Water and Solutions

As usual, be sure to read this chapter for a more comprehensive discussion of water and solutions.

Water molecules have a very strong electric dipole moment which means they are very good at pulling ions out of crystals. It is ubiquitous on the earth and essential for life.

Water molecules at room temperature spontaneously break apart so that at any particular moment, one out of 556 million molecules has turned into an H^+ ion and an OH^- ion. In less than a millionth of a second later, the H^+ ion has joined with a nearby H_2O molecule forming an H_3O^+ hydronium ion. This tiny amount may seem unimportant, but the concentration of H_3O^+ and OH^- ions in water controls much of the chemistry that happens in water solutions, including that in living things. Often, people talk of the H^+ concentration when it is actually H_3O^+.

The concentration of H_3O^+ is so important that there is a quantity called pH which is a measure of the hydronium concentration in units of moles of H_3O^+ per liter of water. This way of expressing concentration is called the molarity and is given the symbol $[H_3O^+]$ with units of moles per liter (mol/L). Remember, the unit name is mole, but its symbol is mol, just like the unit name of length is meter, but its symbol is m. Concentrations of H_3O^+ ions can span an wide range from 10^{-1} mol/L for concentrated acids to 10^{-14} mol/L for concentrated bases. Tables 11.4 and 11.5 in your text show this for some common substances.

Your book does not explain how you calculate the pH form the $[H_3O^+]$, but I will expect you to be able to do this with your calculators. The formulas are:

$$\mathrm{pH} = -\log_{10}[\mathrm{H_3O^+}] \quad \text{and} \quad [\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$$

Note: pH is written without units, but when these formulas are used the concentration must be in mol/L.

First, try these formulas with pure water where $[\mathrm{H_3O^+}] = 1.0\times10^{-7}\,\mathrm{mol/L} = 10^{-7}\,\mathrm{mol/L}$:

$$\mathrm{pH} = -\log_{10}[\mathrm{H_3O^+}] = -\log_{10}(1.0\times10^{-7.0}\,\mathrm{mol/L}) = -(-7.0) = +7.0$$

and the reverse calculation

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}} = 10^{-(+7.0)} = 1.0\times10^{-7}\,\mathrm{mol/L}$$

Blood has a pH of 7.4 so

$$[\mathrm{H_3O^+}] = 10^{-(+7.4)} = 4.0\times10^{-8}\,\mathrm{mol/L} \quad \text{and} \quad \mathrm{pH} = -\log_{10}(4.0\times10^{-8.0}) = -(-7.4) = +7.4$$

The active component of vinegar, acetic acid CH_3COOH, loses a hydrogen when in water by the reaction

$$\mathrm{CH_3COOH + H_2O \Rightarrow CH_3COO^- + H_3O^+}$$

This makes it acidic with a pH of 4.76 so

$$[\mathrm{H_3O^+}] = 10^{-(+4.76)} = 1.74\times10^{-5}\,\mathrm{mol/L}$$

When sodium hydroxide NaOH is added to water, it separates into Na^+ and OH^- ions. These additional OH^- ions combine with hydronium to make water. As a result, the hydronium concentration is reduced and the pH is therefore increased. The pH of NaOH is 13.0 so the hydronium concentration has been reduced to

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}} = 10^{-(+13)} = 1.0\times10^{-13}\,\mathrm{mol/L}$$

Water is constantly undergoing the reaction $\mathrm{H_2O \Leftrightarrow H^+ + OH^-}$ which keeps the product $[H_3O^+]$ x $[OH^-]$ = 1.0 x 10^{-14} $(mol/L)^2$ even as $[H_3O^+]$ or $[OH^-]$ are individually changed.

When we built amino acids, I said you could put a H on the O^- and make NH_2 instead of NH_3^+. Once the amino acids are in the pH of a living body, however, the OH will lose its H and the NH_2 will gain an extra H. This is why pH is so important, chemicals change their character depending on the pH of the aqueous solution they are contained in.