# 5.2: pH and pOH

Learning Objectives

- Explain the characterization of aqueous solutions as acidic, basic, or neutral
- Express hydronium and hydroxide ion concentrations on the pH and pOH scales
- Perform calculations relating pH and pOH

As discussed earlier, hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water ($K_w$). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviors of its other solutes, and specific vocabulary has been developed to describe these concentrations in relative terms. A solution is neutral if it contains equal concentrations of hydronium and hydroxide ions; acidic if it contains a greater concentration of hydronium ions than hydroxide ions; and basic if it contains a lesser concentration of hydronium ions than hydroxide ions.

A common means of expressing quantities, the values of which may span many orders of magnitude, is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where "X" is the quantity of interest and "log" is the base-10 logarithm:

$$\mathrm{pX} = -\log \mathrm{X} \tag{5.2.1}$$

The pH of a solution is therefore defined as shown here, where $[H_3O^+]$ is the molar concentration of hydronium ion in the solution:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] \tag{5.2.2}$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression:

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}} \tag{5.2.3}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or pOH:

$$\mathrm{pOH} = -\log[\mathrm{OH^-}] \tag{5.2.4}$$

or

$$[\mathrm{OH^-}] = 10^{-\mathrm{pOH}} \tag{5.2.5}$$

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the $K_w$ expression:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] \tag{5.2.6}$$

$$-\log K_\mathrm{w} = -\log([\mathrm{H_3O^+}][\mathrm{OH^-}]) = -\log[\mathrm{H_3O^+}] + -\log[\mathrm{OH^-}] \tag{5.2.7}$$

$$\mathrm{p}K_\mathrm{w} = \mathrm{pH} + \mathrm{pOH} \tag{5.2.8}$$

At 25 °C, the value of $K_w$ is $1.0 \times 10^{-14}$, and so:

$$14.00 = \mathrm{pH} + \mathrm{pOH} \tag{5.2.9}$$

The hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7}\ M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{5.2.10}$$

$$\mathrm{pOH} = -\log[\mathrm{OH^-}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{5.2.11}$$

And so, at this temperature, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities less than $1.0 \times 10^{-7}\ M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7}\ M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

### When $pH = 7$ Solutions are not Neutral

Since the autoionization constant $K_w$ is temperature dependent, these correlations between pH values and the acidic/neutral/basic adjectives will be different at temperatures other than 25 °C. For example, the hydronium molarity of pure water at 80 °C is $4.9 \times 10^{-7}$ $M$, which corresponds to pH and pOH values of:

$$\begin{aligned} pH &= -\log[\mathrm{H_3O^+}] \\ &= -\log(4.9 \times 10^{-7}) \\ &= 6.31 \end{aligned}$$

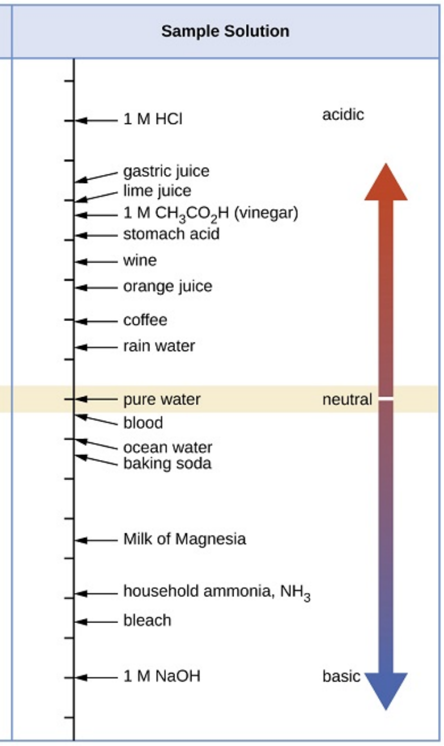
$$\begin{aligned} pOH &= -\log[\mathrm{OH^-}] \\ &= -\log(4.9 \times 10^{-7}) \\ &= 6.31 \end{aligned}$$

At this temperature, then, neutral solutions exhibit pH = pOH = 6.31, acidic solutions exhibit pH less than 6.31 and pOH greater than 6.31, whereas basic solutions exhibit pH greater than 6.31 and pOH less than 6.31. This distinction can be important when studying certain processes that occur at nonstandard temperatures, such as enzyme reactions in warm-blooded organisms. Unless otherwise noted, references to pH values are presumed to be those at standard temperature (25 °C) (Table 5.2.1).

Table 5.2.1: Summary of Relations for Acidic, Basic and Neutral Solutions

| Classification | Relative Ion Concentrations | pH at 25 °C |
|---|---|---|
| acidic | $[H_3O^+] > [OH^-]$ | pH < 7 |
| neutral | $[H_3O^+] = [OH^-]$ | pH = 7 |
| basic | $[H_3O^+] < [OH^-]$ | pH > 7 |

Figure 5.2.1 shows the relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH, and gives values for these properties at standard temperatures for some common substances.




| $[H_3O^+]$ (M) | $[OH^-]$ (M) | pH | pOH |
|---|---|---|---|
| $10^{1}$ | $10^{-15}$ | –1 | 15 |
| $10^{0}$ or 1 | $10^{-14}$ | 0 | 14 |
| $10^{-1}$ | $10^{-13}$ | 1 | 13 |
| $10^{-2}$ | $10^{-12}$ | 2 | 12 |
| $10^{-3}$ | $10^{-11}$ | 3 | 11 |
| $10^{-4}$ | $10^{-10}$ | 4 | 10 |
| $10^{-5}$ | $10^{-9}$ | 5 | 9 |
| $10^{-6}$ | $10^{-8}$ | 6 | 8 |
| $10^{-7}$ | $10^{-7}$ | 7 | 7 |
| $10^{-8}$ | $10^{-6}$ | 8 | 6 |
| $10^{-9}$ | $10^{-5}$ | 9 | 5 |
| $10^{-10}$ | $10^{-4}$ | 10 | 4 |
| $10^{-11}$ | $10^{-3}$ | 11 | 3 |
| $10^{-12}$ | $10^{-2}$ | 12 | 2 |
| $10^{-13}$ | $10^{-1}$ | 13 | 1 |
| $10^{-14}$ | $10^{0}$ or 1 | 14 | 0 |
| $10^{-15}$ | $10^{1}$ | 15 | –1 |

Figure 5.2.1: The pH and pOH scales represent concentrations of $[H_3O^+]$ and $OH^-$, respectively. The pH and pOH values of some common substances at standard temperature (25 °C) are shown in this chart.

### ✓ Example 5.2.1: Calculation of pH from $[H_3O^+]$

What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}\ M$?

**Solution**

$$\begin{aligned} pH &= -\log[H_3O^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) \\ &= 2.92 \end{aligned}$$

### ? Exercise 5.2.1

Water exposed to air contains carbonic acid, $H_2CO_3$, due to the reaction between carbon dioxide and water:

$$\mathrm{CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)}$$

Air-saturated water has a hydronium ion concentration caused by the dissolved $\mathrm{CO_2}$ of $2.0 \times 10^{-6}\ M$, about 20-times larger than that of pure water. Calculate the pH of the solution at 25 °C.

**Answer**

5.70

### ✓ Example 5.2.2: Calculation of Hydronium Ion Concentration from pH

Calculate the hydronium ion concentration of blood, the pH of which is 7.3 (slightly alkaline).

**Solution**

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = 7.3$$

$$\log[H_3O^+] = -7.3$$

$$[H_3O^+] = 10^{-7.3}$$

or

$$[H_3O^+] = \text{antilog of} - 7.3$$

$$[H_3O^+] = 5 \times 10^{-8}\ M$$

(On a calculator take the antilog, or the “inverse” log, of −7.3, or calculate $10^{-7.3}$.)

### Exercise 5.2.2

Calculate the hydronium ion concentration of a solution with a pH of −1.07.

**Answer**

12 M

### Example 5.2.3: Calculation of pOH

What are the pOH and the pH of a 0.0125-M solution of potassium hydroxide, KOH?

**Solution**

Potassium hydroxide is a highly soluble ionic compound and completely dissociates when dissolved in dilute solution, yielding $[OH^-]$ = 0.0125 M:

$$\text{pOH} = -\log[OH^-] = -\log 0.0125 \tag{5.2.12}$$

$$= -(-1.903) = 1.903 \tag{5.2.13}$$

The pH can be found from the pOH:

$$\text{pH} + \text{pOH} = 14.00 \tag{5.2.14}$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.903 = 12.10 \tag{5.2.15}$$

### Exercise 5.2.3

The hydronium ion concentration of vinegar is approximately $4 \times 10^{-3}\ M$. What are the corresponding values of pOH and pH?

**Answer**

pOH = 11.6,

pH = 14.00 - pOH = 2.4

The acidity of a solution is typically assessed experimentally by measurement of its pH. The pOH of a solution is not usually measured, as it is easily calculated from an experimentally determined pH value. The pH of a solution can be directly measured using a pH meter (Figure 5.2.2).

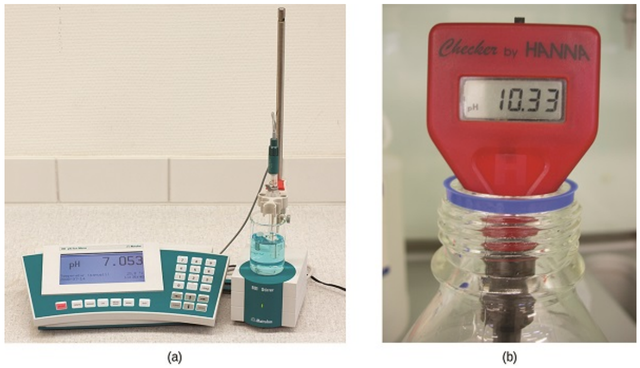



Figure 5.2.2: (a) A research-grade pH meter used in a laboratory can have a resolution of 0.001 pH units, an accuracy of ± 0.002 pH units, and may cost in excess of $1000. (b) A portable pH meter has lower resolution (0.01 pH units), lower accuracy (± 0.2 pH units), and a far lower price tag. (credit b: modification of work by Jacopo Werther)

The pH of a solution may also be visually estimated using colored indicators (Figure 5.2.3).


Figure 5.2.3: (a) A universal indicator assumes a different color in solutions of different pH values. Thus, it can be added to a solution to determine the pH of the solution. The eight vials each contain a universal indicator and 0.1-M solutions of progressively weaker acids: HCl (pH = l), $CH_3CO_2H$ (pH = 3), and $NH_4Cl$ (pH = 5), deionized water, a neutral substance (pH = 7); and 0.1-M solutions of the progressively stronger bases: KCl (pH = 7), aniline, $C_6H_5NH_2$ (pH = 9), $NH_3$ (pH = 11), and NaOH (pH = 13). (b) pH paper contains a mixture of indicators that give different colors in solutions of differing pH values. (credit: modification of work by Sahar Atwa).

## Summary


The concentration of hydronium ion in a solution of an acid in water is greater than $1.0 \times 10^{-7}\ M$ at 25 °C. The concentration of hydroxide ion in a solution of a base in water is greater than $1.0 \times 10^{-7}\ M$ at 25 °C. The concentration of $H_3O^+$ in a solution can be expressed as the pH of the solution; $\mathrm{pH} = -\log \mathrm{H_3O^+}$. The concentration of $OH^-$ can be expressed as the pOH of the solution: $\mathrm{pOH} = -\log[\mathrm{OH^-}]$. In pure water, pH = 7.00 and pOH = 7.00


## 5.2.1: Key Equations

- $\mathrm{pH} = -\log[\mathrm{H_3O^+}]$
- $\mathrm{pOH} = -\log[\mathrm{OH^-}]$
- $[H_3O^+] = 10^{-pH}$
- $[OH^-] = 10^{-pOH}$
- pH + pOH = $pK_w$ = 14.00 at 25 °C

## Glossary

**acidic**

describes a solution in which $[H_3O^+] > [OH^-]$

**basic**

describes a solution in which $[H_3O^+] < [OH^-]$

**neutral**

describes a solution in which $[H_3O^+] = [OH^-]$

**pH**

logarithmic measure of the concentration of hydronium ions in a solution

**pOH**

logarithmic measure of the concentration of hydroxide ions in a solution

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