# Acids & Bases

**Unit 14B:** The pH scale & pH calculations Acids & Bases

### Knowledge/Understanding Goals:

* the pH scale
* *K*w
* pH indicators

### Skills:

* calculate pH from [H3O+] and [H3O+] from pH

### Notes:

## H+ *vs.* H3O+

In water, H2O molecules exist in equilibrium with H3O+ (hydronium) and OH− (hydroxide) ions:

2 H2O (ℓ)  H3O+ (aq) + OH− (aq)

Most first-year inorganic chemistry courses[[1]](#footnote-1) use the Brønsted-Lowry definition that acids are H+ donors. However, H+ never actually exists in aqueous solution. The acid transfers the H+ directly to an H2O molecule creating an H3O+ ion. Thus while it is common to write H+ in a chemical equation (especially a net ionic equation), it is understood that the ion that is actually present in water is always H3O+. Thus, the above chemical equation is equivalent to:

H2O (ℓ)  H+ (aq) + OH− (aq)

## The Water Constant *K*w

The equilibrium for the reaction:

H2O (ℓ)  H3O+ (aq) + OH− (aq)

lies far to the left, so the equilibrium constant is very small:



(Recall that H2O is a pure liquid, so [H2O] is left out of the equilibrium expression.)

In water at 25°C, the product of the concentrations of H3O+ and OH− ions (in ) is therefore 1.008 × 10−14. This equilibrium constant is called the “water dissociation constant” *Kw*.

Note that as with any other equilibrium constant, the value of *Kw* is temperature dependent.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Temp. (°C) | *Kw* | pH | Temp. (°C) | *Kw* | pH |
| 0 | 0.114 × 10−14 | 7.47 | 40 | 2.916 × 10−14 | 6.77 |
| 5 | 0.186 × 10−14 | 7.37 | 50 | 5.476 × 10−14 | 6.63 |
| 10 | 0.293 × 10−14 | 7.27 | 60 | 9.550 × 10−14 | 6.51 |
| 15 | 0.457 × 10−14 | 7.17 | 70 | 15.85 × 10−14 | 6.40 |
| 20 | 0.681 × 10−14 | 7.08 | 80 | 25.12 × 10−14 | 6.30 |
| **25** | **1.008 × 10−14** | **7.00** | 90 | 38.02 × 10−14 | 6.21 |
| 30 | 1.471 × 10−14 | 6.92 | 100 | 51.3 × 10−14 | 6.14 |

This table shows that the pH of a neutral solution (*i.e.,* one for which [H3O+] = [OH−]) can vary from 6.14 to 7.47, depending on the temperature. Unless otherwise stated, you should assume a temperature of 25°C.

Remembering Le Châtelier’s principle, in the following equation at 25°C:

2 H2O  H3O+ + OH−

If we add acid, [H3O+] increases. This shifts the equilibrium to the left, which means [OH−] decreases, and [H3O+] [OH−] = 1.008 × 10−14 = *Kw*.

Similarly, if we add base, [OH−] increases and [H3O+] decreases, and [H3O+] [OH−] = 1.008 × 10−14 = *Kw*.

## The pH scale

If we have exactly the same amount of acid and base, then [H+] = [OH−] and both are equal (1.004 × 10−7*M* at 25°C). A solution with the same amount of acid and base is said to be *neutral*.

pH: a measure of the strength of an acidic or basic solution.

**pH = −log [H3O+]**

(The “p” in pH is a mathematical function that literally means “−log”.)

Examples:

if [H3O+] = 0.001 M, then pH = −log (0.001) = 3

if [H3O+] = 0.000 000 01 M (= 1 × 10−8 M) then pH = −log (1 × 10−8) = 8

Note that the higher the concentration of H+ ions (higher value of [H+], and therefore more acidic), the lower the pH. Similarly, the higher the concentration of OH− ions (and therefore more basic), the lower the concentration of H+ ions, and the higher the pH.

For some calculations, it is also useful to define the pOH scale:

**pOH = −log[OH−]**

### pH and pOH formulas

pH = −log[H3O+]  (or pH = −log[H+]) pOH =  −log[OH−]

[H3O+] = 10−pH [OH−] = 10−pOH

[H+][OH−] = 1.0 × 10−14 (at 25°C) pH + pOH = 14 (at 25°C)

Low pH = acidic = more H3O+ = less OH−  
High pH = basic = less H3O+ = more OH−

|  |  |  |  |
| --- | --- | --- | --- |
| **[H3O+]** | **[OH−]** | **pH** | **Acidic/Basic?** |
| 1*M* (= 1 × 100*M*) | 1 × 10−14*M* | 0 | very acidic |
| 0.1*M* (= 1 × 10−1*M*) | 1 × 10−13*M* | 1 |  |
| 0.01*M* (= 1 × 10−2*M*) | 1 × 10−12*M* | 2 |  |
| 1 × 10−3*M* | 1 × 10−11*M* | 3 |  |
| 1 × 10−4*M* | 1 × 10−10*M* | 4 |  |
| 1 × 10−5*M* | 1 × 10−9*M* | 5 | slightly acidic |
| 1 × 10−6*M* | 1 × 10−8*M* | 6 |  |
| 1 × 10−7*M* | 1 × 10−7*M* | 7 | neutral |
| 1 × 10−8*M* | 1 × 10−6*M* | 8 |  |
| 1 × 10−9*M* | 1 × 10−5*M* | 9 | slightly basic |
| 1 × 10−10*M* | 1 × 10−4*M* | 10 |  |
| 1 × 10−11*M* | 1 × 10−3*M* | 11 |  |
| 1 × 10−12*M* | 0.01*M* (= 1 × 10−2*M*) | 12 |  |
| 1 × 10−13*M* | 0.1*M* (= 1 × 10−1*M*) | 13 |  |
| 1 × 10−14*M* | 1*M* (= 1 × 100*M*) | 14 | very basic |

### Examples:

Q: What is the pH of a solution with [H3O+] = 2.5 x 10−4*M*?

A:

Q: What is the concentration of H3O+ ions in a solution with a pOH of 11.4?

A:

## pH of Strong *vs.* Weak Acid & Base Solutions

By definition, strong acids and bases dissociate completely in water. This means all of a strong acid will be converted to H3O+, and all of a strong base will be converted to OH−.

Therefore, in a strong acid:

[H3O+] = [acid] and pH = −log [acid]

(However, note that for extremely dilute solutions, the H3O+ and OH− from water are important. For example, if you start with water and add 10−12 *M* HCl, the pH would be 7, not 12.)

Weak acids and bases dissociate partially in water. This means that for a weak acid, [H3O+] < [acid] and pH > −log [acid]. By a similar argument, for a weak base, pH < −log [acid]. An easy way to remember is that if the acid or base is only partially dissociated, the pH will be closer to 7 (the pH of pure H2O).

We will discuss how to calculate pH values for these weak solutions later.

**Indicators**

indicator: a substance that changes color at a specific pH value. Indicators are used as a visual way to measure pH.

|  |  |  |  |
| --- | --- | --- | --- |
| **Name of  Indicator** | **color in acid** | **color in base** | **pH range of color change** |
| bromophenol blue | yellow | purple | 3.0-4.6 |
| methyl red | red | yellow | 4.4-6.2 |
| Litmus | red | blue | 5.5-8.2 |
| bromothymol blue | yellow | blue | 6.0-7.6 |
| phenol red | yellow | red | 6.8-8.4 |
| Phenolphthalein | clear | pink | 8.2-10.0 |

1. *I.e., college* first-year chemistry courses, which includes AP Chemistry [↑](#footnote-ref-1)