains at 0°C until all the solid has melted. Then, if the heating process is continued, the temperature of the liquid water begins to rise.

Fig. 7-12. Graph of temperature vs. heat added, for water. Caption: *This graph shows the effect of adding heat to water, beginning at a temperature of -10°C. The temperature stays constant during the phase changes of solid to liquid and liquid to gas.*

**Problem example 7-1**: You are given a glass which contains nothing but ice cubes and pure water. Can you determine the temperature of the contents of the glass without a thermometer?

> First, make sure that the ice and water have been in contact long enough to make sure that they are at the same temperature. Then, without measuring the temperature, you can state with confidence that the contents of the glass are at 0°C, the freezing point of water. The temperature will remain at 0°C until all the ice has melted, for the heat energy is used to overcome the attractive forces holding together the solid.

As liquid water is heated, the temperature rises slowly, as the molecules of water gain energy and begin to move more quickly. We can view this process indirectly if we put a few drops of dye into a glass beaker of water that is being heated. As the dye diffuses through the liquid, it is pushed about by the water molecules more rapidly in the area that is being heated.

Fig. 7-13. Photo of dye being diffused in solution. Caption: Watching dye mix with water helps us to "see" the action of water molecules.

As we learned in Chapter 1, one calorie of heat is required to raise the temperature of one gram of water by one degree Celsius. The specific heat of water, 1 cal/g-°C, is a high one compared to many other substances. We have already observed in Chapter 1 that an iron ring supporting a beaker of water that is being heated becomes very hot long before the water is warmed, because the specific heat of the water is much higher than that of the iron. Now our understanding of intermolecular forces gives us an explanation for the high specific heat of water. The hydrogen bonding forces that attract the water molecules to one another are very strong ones, and these forces must be overcome as the water molecules move away from one another.

The temperature of the water will continue to rise as heat is added until the boiling point of water, 100°C, is reached. At this temperature the molecules gain enough kinetic energy to break way from the attractive forces of hydrogen bonding that hold them together. The molecules go into the gas state; a gas like this that has escaped from a liquid is often called a **vapor**, and the change from a liquid state to a gas, or vapor, state is called **evaporation**. Some evaporation occurs at all temperatures, as a few molecules gain enough kinetic energy to escape the liquid. The higher the temperature of the liquid, the greater the proportion of evaporating molecules, until the boiling point is reached. The boiling temperature of water is 100°C at normal atmospheric pressure. Higher pressures have the effect of making it more difficult for gas molecules to escape into the vapor phase and raising the boiling point; lower pressures result in a lower boiling point. The temperature of boiling water will remain at the boiling point until all the water has evaporated. Then the gas temperature will rise if it is heated further. Fig. 7-12 shows in graphical form the changes of temperature and state that our water has undergone in passing from the solid state at -10°C to the gas state.

The changes of state that have taken place in our small sample of water happen around us every day. Ice absorbs heat from its surroundings on a warm winter day, remaining at 0°C until it is all melted. Water in an ocean or lake in the summer remains cooler than the surrounding land because of the high specific heat of water, creating a refreshing relief from the heat for the swimmer or boater. The water in our sweat evaporates, absorbing heat energy from our bodies to overcome intermolecular attractions and enter the gas state. Chemical concepts we have learned enable to understand these phenomena, and many more as well.

**Solutions**

If we take a liter of water, add one mole of sodium chloride (58.5 g), and stir, the solid soon disappears, leaving a clear liquid that seems just like the water we started with. Tasting that water, though, would prove to us that salt had been added. By dissolving the sodium chloride in the water, we have created a **solution** of sodium chloride in water. Solutions, particularly those made with water, which are often called **aqueous solutions**, are important to us. Blood, urine, and sweat are all aqueous solutions. The ocean, the rivers, rainwater, and our drinking water are aqueous solutions. Most water we encounter contains some dissolved substances. Totally pure water is very rare, prepared only with difficulty in the laboratory. How do substances dissolve one another? What are the properties of the resulting solutions? The chemical concepts we have already learned can help us to answer these questions.

When two substances mix together to form a solution, the substance in greater quantity is called the **solvent**; the substance in lesser quantity, the **solute**. We say that the solvent dissolves the solute. Removing a spot from clothing, for instance, may be a matter of finding the right solvent to dissolve the spot. The solvent is usually a liquid. The solute may be a solid, like the sodium chloride, or table salt, in our earlier example. The solute may, however, be a liquid, like alcohol mixed with water as in a cocktail, or a gas, like the carbon dioxide which gives fizz to soda pop.

Fig. 7-13. (3-part figure) Photos of three types of solutions. (a) A cup of tea. Caption: *A cup of tea gets its color from solids dissolved in aqueous solution.* (b) A bottle of vodka. Caption: *Vodka is an aqueous solution of ethyl alcohol.* (c) A bottle of seltzer water. Caption: *Seltzer water is an aqueous solution of carbon dioxide gas.*

The rule that is often quoted to determine whether two substances will mix to form a solution is "like dissolves like." This rule refers to the fact that polar substances dissolve polar substances, while nonpolar substances dissolve nonpolar substances. On examination, this rule is simple a restatement of a fact that we have already learned: opposite charges attract, and the greater the magnitude of the charge, the greater the attraction. For example, why doesn't gasoline mix with water? Gasoline is a mixture of hydrocarbons, primarily octane, $C_8H_{18}$. Since the electronegativities of carbon and hydrogen are very similar, the covalent bonds in a hydrocarbon are nonpolar, and octane is a nonpolar molecule. The polar water molecules are attracted to one another by the strong intermolecular attraction called hydrogen bonding. There is no such strong force to attract them to the octane molecules of gasoline. Hence, when added to gasoline, the water molecules stay together, held by strong intermolecular attractive forces. The gasoline molecules, lacking the partial positive and negative charges that attract polar molecules to one another, remain separate, left out of the electrostatic attraction game. Being less dense than water, the gasoline floats on top of the water. Oil spill recovery teams must make use of these solubility and density properties of oil and water in cleaning up oil spills.

Fig. 7-14. Photo of oil spill (Caption to identify the source of the oil spill). Caption: *Nonpolar oil molecules do not mix with water molecules.*

Though hydrocarbons and other nonpolar substances do not mix with water, they mix easily with one another, since all nonpolar substances have similar intermolecular forces. Or, to apply the simplified rule, "like dissolves like." These forces result from the slight polarities that arise from small electronegativity differences in their atoms and from dispersion forces.

If hydrocarbons and other nonpolar substances do not mix with water, what types of substances form aqueous solutions? Since water itself features hydrogen bonding as an intermolecular force, the rule "like dissolves like" suggests that other molecules which form hydrogen bonds should mix with water. Alcohols like ethanol, $C_2H_5OH$, although they contain hydrogen and carbon, also feature a hydrogen atom attached to an oxygen atom and thus fulfill the requirements for hydrogen bonding. The end of the molecule with the -OH group attaches firmly to a water molecule by hydrogen bonding, and so ethanol is soluble in water. Ethanol is the alcohol found in the aqueous solutions of beer and wine. Table sugar, or sucrose, with the formula $C_{12}H_{22}O_{11}$, is another example of a substance with -OH groups that dissolves in water because of its ability to form hydrogen bonds.

In addition to those molecules which can form hydrogen bonds, water readily dissolves other polar molecules. The HCl molecule, for example, dissolves readily in water, as does ammonia, $NH_3$.

How well do ionic compounds dissolve in water? It is difficult for us to make predictions with certainty from our theories of intermolecular attractions in the case of ionic substances. On the one hand, water is a highly polar substance, so it should attract the charged ions of ionic substances. On the other hand, can we expect ions to be attracted to the partial charges on a water molecule in preference to the full charges on other ions? Conducting experiments to find out the solubilities of ionic substances by attempting to mix them with water is easy to do, and the results have been tabulated in chemistry handbooks. Some ionic compounds are found to be water-soluble; others have extremely low solubilities in water. Several factors are probably involved in determining the solubilities of ions in water, like the number of water molecules that can cluster in the space around an ion and the magnitude of the charge on an ion. The chemistry of dissolved ions in solution is an important topic. Both sea water and the fluid in our body cells, for example, contain dissolved ions. The remarkable similarity of the ion content of these two fluids has been cited in support of the theory that the first cells of living creatures formed by encapsulating the sea water in which life began.

**Solution Concentrations**

It is important to know not only what substances are present in a solution, but also how much. For example, a 5% solution of acetic acid is normal household vinegar, useful for preparing a tasty salad dressing. A 50% solution of acetic acid is hazardous, and can cause serious skin burns.

The question "How much?" with respect to solutions is best answered in terms of concentration units. To understand the principle of solution concentration, let us return to our early example of salt dissolved in water. First, consider a solution in which one mole, or 58.5 g, of sodium chloride is mixed with water to make one liter of solution. Compare that solution to one in which 58.5 g of sodium chloride is mixed with water to make 2 liters of solution (Fig. 7-15).

Fig. 7-15. Two-part sketch. (a) One-liter beaker next to dish of crystals labeled "58.5 g NaCl." Caption: *58.5 g of NaCl (1 mole NaCl) dissolved in 1 liter makes a solution of 1 mole NaCl per liter.* (b) Two-liter beaker next to dish of crystal identical to the one pictured in (a). Caption: *58.5 g of NaCl (1 mole NaCl) dissolved in 2 liters makes a solution of 0.5 mole NaCl per liter.*

Both solutions contain the same amount of NaCl. The second solution, with twice as much water, is half as concentrated as the first. Clearly, to know the concentration of a solution we need to know both the amount of solute that has been dissolved and the volume of solvent.

Since substances react chemically according to mole ratios, it is useful to measure concentrations using moles. The most common concentration used in the chemical laboratory is the number of moles per liter, or the **molarity**, abbreviated M. To find the molarity of a solution, we divide the number of moles in a solution by the number of liters. In equation form,

$$M = \frac{\text{moles}}{\text{liter}} \qquad (7\text{-}1)$$

**Problem example 7-2**: Find the molarity of a solution in which 58.5 g of NaCl has been dissolved to make 2.00 liters of solution.

First, we define our problem objective, the molarity, in terms of Eqn. 7-1.

$$M = \frac{\text{moles}}{\text{liter}} \qquad (7\text{-}1)$$

This equation shows that we need to know both the number of moles and the number of liters, then divide moles by liters to find molarity. The simplest way to solve this kind of problem is to do each of these steps in turn:

(1) Find the number of moles.
(2) Find the number of liters.
(3) Divide moles by liters to find moles per liter, or molarity.

(1)In this problem, first we analyze the problem to determine what we are given and what we want to know to solve the first step of the problem, finding moles NaCl:

| *WE ARE GIVEN* | *WE WANT* |
|---|---|
| 58.5 g NaCl | moles NaCl |

To go from what we are given to what we want, we need to know the number of grams in 1 mole of NaCl, found as in Chapter 6 by adding the gram atomic weights from the periodic table for each element in the compound:

| | |
|---|---|
| Na | 23.0 g |
| Cl | 35.5 g |
| | 58.5 g in 1 mole NaCl |

From this information we construct a conversion factor to give us the units we want, taking care to put the units we want on top of the conversion factor:

| *WE ARE GIVEN* | | *CONVERSION FACTOR* | | *WE WANT* |
|---|---|---|---|---|
| 58.5 g NaCl | X | $\frac{\text{1 mole NaCl}}{\text{58.5 g NaCl}}$ | = | moles NaCl |

Solving, we find that we have 1.00 mole NaCl.
(2) The volume is already given in liters, the units we require to find molarity.
(3) We are ready to divide moles by liters to find molarity:

$$M = \frac{\text{moles}}{\text{liter}} = \frac{\text{1.00 mole}}{\text{2.00 liters}} = 0.500$$

**Problem example 7-3:** 5.85 g of NaCl is dissolved in water to make 500 mL of solution. What is the molarity of the resulting solution?

Analyzing the problem as in example 7-3,

| *WE ARE GIVEN* | *WE WANT* |
|---|---|
| 5.85 g NaCl | moles NaCl |

The conversion factor which will give us the moles NaCl is the number of grams in a mole, which we found in the previous example as 58.5 g/mole.

| *WE ARE GIVEN* | | *CONVERSION FACTOR* | | *WE WANT* |
|---|---|---|---|---|
| 5.85 g NaCl | X | $\frac{\text{1 mole NaCl}}{\text{58.5 g NaCl}}$ | = | moles NaCl |

Solving, we find that we have 0.100 moles NaCl.

(2) In this problem the volume is given in units of mL:

| *WE ARE GIVEN* | *WE WANT* |
|---|---|
| 500 mL | liters |

Using the fact that there are 1000 milliliters in 1 liter of solution, we construct a conversion factor that gives us the desired unit of liters:

| *WE ARE GIVEN* | | *CONVERSION FACTOR* | | *WE WANT* |
|---|---|---|---|---|
| 500 ml | x | $\frac{\text{1 liter}}{\text{1000 mL}}$ | = | liters |

Solving, we find that there is 0.500 liter.

(3) Using our values of moles and liters to find moles per liter,

$$M = \frac{\text{moles}}{\text{liter}} = \frac{\text{0.100 mole}}{\text{0.500 liters}} = 0.200$$

**CONCEPTS TO UNDERSTAND FROM CHAPTER 7**

The three states of matter are solid, liquid, and gas.

Changes of state are physical changes, not chemical changes.

The properties of a gas are explainable by the fact that a gas is made of molecules moving in space. The temperature of a gas is a measure of the kinetic energy, or energy of motion, of the molecules of a gas. This concept is called the kinetic molecular theory. You should be able to explain in your own words the following "laws", using the kinetic molecular theory, that is, the behavior of the moving molecules of gas, as the basis of your explanation:

> The volume of a gas in an expandable container increases when the temperature increases.
>
> The pressure of a gas in a rigid container increases when the temperature increases.

Solids of different types (ionic, molecular, covalent, metallic) are held together by different types of characteristic forces (as detailed in Table 7-1).

Molecules in which hydrogen is bonded to an oxygen, a nitrogen, or a fluorine exhibit a special kind of intermolecular force called hydrogen bonding.

The prevalence of water both as a major component of our bodies and of our planet make the properties of water and of aqueous solutions especially important to us.

Some substances, like carbon dioxide and iodine, change directly from the solid state to the gas state without passing through a liquid state. This process is called sublimation.
The solubilities of substances, or their tendencies to mix with one another, depend on their mutual attractive forces. When two substances mix, the one in greater quantity is called the solvent; the one in lesser quantity is called the solute.

**SKILLS TO LEARN FROM CHAPTER 7**

Give both the amount of solute present and the volume of a solution, be able to calculate the molarity of the solution (concentration in moles/liter) from the relationship

$$M = \frac{\text{moles}}{\text{liter}} \qquad (7\text{-}1)$$

**Name________________________________** **Date ______________**

## PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 7

7-1. Why do auto tires need more air as seasons change from summer to winter and temperatures fall? Explain using the kinetic molecular theory.

7-2. A balloon is plunged into ice water. Predict what will happen to the appearance of the balloon, and explain why, using the kinetic molecular theory.

7-3. Why does ice float on water? Explain, using your knowledge of attractive forces in solids. Hint: check the table of molecular solids and their properties.

7-4. Why is a diamond hard? Explain, using your knowledge of attractive forces in solids.

7-5. Predict the properties of granite. Give your reasoning, using your knowledge of attractive forces in solids. (Hint: check both the table and the chapter figure labels.)

7-6. Why does copper wiring conduct electricity? Explain, using your knowledge of attractive forces in solids.

7-7. Why is $H_2O$ a liquid at room temperature, while $H_2S$ is a gas at room temperature?

7-8. Which of the following compounds exhibit hydrogen bonding? (Circle those compounds.)

a. HCl
b. HF
c. $NH_3$
d. $PH_3$
e. $CH_3OH$
f. $CH_3NH_2$
g. $CH_3SH$

7-9. What kind of forces must be overcome in order to change each of the following substances from a liquid to a gas?

a. He
b. HF
c. $H_2S$
d. $NH_3$
e. $H_2O$
f. $Cl_2$

7-10. What kind of forces must be overcome in order to change each of the following substances from a solid to a liquid?

a. diamond
b. $Li_2O$
c. Fe
d. Na
e. NaF
f. $SiO_2$

7-11. Give an example of each of the following from your personal experience.

a. Evaporation

b. Sublimation

c. Solution

7-12. Why does sweating make you feel cooler? What intermolecular forces are being overcome in this process?

7-13. Why does water dissolve ethyl alcohol ($CH_3CH_2OH$)? Explain using your knowledge of intermolecular forces.

7-14. Why doesn't water mix with gasoline? Expalin using your knowledge of intermolecular forces.

7-15. Find the molarity of the solution that results when 100 grams of table sugar (sucrose, $C_{12}H_{22}O_{11}$) is dissolved to make one liter of solution.

7-16. Find the molarity of the solution that results when 100 grams of sucrose is dissolved to make 500 mL of solution.

7-17. Find the molarity of the solution that results when 100 grams of potassium nitrate is dissolved to make 250 mL of solution.

7-18. A pot of water is boiling on the top of a stove under normal atmospheric pressure. Without using a thermometer, can you tell the temperature of the water?

7-19. Tell whether the following describe chemical changes or physical changes.

a. Calcium carbonate when heated forms calcium oxide and carbon dioxide.
b. Solid ice when heated forms liquid water.
c. Liquid water when heated forms water vapor.
d. Hydrogen gas and oxygen gas react explosively to form liquid water.

7-20. Atmospheric pressure is lower in the Rocky Mountains than at sea level. Is the boiling point of water higher, lower, or the same in the Rockies?

7-21 Which is the correct answer?

When the ice in a glass of icewater melts, the level of water in the glass will:

a) get higher
b) get lower
c) remain the same