



Chemistry

Teach Yourself Series

Topic 1: pH

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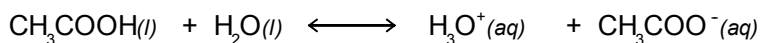
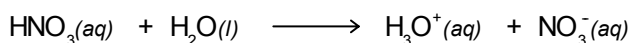
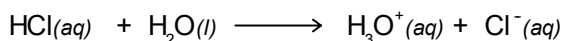
pH

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

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

pH formula: Derivation of formula

As it appears in Unit 2



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.

2 M HCl	}	dangerous solution
2 M HNO_3		

2 M CH_3COOH	not a dangerous solution
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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

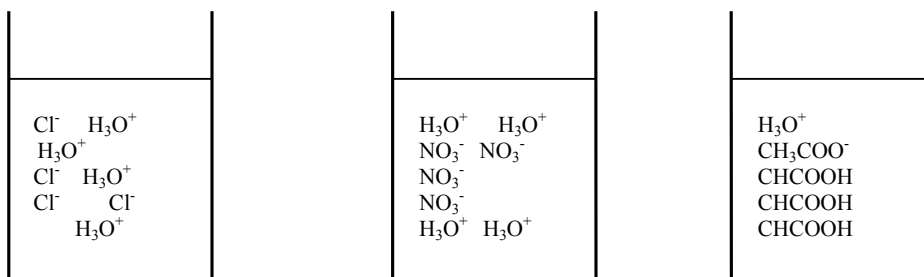


Diagram: The first two solutions are strong acids, the third is a weak acid.

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

As it appears in Unit 2

Example

Calculate the pH of the following strong acids

- a. 0.01 M HCl
- b. 0.0001 M HNO₃

Solution

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

- a. $[HCl] = 0.01 \Rightarrow [H_3O^+] = 0.01 \Rightarrow pH = -\log_{10}(0.01) = 2$
- b. $[HCl] = 0.0001 \Rightarrow [H_3O^+] = 0.0001 \Rightarrow pH = -\log_{10}(0.0001) = 4$

Short cut

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

- a. $pH = -\log_{10}(0.01)$
- $$\left. \begin{array}{l} pH = -\log_{10}(10^{-2}) = 2 \\ pH = -\log_{10}(0.0001) \\ pH = -\log_{10}(10^{-4}) = 4 \end{array} \right\} \text{note that the answer = the power of 10 with sign changed}$$

This technique is helpful when working in reverse;

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

Review 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. $\text{pH} = 1.0$ _____

b. $\text{pH} = 5.0$ _____

Alkaline solutions

As it appears in Unit 2

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

The amount of $[\text{H}_3\text{O}^+]$ has to be found using the formula

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14} \text{ M}^2 \text{ at } 25^\circ\text{C}.$$

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. $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$

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

$$\Rightarrow [\text{H}_3\text{O}^+] \times 10^{-2} = 10^{-14}$$

$$\Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{10^{-2}} = 10^{-12} \quad \Rightarrow \text{pH} = -\log(10^{-12}) = 12$$

b. $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$

$$[\text{OH}^-] = 0.5$$

$$\Rightarrow [\text{H}_3\text{O}^+] \times 0.5 = 10^{-14}$$

$$\Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{0.5} = 2.0 \times 10^{-14} \quad \Rightarrow \text{pH} = -\log(10^{-12}) = 13.7$$

Example

- c. Calculate the pH of a 0.005 M $\text{Mg}(\text{OH})_2$

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

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

Review Question

4. Calculate the pH of a solution of

a. 0.001 M NaOH

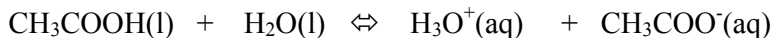
Care: watch for 2 here

b. b. 0.01 M $\text{Mg}(\text{OH})_2$ 

pH of weak acids

As it appears in Unit 4

CH₃COOH is a weak acid. It only donates a small percentage of its protons.



The [H₃O⁺] concentration is much less than the CH₃COOH concentration, therefore the process of calculating pH is different.

$$K = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}][\text{H}_2\text{O}]} \quad \text{hence} \quad K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

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 hydrocyanic acid is 6.3×10^{-10} . Calculate the pH of the solution.

Solutions

a. $K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$

If pH = 3.2, then $[\text{H}_3\text{O}^+] = 10^{-3.2} = 6.31 \times 10^{-4}$

Since CH₃COOH was the only acid added, $[\text{H}_3\text{O}^+] = [\text{CH}_3\text{COO}^-] = 6.31 \times 10^{-4}$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{6.31 \times 10^{-4} \times 6.31 \times 10^{-4}}{0.05} = 7.96 \times 10^{-6} \text{ M}$$

b. $K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$

To calculate pH, the [H₃O⁺] needs to be determined.

If $[\text{H}_3\text{O}^+] = X$, then $[\text{CN}^-]$ also = X

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

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

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

$$\begin{aligned}\text{pH} &= -\log_{10}[\text{H}_3\text{O}^+] \\ &= -\log_{10}(2.5 \times 10^{-6}) = 5.6\end{aligned}$$

Review Questions

5. Calculate the K_a 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.

Solutions to Review Questions

1.

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

2.

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

3.

- a. $[\text{H}_3\text{O}^+] = 10^{-1} = 0.1$
- b. $[\text{H}_3\text{O}^+] = 10^{-5} = 0.00001$

4.

- a. $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$
 $[\text{OH}^-] = 0.001 = 10^{-3}$
 $\Rightarrow [\text{H}_3\text{O}^+] \times 10^{-3} = 10^{-14}$
 $\Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{10^{-3}} = 10^{-11} \quad \Rightarrow \text{pH} = -\log(10^{-11}) = 11$
- b. $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$
 $[\text{OH}^-] = 0.01 \times 2 = 0.02$
 $\Rightarrow [\text{H}_3\text{O}^+] \times 0.02 = 10^{-14}$
 $\Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{0.02} = 5 \times 10^{-13} \quad \Rightarrow \text{pH} = -\log(5 \times 10^{-13}) = 12.3$

5. If $\text{pH} = 4.7$, then $[\text{H}_3\text{O}^+] = 10^{-4.7} = 2.0 \times 10^{-5}$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = \frac{2.0 \times 10^{-5} \times 2.0 \times 10^{-5}}{0.05} = 8.0 \times 10^{-9} \text{ M}$$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

To calculate pH, the $[\text{H}_3\text{O}^+]$ needs to be determined.

If $[\text{H}_3\text{O}^+] = X$, then $[\text{CH}_3\text{COO}^-]$ also = X

$$\Rightarrow K_a = \frac{X \times X}{0.01} = 1.7 \times 10^{-5}$$

$$\Rightarrow X^2 = 1.7 \times 10^{-5} \times 0.01 = 1.7 \times 10^{-7}$$

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

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

$$= 3.38$$