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Chemistry Teach Yourself Series

Topic 1: pH

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pH

What is it?: a measure of the acidity or alkalinity of a solution.

$$pH = -\log_{10}[H_3O^+]$$

pH formula: Derivation of formula

Unit 2 Chemistry

$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$

$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$

$CH_3COOH(l) + H_2O(l) \longleftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

Acid definition. A substance that can donate a proton (H^+)

CH_3COOH is a weak acid. This means that it only donates a proportion of its protons.

$\left.\begin{array}{l} \text{2 M HCl} \\ \text{2 M } HNO_3 \end{array}\right\}$ dangerous solution

2 M CH_3COOH not a dangerous solution

Concentration of the acid itself does not indicate the acidity of the solution that will form.

2 M HCl	leads to 2 M	H_3O^+	solution
2 M HNO_3	leads to 2 M	H_3O^+	solution
2 M CH_3COOH	leads to 0.001 M	H_3O^+	solution

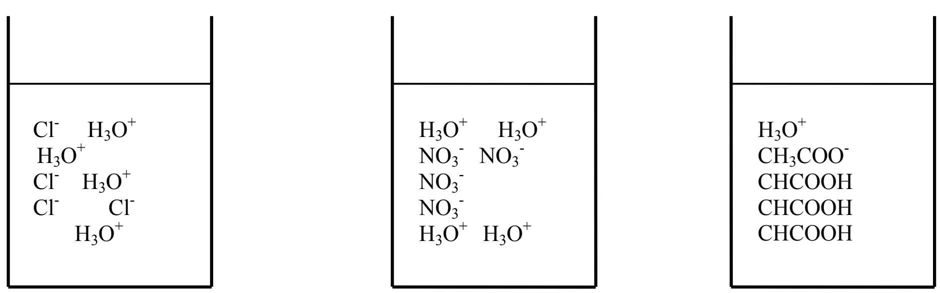

Diagram: The first two solutions are strong acids, the third is a weak acid.

Note that $[H_3O^+]$ is a better measure of acidity, hence it is used in pH formula.

Lactic acid has a $[H_3O^+]$ of 0.000001 M.

That is a very small number, hence the pH formula uses log to make this a more familiar number $\log_{10}(0.000001) = -6$.

$[H_3O^+]$ is the best measure of acid strength

$$pH = -\log_{10}[H_3O^+]$$

Negative sign converts most answers to positive number

log scale makes small numbers manageable

pH of strong acids

Unit 2 Chemistry

Example

Calculate the pH of the following strong acids

a. 0.01 M HCl
b. 0.0001 M HNO_3

Solution

Strong acid => $[H_3O^+]$ assumed = concentration of the acid itself

a. [HCl] = 0.01 => $[H_3O^+]$ = 0.01 => pH = $-\log_{10}(0.01)$ = 2

b. [HCl] = 0.0001 => $[H_3O^+]$ = 0.0001 => pH = $-\log_{10}(0.0001)$ = 4

Short cut
If the concentration is a simple power of 10, a calculator is not required i.e.

a. pH = $-\log_{10}(0.01)$

pH = $-\log_{10}(10^{-2})$ = 2

b. pH = $-\log_{10}(0.0001)$

pH = $-\log_{10}(10^{-4})$ = 4

note that the answer = the power of 10 with sign changed

This technique is helpful when working in reverse;

If the pH - is 4, then $[H_3O^+]$ = 10^{-4}
- is 2, then $[H_3O^+]$ = 10^{-2}

Questions

1. Calculate the pH of the following solutions of nitric acid, HNO_3

a. 1.0 M ____________________

b. 0.1 M ____________________

c. 0.00001 M ____________________

2. Complete the table below

concentration M	**pH**
1.0	0
0.10	
0.001	
	1
	4
	6

3. Given the pH of the following solutions, what is the hydronium ion concentration?

a. pH = 1.0 ________________________

b. pH = 5.0 ________________________

Alkaline solutions

Unit 2 Chemistry

Solutions containing or producing OH^- ions are alkaline i.e. NaOH or NH_3

The amount of $[H_3O^+]$ has to be found using the formula

$$[H_3O^+][OH^-] = 10^{-14} \quad M^2 \quad \text{at } 25\ ^0C.$$

Example

Calculate the pH of the following alkaline solutions

a. 0.01 M LiOH
b. 0.5 M NaOH

Solutions

Example a. can be completed without a calculator because powers of 10 are used.
Example b. is best completed with a calculator.

a. $[H_3O^+][OH^-] = 10^{-14}$

$[OH^-] = 0.01 = 10^{-2}$

$\Rightarrow \quad [H_3O^+] \times 10^{-2} = 10^{-14}$

$\Rightarrow \quad [H_3O^+] = \frac{10^{-14}}{10^{-2}} = 10^{-12}$ $\qquad \Rightarrow \quad pH = -\log(10^{-12}) = 12$

b. $[H_3O^+][OH^-] = 10^{-14}$

$[OH^-] = 0.5$

$\Rightarrow \quad [H_3O^+] \times 0.5 = 10^{-14}$

$\Rightarrow \quad [H_3O^+] = \frac{10^{-14}}{0.5} = 2.0x10^{-14}$ $\qquad \Rightarrow \quad pH = -\log(10^{-12}) = 13.7$

Example

c. Calculate the pH of a 0.005 M $Mg(OH)_2$

Note: pH uses the $[OH^-]$. If the $Mg(OH)_2$ is 0.005, then $[OH^-]$. = 2 x 0.005 = 0.01 M

Therefore the working is exactly the same as example **a** above and the answer is the same, 12.

Question

4. Calculate the pH of a solution of

a. 0.001 M NaOH

Care: watch for 2 here

b. b. 0.01 M $Mg(OH)_2$

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pH of weak acids

Unit 4 Chemistry

CH_3COOH is a weak acid. It only donates a small percentage of its protons.

$CH_3COOH(l) + H_2O(l) \Leftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

The $[H_3O^+]$ concentration is much less than the CH_3COOH concentration, therefore the process of calculating pH is different.

$$K = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH][H_2O]} \quad \text{hence} \quad K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

Questions will – give K_a and ask for pH **OR** give pH and ask for K_a

Examples

a. A 0.05 M ethanoic acid solution has a pH of 3.2. Calculate the value of K_a for the ethanoic acid.

b. The K_a value for a 0.01 M sample of hydocyanic acid is 6.3×10^{-10}. Calculate the pH of the solution.

Solutions

a. $K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$

If pH = 3.2, then $[H_3O^+] = 10^{-3.2} = 6.31 \times 10^{-4}$

Since CH_3COOH was the only acid added, $[H_3O^+] = [CH_3COO^-] = 6.31 \times 10^{-4}$

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{6.31 \times 10^{-4} \times 6.31 \times 10^{-4}}{0.05} = 7.96 \times 10^{-6} \text{ M}$$

b. $K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$

To calculate pH, the $[H_3O^+]$ needs to be determined.

If $[H_3O^+]$ = X, then $[CN^-]$ also = X

$\Rightarrow Ka = \frac{X \times X}{0.01} = 6.3 \times 10^{-10}$

$\Rightarrow X = 6.3 \times 10^{-10} \times 0.01 = 6.3 \times 10^{-12}$

$\Rightarrow X = 2.5 \times 10^{-6}$

$pH = -\log_{10}[H_3O^+]$

$= -\log_{10}(2.5 \times 10^{-6}) = 5.6$

Questions

5. Calculate the *Ka* of a 0.01 M solution of hydrocyanic acid if the pH is 4.7.

6. The acidity constant for ethanoic acid is 1.7×10^{-5}. Calculate the pH of a 0.05 M solution of acid.

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Solutions to Questions

1.
a. pH = $-\log_{10}(1.0) = 0$
b. pH = $-\log_{10}(0.10) = 1$
c. pH = $-\log_{10}(0.00001) = 5$

2.

concentration M	pH
1.0	0
0.10	1
0.001	3
0.10	1
0.0001	4
0.000001	6

3.
a. $[H_3O^+] = 10^{-1} = 0.1$
b. $[H_3O^+] = 10^{-5} = 0.00001$

4.
a. $[H_3O^+][OH^-] = 10^{-14}$
$[OH^-] = 0.001 = 10^{-3}$
=> $[H_3O^+] \times 10^{-3} = 10^{-14}$
=> $[H_3O^+] = \dfrac{10^{-14}}{10^{-3}} = 10^{-11}$ => pH $= -\log(10^{-11}) = 11$

b. $[H_3O^+][OH^-] = 10^{-14}$
$[OH^-] = 0.01 \times 2 = 0.02$
=> $[H_3O^+] \times 0.02 = 10^{-14}$
=> $[H_3O^+] = \dfrac{10^{-14}}{0.02} = 5x10^{-13}$ => pH $= -\log(5 \times 10^{-13}) = 12.3$

5. If pH = 4.7, then $[H_3O^+] = 10^{-4.7} = 2.0 \times 10^{-5}$

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} = \frac{2.0 \times 10^{-5} \times 2.0 \times 10^{-5}}{0.05} = 8.0 \times 10^{-9}\ M$$

6. The acidity constant for ethanoic acid is $1.7x10^{-5}$. Calculate the pH of a 0.05 M solution of acid.

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

To calculate pH, the $[H_3O^+]$ needs to be determined.

If $[H_3O^+]$ = X, then $[CH_3COO^-]$ also = X

$$\Rightarrow Ka = \frac{X \text{ x } X}{0.01} = 1.7x10^{-5}$$

$$\Rightarrow X^2 = 1.7x10^{-5} \text{ x } 0.01 = 1.7x10^{-7}$$

$$\Rightarrow X = 4.1 \text{ x } 10^{-4}$$

$$pH = -\log_{10}[4.1 \text{ x } 10^{-4}]$$

$$= 3.38$$