Analytical chemistry

Chemical engineering department

First class / first term

Al-Mustaqbal-college

Lecture five

acidity or the basicity of a substance

By

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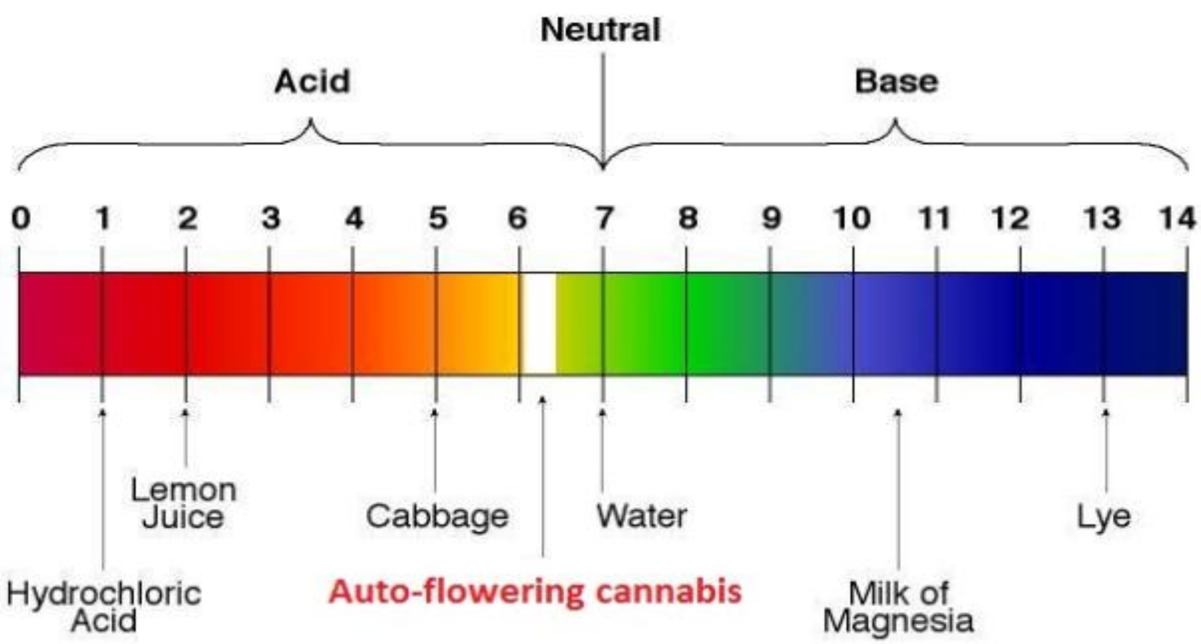
Lecture Three

⇒ pH and pOH

pH scale is a commonly used scale to measure the acidity or the basicity of a substance. The possible values on the pH scale range from 0 to 14. Acidic substances have pH values ranging from 1 to 7 (1 being the most acidic point on the pH scale) and alkaline or basic substances have pH values ranging from 7 to 14. A perfectly neutral substance would have a pH of exactly 7.

pH which is an abbreviation of 'potential for hydrogen' or 'power of hydrogen' of a substance can be expressed as the negative logarithm (with base 10) of the hydrogen ion concentration in that substance. Similarly, the pOH of a substance is the negative logarithm of the hydroxide ion concentration in the substance. These quantities can be expressed via the following formulae:

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



The pH Scale

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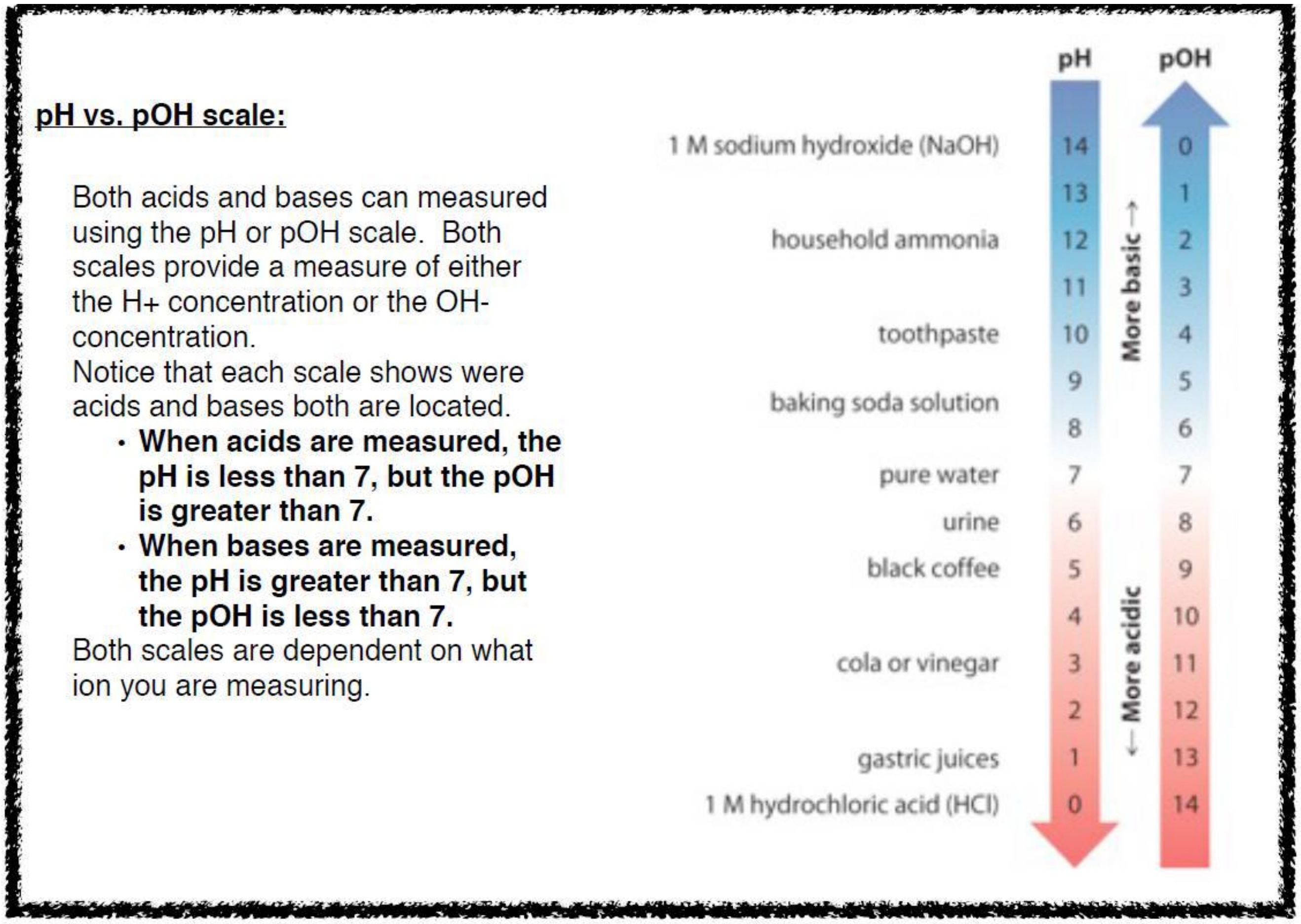
⇒ Relation between p[H^+] and p[OH^-] :



Acid solutions usually have protons and basic solutions have hydroxide ions. Concentrations of the ions are low (negative power of ten). pH scale is a convenient way of expressing these low concentrations in simple numbers between 1 and 14.

pH is the negative logarithm to the base ten of hydrogen ion concentration in moles per liter.

pH = -log[H^+]

p(OH) is the negative logarithm to the base ten of hydroxide ion concentration in moles per liter.

pOH = -log[OH^-]

In aqueous solutions, **pH + p(OH) = 14.**

pH scale is based on neutral water, where **[H+] = [OH–] = 10-7**

For a neutral solution **pH = – log [H+] = – log [10-7] = +7**

Finding p[OH] = - log[OH$^-$] , pH = 14 – pOH

Finding [OH] in bases :

pH = [H$^+$] [OH$^-$]

1x 10^{-14} = [H$^+$] [OH$^-$]

<u>**Example 1**</u> **\\ A solution of acetic acid (CH_3O_2H) has an H^+ concentration of 5x10^{-5} M. what is the pH of the solution?**
Sol. :
pH = -log[H$^+$]
pH = -log[5x10^{-5}]
= 4.3 , it is relatively weak acid .

<u>**Example 2**</u> **\\ find the pH of a 0.012 M sodium hydroxide (NaOH) solution.**
Sol. :
pOH = -log[OH$^-$]

= -log [0.012] = 1.92

pH + p(OH) = 14

pH = 14 – 1.92 = 12.08 , the solution has **pH of 12.08** and is a **strong base**.

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pH and pOH are related to one another; **THEY ARE <u>NOT</u> INDEPENDENT OF EACH OTHER**. As pH increases, pOH decreases. As pH decreases, pOH increases. By knowing what ion you are measuring on which scale, this will tell you whether or not the solution is acidic or basic.

Ion Concentration	Solution Type	pH	pOH
[H+] > [OH-]	**Acidic**	pH < 7	pOH > 7
[H+] < [OH-]	**Basic**	pH > 7	pOH < 7
[H+] = [OH-]	**Neutral**	pH = 7	pOH = 7

Because these scales are related, an equation can be used to explain their correlation.

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

$$pOH = 14 - pH$$

The bottom equations are manipulations of the top equation

The **<u>KEYS</u>** to calculations are knowing:

1. Knowing what type of solution you are working with
2. What equation to use first

Example Calculations:

1. Calculate the pH and pOH of a 0.33 M H_2SO_4 solution.

- Are you working with an acid or a base? ____________
- Are you given the concentration, or *Molarity* of that solution? ____________
- Next, plug the concentration or Molarity into the correct equation.

Because this is an acidic solution, we have to calculate pH first.

$$pH = -\log[0.33] = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

Now we can calculate the pOH, because we have calculated the pH.

$$pOH = 14 - pH = \_\_\_\_\_\_\_\_\_\_\_\_ = \_\_\_\_\_\_\_\_\_\_\_\_$$

2. Calculate the pH and pOH of a 0.25 M NaOH solution.

- Are you working with an acid or a base? ____________
- Are you given the concentration, or *Molarity* of that solution? ____________
- Next, plug the concentration or Molarity into the correct equation.

Because this is a basic solution, we have to calculate pOH first.

$$pOH = -\log[0.25] = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

Now we can calculate the pH, because we have calculated the pOH.

$$pH = 14 - pOH = \_\_\_\_\_\_\_\_\_\_\_\_ = \_\_\_\_\_\_\_\_\_\_\_\_$$

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Ex/ What is the pH of a 0.0005 M solution of NaOH at 25 °C ?

Solution / $NaOH \rightarrow Na^+ + OH^-$

$[OH^=] = 0.0005\ M = 5 \times 10^{-4}\ M$

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

$= -\log 5 \times 10^{-4}$

$= -\log 5 + 4 \log 10$

$= -0.699 + 4$

$= 3.301$

$pH = 14 - 3.401 = 10.7$

Ex/ What is the pH of a 0.1 M NH_3 solution ? K_b 1.8×10^{-5}

Solution / $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

0.1	0	0
0.1 – X	X	X

$K_b = [NH_4^+]\,[OH^-] / [NH_3]$

$1.8 \times 10^{-5} = (X)(X) / 0.1 - X$

$1.8 \times 10^{-5} = X^2 / 0.1$

$X^2 = 1.8 \times 10^{-6}$

$X = 1.34 \times 10^{-3} = [\,OH^-]$

$pOH = -\log [OH^-] = -\log 1.34 \times 10^{-3} = 2.87$

$pOH + pH = 14$

$pH \; 14 - 2.87 = 11.12$