# Chemistry 20

# Lesson 26 – pH and pOH

Refer to pages 238 to 244 in the Nelson text.

## I. Ionization of water

Pure water has a very slight conductivity that is observable only if measurements are made with very sensitive instruments. According to Arrhenius' theory, conductivity is due to the presence of ions. Therefore, the conductivity observed in pure water must be the result of ions produced by the ionization of some water molecules into hydronium (hydrogen) ions and hydroxide ions. (Note that in this context, it is understood that when we talk about hydrogen ions we are actually referring to hydronium ions. Sometimes it is easier to just use the hydrogen ion idea.) Because the conductivity is so slight, very few water molecules, less than $10^{-6}$ %, will actually be in hydrogen and hydroxide form.

$$H_2O_{(l)} \xrightarrow{\text{less than } 10^{-6}\,\%} H^+_{(aq)} + OH^-_{(aq)}$$

Evidence indicates that fewer than two water molecules in one billion ionize at SATP.

The equation for the ionization of water shows that hydrogen ions and hydroxide ions are formed in a 1:1 ratio. Therefore, the concentration of hydrogen ions and hydroxide ions in pure water and neutral solutions must be equal. The measured values at SATP are: $[H^+_{(aq)}] = 10^{-7}$ mol/L and $[OH^-_{(aq)}] = 10^{-7}$ mol/L.

$$H_2O_{(l)} + H_2O_{(l)} \longrightarrow \underset{1.0 \times 10^{-7}}{H_3O^+_{(aq)}} + \underset{1.0 \times 10^{-7}}{OH^-_{(aq)}}$$

The ionization of water is especially important in the empirical and theoretical study of acidic and basic solutions. According to Arrhenius theory, an acid is a substance that ionizes in water to produce hydrogen ions. The additional hydrogen ions provided by the acid increase the hydrogen ion concentration in the water; $[H^+_{(aq)}]$ will be greater than $10^{-7}$ mol/L, so the solution is acidic. A basic solution is one in which $[OH^-_{(aq)}]$ is greater than $10^{-7}$ mol/L. A basic solution is produced, for example, by the dissociation in water of an ionic hydroxide such as sodium hydroxide.

A number called the ionization constant for water ($K_w$) is equal to the product of the hydrogen and hydroxide concentrations.

$$K_w = [H^+_{(aq)}] \times [OH^-_{(aq)}]$$

Since $[H^+_{(aq)}] = 1.0 \times 10^{-7}$ mol/L and $[OH^-_{(aq)}] = 1.0 \times 10^{-7}$ mol/L

$$K_w = 1.0 \times 10^{-7} \text{ mol/L} \times 1.0 \times 10^{-7} \text{ mol/L}$$
$$K_w = 1.0 \times 10^{-14} \text{ (mol/L)}^2$$

One important observation is that the ionization constant, $K_w$, applies to all aqueous solutions. Therefore, $K_w$ may be used to calculate either $[H^+_{(aq)}]$ or $[OH^-_{(aq)}]$ if the other concentration is known.

*Example 1*

A 0.15 mol/L solution of hydrochloric acid at 25°C is found to have a hydrogen ion concentration of 0.15 mol/L. Calculate the concentration of the hydroxide ions.

$$HCl_{(aq)} \xrightarrow{100\%} H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$\left[H^{+}_{(aq)}\right] = \left[HCl_{(aq)}\right] = 0.15\,\text{mol/L}$$

$$\left[H^{+}_{(aq)}\right] \times \left[OH^{-}_{(aq)}\right] = K_w$$

$$\left[OH^{-}_{(aq)}\right] = \frac{K_w}{\left[H^{+}_{(aq)}\right]}$$

$$\left[OH^{-}_{(aq)}\right] = \frac{1.0 \times 10^{-14}\,(\text{mol/L})^2}{0.15\,\text{mol/L}}$$

$$\left[OH^{-}_{(aq)}\right] = \mathbf{6.7 \times 10^{-14}\ mol/L}$$

*Example 2*

Determine the hydrogen ion and hydroxide ion concentrations in 500 mL of an aqueous solution containing 2.6 g of dissolved sodium hydroxide.

$$n_{NaOH} = \frac{2.6\,\text{g}}{40.00\,\text{g/mol}} = 0.065\,\text{mol}$$

$$\left[NaOH_{(aq)}\right] = \frac{0.065\,\text{mol}}{0.500\,\text{L}} = 0.13\,\text{mol/L}$$

$$NaOH_{(s)} \longrightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

$$\left[OH^{-}_{(aq)}\right] = \left[NaOH_{(aq)}\right] = \mathbf{0.13\,mol/L}$$

$$\left[H^{+}_{(aq)}\right] \times \left[OH^{-}_{(aq)}\right] = K_w$$

$$\left[H^{+}_{(aq)}\right] = \frac{K_w}{\left[OH^{-}_{(aq)}\right]}$$

$$\left[H^{+}_{(aq)}\right] = \frac{1.0 \times 10^{-14}\,(\text{mol/L})^2}{0.13\,\text{mol/L}}$$

$$\left[H^{+}_{(aq)}\right] = \mathbf{7.7 \times 10^{-14}\ mol/L}$$

## II. Communicating Concentrations: pH and pOH

There is a tremendous range of hydrogen ion and hydroxide ion concentrations that can occur in different solutions, and there are two ways that they are communicated. The first way you are already familiar with: molar concentration (mol/L). Unfortunately, for many concentration values like $4.7 \times 10^{-11}$ mol/L, for example, the scientific notation becomes rather tiresome. A second system, called the **pH scale**, was developed by Danish chemist Sören Sörenson in 1909. pH stands for "power of hydrogen" which means "exponent of hydrogen." In exponent form the number $4.7 \times 10^{-11}$ is equivalent to $10^{-10.33}$. To calculate the –10.33 value, we calculate the **log** (short for logarithm) of the number. (Try it in your calculator.) Since most concentrations of hydrogen yield negative exponents, and it would be tiresome to always write the minus sign in front of the pH value, the pH value for a hydrogen ion concentration of $4.7 \times 10^{-11}$ mol/L is 10.33. Speaking formally: Expressed as a numerical value without units, the pH of a solution is **the negative of the logarithm to the base ten of the hydrogen ion concentration**.

On TI graphing calculators, log($4.7 \times 10^{-11}$) may be entered by pushing the following sequence of keys:

| log | 4 | • | 7 | 2nd | , |
|---|---|---|---|---|---|
| (–) | 1 | 1 | enter | | |

$$pH = -\log[H^{+}_{(aq)}]$$

However, when converting from pH to $[H^+_{(aq)}]$ the negative sign is reintroduced in the calculation.

$$[H^+_{(aq)}] = 10^{-pH}$$

Note that the higher the pH value the smaller the hydrogen ion concentration. Further, a difference of 1 pH translates into a 10 fold difference in hydrogen concentration. If, for example, one solution has a pH of 4.00 and second solution has a pH of 8.00, the second solution has a hydrogen ion concentration $10^4$ or 10000 times less than the first solution.

On TI graphing calculators, $10^{-10.33}$ may be entered by pushing the following sequence of keys:

$\boxed{1}$ $\boxed{0}$ $\boxed{\wedge}$ $\boxed{(-)}$ $\boxed{1}$ $\boxed{0}$ $\boxed{\bullet}$ $\boxed{3}$ $\boxed{3}$ $\boxed{\text{enter}}$

**or**

$\boxed{\text{2nd}}$ $\boxed{\log}$ $\boxed{(-)}$ $\boxed{1}$ $\boxed{0}$ $\boxed{\bullet}$ $\boxed{3}$ $\boxed{3}$ $\boxed{)}$ $\boxed{\text{enter}}$

Values of pH can be calculated from the hydrogen ion concentration, as shown in the following examples. The digits preceding the decimal point in a pH value are determined by the digits in the exponent of the given hydrogen ion concentration. These digits serve to locate the position of the decimal point in the concentration value and have no connection with the certainty of the value. However, **the number of digits following the decimal point in the pH value is equal to the number of significant digits in the hydrogen ion concentration**. For example, $[H^+_{(aq)}] = 2.7 \times 10^{-3}$ mol/L corresponds to a pH of 2.57 and $[H^+_{(aq)}] = 2.70 \times 10^{-3}$ mol/L corresponds to a pH of 2.569.

### *Example 3*

Communicate a hydrogen ion concentration of $4.72 \times 10^{-6}$ mol/L as a pH value.

$pH = -\log[H^+_{(aq)}]$

$= -\log(4.72 \times 10^{-6})$ (three significant digits)

$= \mathbf{5.326}$ (three digits following the decimal point)

### *Example 4*

Communicate a pH of 8.33 as a hydrogen ion concentration.

$[H^+_{(aq)}] = 10^{-pH}$

$= 10^{-8.33}$ (two digits following the decimal point)

$= \mathbf{4.7 \times 10^{-9}\ mol/L}$ (two significant digits)

Although pH is used in most applications, for some applications it may be convenient to describe hydroxide ion concentrations in a similar way. The definition of **pOH** follows the same format and the same certainty rule as for pH.

$$pOH = -\log[OH^-_{(aq)}]$$

$$[OH^-_{(aq)}] = 10^{-pOH}$$

The mathematics of logarithms (which you will learn about in Mathematics 30) allows us to express a simple relationship between pH and pOH.

Since $K_w = [H^+_{(aq)}]\,[OH^-_{(aq)}] = 1.0 \times 10^{-14}\ (mol/L)^2$

$$pH + pOH = 14.00 \qquad \text{(at SATP)}$$

This relationship enables a quick conversion between pH and pOH. If, for example a solution has a pH of 11.68, the pOH can be calculated: pOH = 14.00 – 11.68 = 2.32.

# III. Assignment

1. The hydrogen ion concentration in an industrial effluent is 4.40 mmol/L. Determine the concentration of hydroxide ions in the effluent.

2. The hydroxide ion concentration in a household cleaning solution is 0.299 mmol/L. Calculate the hydrogen ion concentration in the cleaning solution.

3. Calculate the hydroxide ion concentration in a solution prepared by dissolving 0.37 g of hydrogen chloride in 250 mL of water.

4. Calculate the hydrogen ion concentration in a saturated solution of calcium hydroxide (limewater) that has a solubility of 6.9 mmol/L.

5. What is the hydrogen ion concentration in a solution made by dissolving 20.0 g of potassium hydroxide in water to form 500 mL of solution?

6. Food scientists and dieticians measure the pH of foods when they devise recipes and special diets.
   (a) Complete the following **Acidity of Foods** table.

| **Food** | **$[H^+_{(aq)}]$ (mol/L)** | **$[OH^-_{(aq)}]$ (mol/L)** | **pH** | **pOH** |
|---|---|---|---|---|
| oranges | $5.5 \times 10^{-3}$ | | | |
| asparagus | | | | 5.6 |
| olives | | $2.0 \times 10^{-11}$ | | |
| blackberries | | | | 10.60 |

   (b) Based on pH only, predict which of the foods would taste most sour.

7. To clean a clogged drain, 26 g of sodium hydroxide is added to water to make 150 mL of solution. What are the pH and pOH values for the solution?

8. What mass of potassium hydroxide is contained in 500 mL of solution that has a pH of 11.50?

9. How does the hydrogen ion concentration compare with the hydroxide ion concentration if a solution is (a) neutral? (b) acidic? (c) basic?

10. At 25°C, the hydroxide ion concentration in normal human blood is $2.5 \times 10^{-7}$ mol/L. Calculate the hydrogen ion concentration and the pH of blood.

11. Acid rain has a pH less than that of normal rain. The presence of dissolved carbon dioxide, which forms carbonic acid, gives normal rain a pH of 5.6. What is the hydrogen ion concentration in normal rain?

12. If the pH of a solution changes by 3 pH units as a result of adding a weak acid, by how much does the hydrogen ion concentration change?

  

13. What mass of hydrogen chloride gas is required to produce 250 mL of a hydrochloric acid solution with a pH of 1.57?

14. Acetic (ethanoic) acid is the most common weak acid used in industry. When in water only 1.3% of the molecules ionize into hydronium ions. Determine the pH and pOH of an acetic acid solution prepared by dissolving 60.0 kg of pure, liquid acetic acid to make 1.25 kL of solution.

15. Determine the mass of sodium hydroxide that must be dissolved to make 2.00 L of a solution with a pH of 10.35.