Paper No. 01

Paper Title: Food Chemistry

Module-02: Water in Food Systems

Water is abundant in all living things and consequently is in almost all foods, unless steps have been taken to remove it. It is essential for life, even though it contributes no calories to the diet. Water also greatly affects the texture of foods, as can be seen when comparing grapes and raisins (dried grapes), or fresh and wilted lettuce. It gives crisp texture or turgor to fruits and vegetables and also affects perception of the tenderness of meat. For some food products, such as potato chips, salt, or sugar, lack of water is an important aspect of their quality and keeping water *out* of such foods is important to maintain quality.

Types of water: Free, Bound, and Entrapped Water

Most natural foods contain water up to 70% of their weight or greater unless they are dehydrated, and fruits and vegetables contain water up to 95% or greater. Water that can be extracted easily from foods by squeezing or cutting or pressing is known as ***free water***, whereas water that cannot be extracted easily is termed as ***bound water***.

Bound water usually is defined in terms of the ways it is measured; different methods of measurement give different values for bound water in a particular food. Many food constituents can bind or hold onto water molecules, such that they cannot be removed easily and they do not behave like liquid water. Some characteristics of bound water include:

• It is not free to act as a solvent for salts and sugars.
• It can be frozen only at very low temperatures (below freezing point of water).
• It exhibits essentially no vapor pressure.
• Its density is greater than that of free water.

Bound water has more structural bonding than liquid or free water; thus, it is unable to act as a solvent. As the vapor pressure is negligible, the molecules cannot escape as vapor; the molecules in bound water are more closely packed than in the liquid state, so the density is greater. An example of bound water is the water present in cacti or pine tree needles—the water cannot be squeezed for pressed out; extreme desert heat or a winter freeze does not negatively affect bound water and the vegetation remains alive. Even upon dehydration, food contains bound water.

Water molecules bind to polar groups or ionic sites on molecules such as starches, pectins, and proteins. Water closest to these molecules is held most firmly; the subsequent water layers are held less firmly and are less ordered, until finally the structure of free water prevails. Water also may be **entrapped** in foods such as pectin gels, fruits, vegetables, and so on. Entrapped water is immobilized in capillaries or cells, but if released during cutting or damage, it flows freely. Entrapped water has properties of free water and no properties of bound water.

Almost all food processing techniques involve the use of water or modification of water in some form: freezing, drying, emulsification (trapping water in droplets), breadmaking, thickening of starch, and making pectin gels are a few examples. Further, because bacteria cannot grow without water, the water content has a significant effect on maintaining quality of the food. This explains why freezing, dehydration, or concentration of foods increases shelf life and inhibits bacterial growth.

Water is important as a solvent or dispersing medium, dissolving small molecules to form true solutions and dispersing larger molecules to form colloidal solutions. Acids and bases ionize in water; water also is necessary for many enzyme-catalyzed and chemical reactions to occur, including hydrolysis of compounds such as sugars. It also is important as a heating and cooling medium and as a cleansing agent.

Because water has so many functions that are important to a food scientist, it is important to be familiar with some of its unique properties. When modifying the water content of a food, it is necessary to understand these functions in order to predict the changes that are likely to occur during processing of such foods. Drinking water is available to the consumer in convenient bottled and aseptic containers in addition to the tap.

Chemistry of Water

The chemical formula for water is H2O. Water contains strong ***covalent bonds*** that hold the two hydrogen atoms and one oxygen atom together. The oxygen can be regarded to be at the center of a tetrahedron, with a bond angle of 105^{o} between the two hydrogen atoms in *liquid water* and a larger angle of $109^{o}\,6^{o}$ between the hydrogens in *ice* (Figure 2.1).

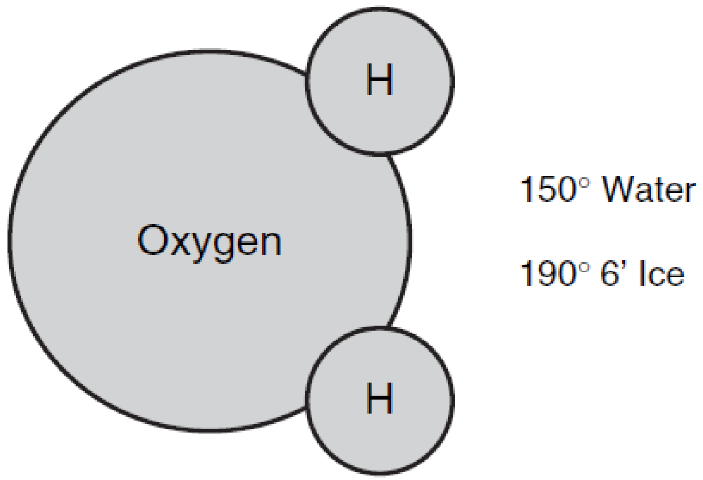

FIGURE 2.1 Bond angle of water and ice.

The bonds between oxygen and each hydrogen atom are polar bonds, having a 40% partial ionic character. This means that the outer-shell electrons are unequally shared between the oxygen and hydrogen atoms, the oxygen atom attracting them more strongly than each hydrogen atom. As a result, each hydrogen atom is slightly positively charged and each oxygen atom is slightly negatively charged. Therefore they are able to form ***hydrogen bonds***.

A hydrogen bond is a *weak* bond between polar compounds where a hydrogen atom of one molecule is attracted to an electronegative atom of another molecule (Figure 2.2). It is a weak bond relative to other types of chemical bonds such as covalent or ionic bonds, but it is very important because it usually occurs in large numbers, and therefore has a significant cumulative effect on the properties of the substance in which it is found. Water can form up to four hydrogen bonds (oxygen can form hydrogen-bond with two hydrogen atoms).

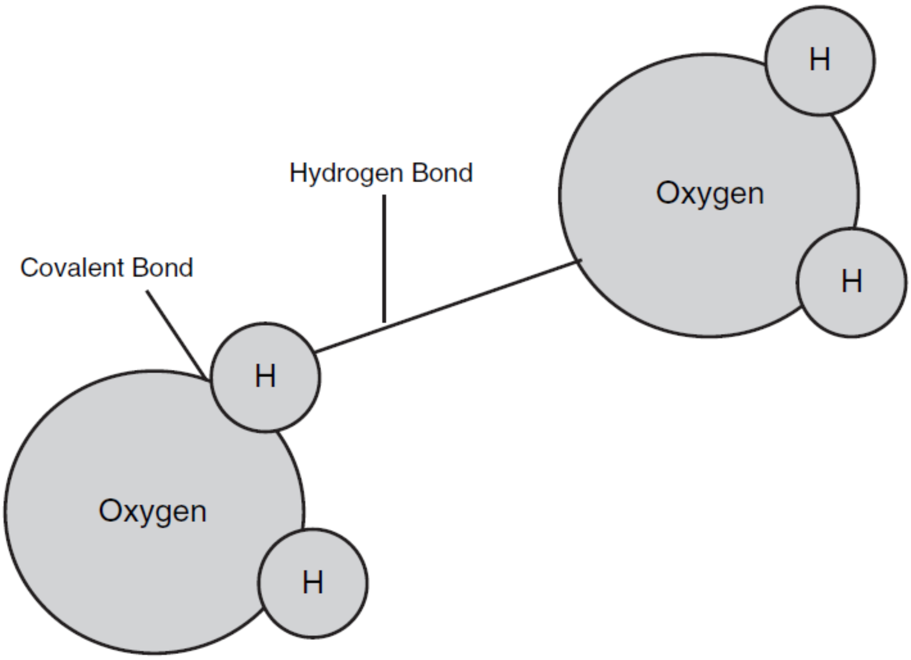

FIGURE 2.2 Hydrogen and covalent bonds in water molecules.

Water would be expected to be gas at room temperature if compared with similar compounds in terms of their positions in the periodic table, but because of the many hydrogen bonds it contains, it is liquid. Hydrogen bonds between hydrogen and oxygen are common, not just between water molecules but between many other types of molecules that are important in foods, such as sugars, starches, pectins, and proteins.

Due to its V-shape, each molecule of water can form up to four hydrogen bonds with its nearest neighbors. Each hydrogen atom can form one hydrogen bond, but the oxygen atom can form two, which results in a three-dimensional lattice in ice. The structure of ice is dynamic and hydrogen bonds are continually breaking and reforming between different water molecules. Liquid water also contains hydrogen bonds, and therefore has a variety of ordered structures that are continually changing as hydrogen bonds break and reform. In liquid water, it is estimated that about 80% of water molecules are involved in hydrogen bonding at any one time at 212^0F (100^0C), whereas 90% are involved in liquid water at 32^0F (0^0C).

Because *liquid* water has a smaller bond angle than ice, the molecules can be packed together more tightly, and so the *coordination number*, or in other words the average number of nearest neighbors is higher for water than for ice. The average distance between water molecules also is affected by temperature and increases with temperature as the molecules have more kinetic energy and can move around faster and farther at higher temperatures. Both of these affect the density of water, but the coordination number has a much more dramatic effect. Ice is less dense than water because the molecules have a smaller coordination number and cannot be packed together as tightly as water. Therefore, ice floats.

As water *freezes*, its density decreases and its *volume increases* by about 9%. This is very significant when freezing foods with high water content. Containers and equipment must be designed to accommodate the volume increase when the product freezes; for example, molds for popsicles must allow room for expansion. This volume increase also contributes to the damage to the structure of soft fruits on freezing. As water is heated above 39^0F (4^0C), the increase in the average distance between molecules causes a slight decrease in density.

Properties of water

Specific Heat and Latent Heat of Water

When ice is heated, the temperature increases in proportion to the amount of heat applied. The ***specific heat*** of water is the energy (in calories or in joules) required to raise the temperature of 1 g of water by 1^0C, and is the same whether heating water or ice. It is relatively high compared to other substances due to the hydrogen bonds. The specific heat of water is 1 cal/g per ^{0}C. This means that it takes 100 cal to raise the temperature of 1 g of water from 0 to 100^0C. Once ice has reached 0^0C, energy needs to be put in to break the hydrogen bonds and enable ice to change to the liquid form. Until the ice has been converted to liquid, there is no further change in temperature until steam is created.

The ***latent heat of fusion*** is the energy required to convert 1 g of *ice to water* at 0^0C and is 80 cal; that is, 1 g of ice at the freezing point absorbs approximately 80 cal as it changes to the liquid state
The ***latent heat of vaporization*** is the energy required to convert 1 g of *water into vapor* at 100^0C and is 540 cal; that is, 1 g of water at the boiling point absorbs approximately 540 cal as it becomes steam.

Both the *specific heat* and *latent heat* for water are fairly high compared with most substances, and this is an important consideration when water is used as a medium of heat transfer. It takes considerable energy to heat water and that energy is then available to be transferred to the food. Foods *heated* in water are slow to heat. Water also must take up considerable heat to evaporate. It takes heat from its surroundings, thus, it is a good *cooling* agent.

When ice is subjected to vacuum and then heated, it is converted into vapor *without* going through the liquid phase. This phenomenon is known as ***sublimation*** and is the basis for the food processing method known as ***freeze-drying***. Coffee is an example of a food product that is freeze-dried. The process is expensive and is used only for foods that can be sold at a high price, such as coffee. The coffee beans are frozen and then subjected to a high vacuum, after which radiant heat is applied until almost all of water is removed by sublimation. Freezer burn also is the result of sublimation.

Vapor Pressure

If a puddle of water is left on the ground for a day or two, it will dry up because the liquid evaporates. The water does not boil, but individual water molecules gain enough energy to escape from the liquid as vapor. Over a period, an *open*, small pool of water will dry up in this way. If the liquid is in a *closed* container, at equilibrium, some molecules are always evaporating and vapor molecules are condensing, so there is no overall change in the system. The *vapor* (gaseous) molecules that have escaped from the *liquid* state exert a pressure on the surface of the liquid known as the ***vapor pressure***.

When the vapor pressure is *high*, the liquid evaporates (is vaporized) easily and *many* molecules exist in the vapor state; the boiling point is low. Conversely, a *low* vapor pressure indicates that the liquid does not vaporize easily and that there are few molecules existing in the vapor state. The boiling point for these liquids is higher. The liquid boils when the vapor pressure reaches the external pressure.

The vapor pressure increases with increasing temperature. At higher temperatures, the molecules have more energy and it is easier for them to overcome the forces holding them within the liquid and to vaporize; so there are more molecules in the vapor state.

The vapor pressure decreases with addition of solutes, such as salt or sugars. In effect, the solutes dilute the water; therefore, there are fewer water molecules (in the same volume) available for vaporization, and thus, there will be fewer molecules in the vapor state and the vapor pressure will be lower. Attraction to the solute also limits evaporation.

Boiling Point

Anything that lowers the vapor pressure (pressure by gas above the liquid) increases the boiling point. This is due to the fact that as the vapor pressure is lowered at a particular temperature, more energy must be put in; in other words, the temperature must be raised to increase the vapor pressure again. The external pressure does not change if salts or sugars are added, but it is more difficult for the molecules to vaporize; so the temperature at which the vapor pressure is the same as the external pressure (boiling point) will be higher. One mole of sucrose elevates the boiling point by 0_52^{0}C and 1 mole of salt elevates the boiling point by 104^{0} C. Salt has double the effect of sucrose because it is ionized, and for every mole of salt, there is 1 mole of sodium ions and 1 mole of chloride ions.

Salts and sugars depress the freezing point of water in a similar fashion. If the external pressure is increased by heating in a pressure cooker or retort (commercial pressure cooker), the boiling point increases and a shorter time than normal is required to cook a particular food (the basis of preserving foods by canning). For example, food may be heated in cans in retorts, and the steam pressure is increased to give a boiling point in the range 239–250^{0}F_115–121^{0}C. Conversely, if the external pressure is decreased, for example, at high altitude, water boils at a lower temperature and so food may require a longer time to cook.

Solution

Water *dissolves* small molecules such as salts, sugars, or water-soluble vitamins to form a true ***solution***, which may be either **ionic** or **molecular**. An ionic solution is formed by dissolving substances that ionize in water, such as salts, acids, or bases. Taking sodium chloride as an example, the solid contains sodium (Na^+) and chloride (Cl^-) ions held together by ionic bonds. When placed in water, the water molecules reduce the attractive forces between the oppositely charged ions, the ionic bonds are broken, and the individual ions become surrounded by water molecules, or **hydrated**. Each ion is usually surrounded by six water molecules; the ions move independently of each other.

Polar molecules, such as sugars, which are associated by hydrogen bonding, dissolve to form **molecular solutions**. When a sugar crystal is dissolved, hydrogen-bond interchange takes place and the hydrogen bonds between the polar hydroxyl groups on the sugar molecules are broken and replaced by hydrogen bonds between water and the sugar molecules. Thus, the sugar crystal is gradually hydrated; each sugar molecule being surrounded by water molecules. Water molecules bind to polar groups on the sugar molecules by hydrogen bonds. The sugar molecules are removed from the sugar crystal and hydrated as water molecules surround them and bind to them by hydrogen bonds. When a hydrogen-bond interchange is involved, solubility increases with increasing temperature. Heating disrupts hydrogen bonds and reduces water– water and sucrose–sucrose attraction, thus facilitating formation of hydrogen bonds between water and sucrose and hydration of sucrose molecules. Therefore, sucrose is much more soluble in hot water than in cold water. Solutes increase the boiling point of water, and the dramatic increase in sucrose solubility with temperature, particularly at temperatures above 100_C (the boiling point of pure water), makes it possible to determine the sucrose concentration by measuring the boiling point of sucrose solution. This is important when making candies or pectin jellies.

Colloidal Dispersion

Molecules that are *too big* to form true solutions may be dispersed in water. Those with a particle size range *1–100 nm* are dispersed to form a ***colloidal dispersion*** or ***sol***. Examples of such molecules include cellulose, *cooked* starch, pectic substances, gums, and some food proteins. Colloidal dispersions are often unstable; thus, food scientists must take care to stabilize them where necessary if they occur in food products. They are particularly unstable to factors such as heating, freezing, or pH change. Changing the conditions in a stable dispersion can cause precipitation or gelation; this is desirable in some cases, for example, when making pectin jellies. Colloid science is important to food scientists as many convenient or packaged foods have colloidal dimensions and their stability and sensitivity to certain types of reactions can be understood only with knowledge of colloid science.

Suspension

Particles that are *larger than 100*nm are too large to form a colloidal dispersion. These form a ***suspension*** when mixed with water. The particles in a suspension separate out over a period, whereas no such separation is observed with colloidal dispersions. An example of a suspension would be *uncooked* starch grains in water. It may be temporarily suspended and then easily settle out, no longer "suspended," but rather falling to the bottom of the container/pan.

Conclusion

The special properties of water are not only essential to life but play a key role in the manufacture of various foods, whether we are considering the solid (ice), liquid (water) or gaseous (vapour) forms. In the formation of doughs and batters, water plays a major role as a plasticiser of polymers and as a solvent for other recipe components. The properties of the solutions formed during mixing strongly influence the rheological behaviour of doughs and batters during processing and contribute directly to the formation of the final product structure. Water plays a part in almost all of the chemical and physical changes that occur in the production of processed foods and the processes that control their qualities during storage.

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