pH & Indicators

Unit: Acids & Bases

MA Curriculum Frameworks (2016): HS-PS1-9(MA)

Mastery Objective(s): (Students will be able to…)

- Calculate pH from [H^+] and pOH from [OH^-].
- Identify acids and bases from their pK_a values.
- Select an appropriate indicator for a desired pH range.

Success Criteria:

- pH and pOH are calculated correctly.
- Acids and bases are correctly identified from their pK_a values.
- Indicator changes color in a pH range that includes the pH of the given acid or base.

Tier 2 Vocabulary: acid, base, indicator

Language Objectives:

- Explain why higher [H^+] results in a lower pH.

Notes:

In water, a very small amount of H_2O dissociates into H^+ and OH^- ions:

$$H_2O \rightleftharpoons H^+ + OH^-$$

The amount of dissociation of any compound in a solvent is a constant that is determined by the attractions of the ions for each other *vs.* the attraction between the ions and the solvent.

In water at 25 °C, the product of the concentrations of H^+ and OH^- ions (in $\frac{\text{mol}}{\text{L}}$) is 1.0×10^{-14}. This number is called the "water dissociation constant" K_w*. In other words, in water at 25 °C:

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

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* K_w is actually the equilibrium constant for the dissociation reaction. $K_{eq} = \frac{[H^+][OH^-]}{[H_2O]}$.

However, because H_2O is a pure liquid, the concentration of H_2O in pure H_2O is constant—it's just the density divided by the molar mass, which works out to 55.6 M. Therefore, we leave [H_2O] out of the equilibrium expression.

Use this space for summary and/or additional notes:

Recall that acids create H^+ (or H_3O^+) in water, and bases create OH^- in water. In the dissociation equation:

$$H_2O \rightleftharpoons H^+ + OH^-$$

Le Châtelier's principle predicts that if we add acid, $[H^+]$ increases. This shifts the equilibrium to the left, which means $[OH^-]$ decreases, and it is still true that $[H^+]\,[OH^-] = 1.0 \times 10^{-14} = K_w$.

Similarly, if we add base, $[OH^-]$ increases and $[H^+]$ decreases and $[H^+]\,[OH^-] = 1.0 \times 10^{-14} = K_w$.

If we have exactly the same amount of acid and base, then $[H^+] = [OH^-]$ and both are equal to the square root of 1×10^{-14}, which is 1×10^{-7} M. A solution with the same amount of acid and base is said to be *neutral*.

Working with concentrations in scientific notation that vary over 14 powers of ten is unwieldy, so we define a function "p" which means "take the logarithm of the quantity and multiply the result by −1." (See the "Logarithms" topic starting on page 99 for a brief description of the logarithm mathematical function.)

Therefore, the quantity "pH" would be $-\log\,[H^+]$.

pH: a measure of the strength of an acidic or basic solution. Equal to $-\log\,[H^+]$.

Examples:

if $[H^+] = 0.001$ M, then $\text{pH} = -\log(0.001) = 3$
if $[H^+] = 0.000\,000\,01$ M ($= 1 \times 10^{-8}$ M) then $\text{pH} = -\log(1 \times 10^{-8}) = 8$

pOH: another measure of the strength of an acidic or basic solution. Equal to $-\log\,[OH^-]$. Much less commonly used that pH.

Examples:

if $[OH^-] = 0.001$ M, then $\text{pOH} = -\log(0.001) = 3$
if $[OH^-] = 0.000\,000\,01$ M ($= 1 \times 10^{-8}$ M) then $\text{pOH} = -\log(1 \times 10^{-8}) = 8$

pH & pOH Equations

$$\text{pH} = -\log[H^+] \qquad \text{pOH} = -\log[OH^-]$$

$$[H^+] = 10^{-\text{pH}} \qquad [OH^-] = 10^{-\text{pOH}}$$

Because the (multiplication) product of $[H^+][OH^-] = 1 \times 10^{-14}$, this means that:

$$\text{pH} + \text{pOH} = 14$$

Use this space for summary and/or additional notes:

Note that the higher the concentration of H^+ ions (higher value of $[H^+]$), the lower the pH.

Low pH = acidic = more H^+ = less OH^-
High pH = basic = less H^+ = more OH^-

$[H^+]$	$[OH^-]$	pH	pOH	Acidic/Basic?
1 M (= 1×10^0 M)	1×10^{-14} M	0	14	very acidic
0.1 M (= 1×10^{-1} M)	1×10^{-13} M	1	13	
0.01 M (= 1×10^{-2} M)	1×10^{-12} M	2	12	
1×10^{-3} M	1×10^{-11} M	3	11	
1×10^{-4} M	1×10^{-10} M	4	10	
1×10^{-5} M	1×10^{-9} M	5	9	slightly acidic
1×10^{-6} M	1×10^{-8} M	6	8	
1×10^{-7} M	1×10^{-7} M	7	7	neutral
1×10^{-8} M	1×10^{-6} M	8	6	
1×10^{-9} M	1×10^{-5} M	9	5	slightly basic
1×10^{-10} M	1×10^{-4} M	10	4	
1×10^{-11} M	1×10^{-3} M	11	3	
1×10^{-12} M	0.01 M (= 1×10^{-2} M)	12	2	
1×10^{-13} M	0.1 M (= 1×10^{-1} M)	13	1	
1×10^{-14} M	1 M (= 1×10^0 M)	14	0	very basic

Sample Problems:

Q: What is the pH of a solution with $[H^+] = 2.5 \times 10^{-4}$ M?

A: $-\log(2.5 \times 10^{-4}) = 3.60$

Q: What is the concentration of H^+ ions in a solution with a pH of 11.4?

A: $10^{-11.4} = 3.98 \times 10^{-12}$ M

Use this space for summary and/or additional notes:

An aqueous solution is neutral when the concentration of H^+ and OH^- are equal. This occurs in water at pH 7.00 at a temperature of 25 °C. However, remember that temperature affects equilibrium; as the temperature increases, more H^+ and OH^- dissociate. This means $[H^+]$ and $[OH^-]$ *both* increase with higher temperatures, which means K_w increases. When that happens, $[H^+]$ and $[OH^-]$ are still equal in a neutral solution, but both are larger, and because $[H^+]$ and $[OH^-]$ are larger, the pH and pOH are both lower.

Temp. (°C)	K_w	**pH of a neutral solution**
0	0.114×10^{-14}	7.47
10	0.293×10^{-14}	7.27
20	0.681×10^{-14}	7.08
25	1.008×10^{-14}	7.00
30	1.471×10^{-14}	6.92
40	2.916×10^{-14}	6.77
50	5.476×10^{-14}	6.63
100	51.3×10^{-14}	6.14

In other words, despite what your previous teachers may have taught you, a pH of 7 is only neutral at 25 °C. In fact, in warm-blooded animals with body temperatures around 37 °C, a neutral pH would be approximately 6.8.

This also means that pH + pOH = 14 is only correct at 25 °C.

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Indicators

<u>indicator</u>: a substance that changes color in a specific range of pH values. Indicators are used as a visual way to measure pH.

The following table lists some common indicators.

Name of Indicator	color in acid	color in base	pH range where color change occurs
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

There are many others, and multiple indicators can be used in order to have different color changes over a broader pH range.

In fact, some clever chemists have developed a "universal indicator," which is typically composed of water, propanol, phenolphthalein, sodium hydroxide, methyl red, bromothymol blue, and thymol blue. This mixture indicates pH over a range from 3 to 11, in ROYGBIV (rainbow) order:

pH range	Description	Color
< 3	Strong acid	red
3–6	Weak acid	orange or yellow
7	Neutral	green
8–11	Weak base	blue
> 11	Strong base	indigo (dark blue) or violet (purple)

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Homework Problems

For each of the following solutions, calculate the information indicated. Choose pH indicators from the "Common Acid-Base Indicators" table in your reference packets.

1. $[H+] = 2.5 \times 10^{-4}$ M
   a. pH =
   b. Is the solution acidic, basic, or neutral?
   c. Which pH indicator would be best for this solution?
2. $[H+] = 4.59 \times 10^{-7}$ M
   a. pH =
   b. Is the solution acidic, basic, or neutral?
   c. Which pH indicator would be best for this solution?
3. pH = 9.1
   a. [H+] =
   b. Is the solution acidic, basic, or neutral?
   c. Which pH indicator would be best for this solution?

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4. pH = 5.5
    a. [H+] =
    b. Is the solution acidic, basic, or neutral?
    c. Which pH indicator would be best for this solution?
5. $[OH^-] = 7.9 \times 10^{-7}$ M
    a. [H+] =
    b. pH =
    c. Is the solution acidic, basic, or neutral?
    d. Which pH indicator would be best for this solution?

Use this space for summary and/or additional notes: