Lect.7.

$$K_w = 1.0 \times 10^{-14}$$

- The product, $[H_3O^+][^-OH]$, is a constant, 1.0×10^{-14}, for all aqueous solutions at 25 °C.

Thus, the value of Kw applies to any aqueous solution, not just pure water.

If we know the concentration of one ion, H_3O^+ or ^-OH, we can find the concentration of the other by rearranging the expression for Kw.

To calculate $[^-OH]$ when $[H_3O^+]$ is known:

$$K_w = [H_3O^+][^-OH]$$

$$[^-OH] = \frac{K_w}{[H_3O^+]}$$

$$[^-OH] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$

To calculate $[H_3O^+]$ when $[^-OH]$ is known:

$$K_w = [H_3O^+][^-OH]$$

$$[H_3O^+] = \frac{K_w}{[^-OH]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[^-OH]}$$

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Pure water and any solution that has an equal concentration of H_3O^+ and –OH ions (1.0×10^{-7}) is said to be neutral.

Other solutions are classified as acidic or basic, depending on which ion is present **in a higher concentration.**

In an acidic solution, $[H_3O^+] > [–OH]$; thus, $[H_3O^+] > 10^{-7}$ M

In a basic solution, $[^-OH] > [H3O^+]$; thus, $[^-OH] > 10^{-7}$ M

Lect.7.

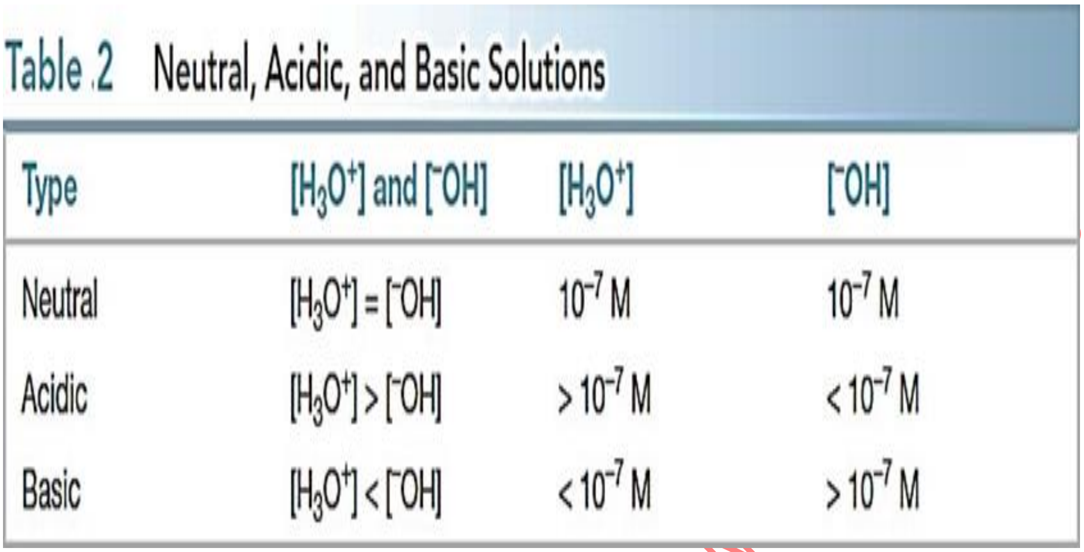
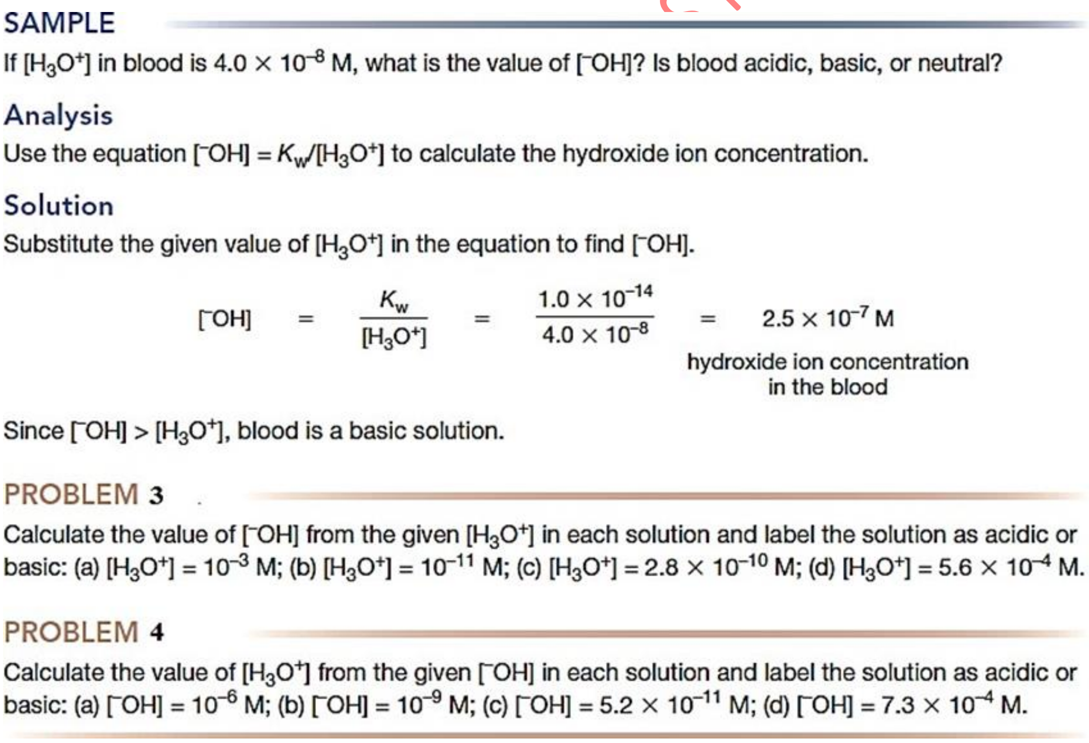


Table 2 Neutral, Acidic, and Basic Solutions

Type	$[H_3O^+]$ and $[^-OH]$	$[H_3O^+]$	$[^-OH]$
Neutral	$[H_3O^+] = [^-OH]$	10^{-7} M	10^{-7} M
Acidic	$[H_3O^+] > [^-OH]$	$> 10^{-7}$ M	$< 10^{-7}$ M
Basic	$[H_3O^+] < [^-OH]$	$< 10^{-7}$ M	$> 10^{-7}$ M



SAMPLE

If $[H_3O^+]$ in blood is 4.0×10^{-8} M, what is the value of $[^-OH]$? Is blood acidic, basic, or neutral?

Analysis

Use the equation $[^-OH] = K_w/[H_3O^+]$ to calculate the hydroxide ion concentration.

Solution

Substitute the given value of $[H_3O^+]$ in the equation to find $[^-OH]$.

$$[^-OH] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-8}} = \underset{\text{hydroxide ion concentration in the blood}}{2.5 \times 10^{-7}\ \text{M}}$$

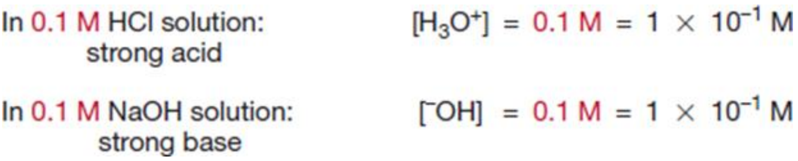
Since $[^-OH] > [H_3O^+]$, blood is a basic solution.

PROBLEM 3

Calculate the value of $[^-OH]$ from the given $[H_3O^+]$ in each solution and label the solution as acidic or basic: (a) $[H_3O^+] = 10^{-3}$ M; (b) $[H_3O^+] = 10^{-11}$ M; (c) $[H_3O^+] = 2.8 \times 10^{-10}$ M; (d) $[H_3O^+] = 5.6 \times 10^{-4}$ M.

PROBLEM 4

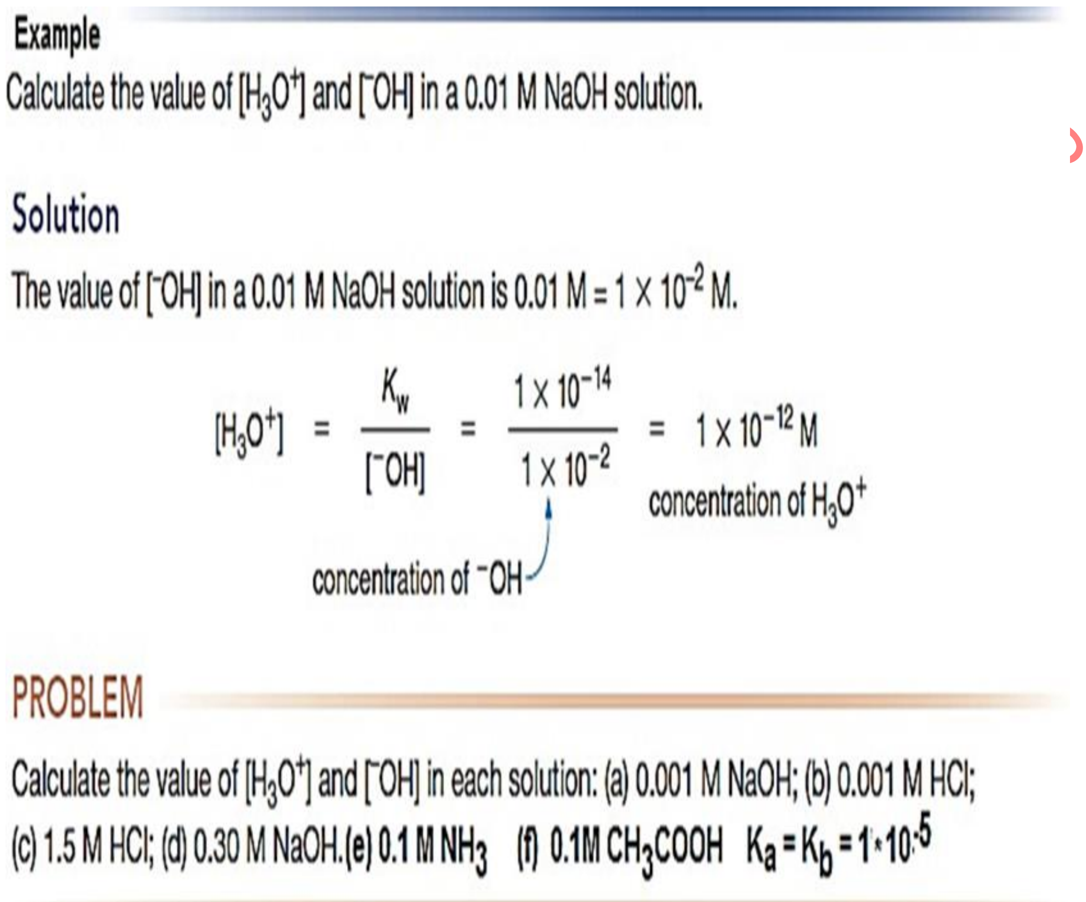
Calculate the value of $[H_3O^+]$ from the given $[^-OH]$ in each solution and label the solution as acidic or basic: (a) $[^-OH] = 10^{-6}$ M; (b) $[^-OH] = 10^{-9}$ M; (c) $[^-OH] = 5.2 \times 10^{-11}$ M; (d) $[^-OH] = 7.3 \times 10^{-4}$ M.



In 0.1 M HCl solution: $[H_3O^+] = 0.1\ \text{M} = 1 \times 10^{-1}$ M
strong acid

In 0.1 M NaOH solution: $[^-OH] = 0.1\ \text{M} = 1 \times 10^{-1}$ M
strong base

Lect.7.



Example

Calculate the value of $[H_3O^+]$ and $[^-OH]$ in a 0.01 M NaOH solution.

Solution

The value of $[^-OH]$ in a 0.01 M NaOH solution is 0.01 M = 1×10^{-2} M.

$$[H_3O^+] = \frac{K_w}{[^-OH]} = \frac{1 \times 10^{-14}}{1 \times 10^{-2}} = 1 \times 10^{-12}\ M$$

concentration of H_3O^+

concentration of ^-OH

PROBLEM

Calculate the value of $[H_3O^+]$ and $[^-OH]$ in each solution: (a) 0.001 M NaOH; (b) 0.001 M HCl; (c) 1.5 M HCl; (d) 0.30 M NaOH.**(e) 0.1 M NH_3 (f) 0.1M CH_3COOH $K_a = K_b = 1*10^{-5}$**

Lect.7.

7.1 The pH Scale.

Knowing the pH value of a solution or fluid is very important for many chemical and analytical tasks and its measurement determines any follow up measurements.

Taking a pH measurement often seems to be trivial, which is the reason why pH measurements are frequently not questioned.

But to make a useful pH measurement close attention must be paid to the measurement's details.

To make a proper pH measurement and avoid errors you must first be familiar with the basics of pH measurement.

Lect.7.

The elementary questions are:

What defines the pH-value?

How do I measure the pH^{-}value?

Where and why are pH measurements made?

The concentrations of hydrogen ions and indirectly hydroxide ions are given by a pH number.

pH is defined as the negative logarithm of the hydrogen ion concentration. The equation is:

$pH = - \log [H^+]$

similarly, $pOH = - \log [OH^-]$

and $p\,Kw = - \log [Kw]$

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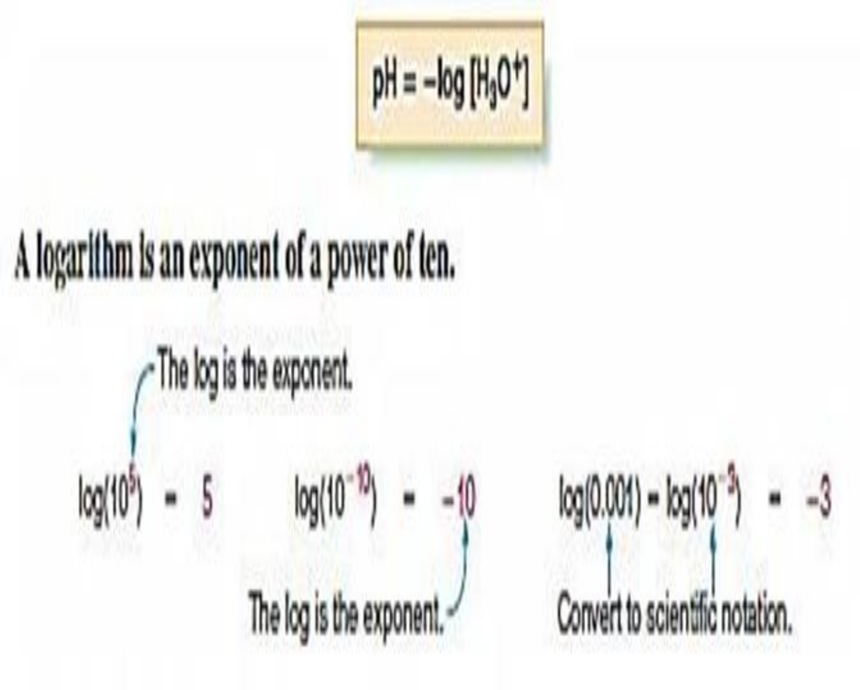

$$pH = -\log [H_3O^+]$$

A logarithm is an exponent of a power of ten.

$\log(10^5) = 5$ $\log(10^{-10}) = -10$ $\log(0.001) = \log(10^{-3}) = -3$

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Logarithms of numbers that are multiples of ten are merely the exponents of the number including the sign. See the table on the left for a review.

The method to find logs of numbers that are not multiples of ten are found by using a calculator.

The method is not discussed here.

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Example

If an acid has an H^+ concentration of 0.0001 M, find the pH?

Solution: First convert the number to exponential notation, find the log, then solve the pH equation.

$H^+ = 0.0001M = 10^{-4}$; log of $10^{-4} = {}^{-}4$

$pH = -\log [H^+] = -\log (10^{-4}) = {}^{-}(-4) = {}^{+}4 = pH$

The purpose of the negative sign in the log definition is to give a positive pH value.

Lect.7.

Example:

If the base has an OH- concentration of 0.001M, find the pH.

Solution: First find the pOH, (similar to finding the pH,) then subtract the pOH from 14.

$OH^- = 0.001M = 10^{-3}$

pOH = -log [OH-] = -log (10-3) = $^+3$ = pOH

pH = 14 - pOH

pH = 14 - 3 = 11 = pH

The pH scale, (0 - 14), is the full set of pH numbers which indicate the concentration of H^+ and OH^- ions in water.

The diagram on the left gives some relationships which summarizes much of the previous discussion.

Whether a solution is acidic, neutral, or basic can now be defined in terms of its pH.

Lect.7.

Acidic solution: pH < 7 $[H_3O^+]$ > 1 × 10^{-7}

Neutral solution: pH = 7 $[H_3O^+]$ = 1 × 10^{-7}

Basic solution: pH > 7 $[H_3O^+]$ < 1 × 10^{-7}

Lect.7

Note the relationship between [H_3O^+] and pH.

The lower the pH, the higher the concentration of H_3O^+

H^+ ion concentration and pH relate inversely.

OH^- ion concentration and pH relate directly.