How does ionic equilibria in aqueous solutions relates to the following foods?

Raspberries (*Rubus idaeus*)

Yogurt sauce

Cocoa pods

Reactions in aqueous solutions are very important in foods and food processing. As a major component and solvent in food, water is the main reaction media in food systems, in fact such reactions would not be possible without water. An important parameter in foods, called water activity or $a_w$, is a measure of the water available in a particular food to participate in reactions and it is an important control factor in the chemical and microbiological stability of food systems. The importance of water in the reactions that occur in foods, including acid-base reactions, is intimately related to its water ability to act as both a weak acid and a weak base, producing $\text{H}_3\text{O}^+$ and $\text{OH}^-$ by proton transfer, which makes it a highly efficient medium for proton exchange. In any aqueous solution at 25°C:

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = \frac{1.00 \times 10^{-14}}{1 \text{ mol dm}^{-3}}
\]

\[
[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}
\]
and concentrations of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ can vary from roughly $10^0$ to $10^{-14}$ mol dm$^{-3}$. This makes it convenient to define pH and pOH as:

$$\text{pH} = -\log\frac{[\text{H}_3\text{O}^+]}{\text{1 mol dm}^{-3}}$$
$$\text{pOH} = -\log\frac{[\text{OH}^-]}{\text{1 mol dm}^{-3}}$$

Figure \(\PageIndex{1}\): Ionization of water

Many reactions in foods involve weak acids or bases or slightly soluble substances, and in such cases one or more equilibria are achieved in solution. Furthermore, the equilibrium state is usually reached almost instantaneously, and so we can use the equilibrium law to calculate the concentrations and amounts of substance of different species in solution. Such information enables us to understand, predict, and control what will happen in solution, and it has numerous practical applications. Equilibrium constants may be used to obtain information about reactions in solution and be applied to practical problems allowing us to predict how some attributes of food including color, texture, and flavor development may be affected.

Since molecules of a strong acid transfer their protons to water molecules completely, $[\text{H}_3\text{O}^+]$ (and hence pH) can be obtained directly from the stoichiometric concentration of the solution. Similarly $[\text{OH}^-]$ and pOH may be obtained from the stoichiometric concentration of a strong base. In the case of weak acids and weak bases, proton-transfer reactions proceed to only a limited extent and a dynamic equilibrium is set up. In such cases an acid constant $K_a$ or a base constant $K_b$ as well as the stoichiometric concentration of weak acid or base are required to calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, or pOH. $K_a$ and $K_b$ for a conjugate acid-base pair are related, and their product is always $K_w$.

For multiple reasons ranging from maintaining proteins in solution to avoiding or favoring the development of certain microorganisms, it is often necessary to restrict the pH of food (aqueous solution) to a narrow range. This can be accomplished by means of a buffer solution—which contains a conjugate weak acid-weak base pair. If a small amount of strong base is added to a buffer, the $\text{OH}^-$ ions are consumed by the conjugate weak acid, so they have little influence on pH. Similarly, a small amount of strong acid can be consumed by the conjugate weak base in a buffer. To a good approximation the $[\text{H}_3\text{O}^+]$ in a buffer solution depends only on $K_a$ for the weak acid and the stoichiometric concentrations of the weak acid and weak base.

Indicators for acid-base titrations are conjugate acid-base pairs, each member of which is a different color. An indicator changes from the color of the conjugate acid to the color of the conjugate base as pH increases from approximately $pK_{In} - 1$ to $pK_{In} + 1$. Some colorants naturally found in food products (i.e., vegetables, fruits, est.) possess acid-base properties and can be used as indicators. However, their use is restricted by their availability in pure forms. For titrations involving only strong acids and strong bases, rarely found when talking about food systems, several indicators are usually capable of signaling the endpoint because there is a large jump in within ± 0.05 cm$^3$ of the exact stoichiometric volume of titrant. In the case of titrations which involve a weak acid or a weak base, a buffer solution is involved and the
jump in pH is smaller. This last type of titrations are widely used in food analysis in the presence of indicators for which the pH of color change is well known.

β-galactosidase

Anthocyanins

Orange Juice

A dynamic equilibrium is set up when a solid compound is in contact with a saturated solution. In the case of an ionic solid, the equilibrium constant for such a process is called the solubility product. $K_{sp}$ can be determined by measurement of the solubility of a compound, and it is useful in predicting whether the compound will precipitate when ionic solutions are mixed. The solubility product of certain salts used as food additives and nutritional supplements is a very important factor to consider when formulating food products. In most cases, we aim to maintain them in solution and avoid the formations of precipitates and crystals. The common-ion effect, in which an increase in the concentration of one ion decreases the concentration of the other ion of an insoluble compound, can be interpreted quantitatively using solubility products. It is also true that removal of one ion of an insoluble compound from solution will increase the concentration of the other ion, and hence the solubility. It is for this reason that salts of weak acids often dissolve in
acidic solutions—protonation of the anion effectively reduces its concentration to the point where the solubility product is not exceeded. Control of pH and the addition of sequestering agents are two strategies often used to maintain in solution ions that may combine into low solubility salts.

From ChemPRIME: 14.0: Prelude to Ionization of Water

Contributors and Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.