



Medical chemistry- year1



Acid And Base

Lecture no(8)

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Medical chemistry- year1



Objectives

- 1- properties of acid and base
- 2- acid , base are defined by three main theories.
- 3- Arrhenius theory
- 4- Bronsted and Lowry Theory
- 5- Lewis Theory
- 6- Strong vs. Weak Acids and Bases
- 7- PH

History Of Acids And Bases

In the early days of chemistry chemists were organizing physical and chemical properties of substances. they different property categories:

ACID	BASE
1- sour taste	1- bitter taste
2- reacts with carbonates to make CO ₂	2- reacts with fats to make soaps
3-reacts with metals to produce H ₂	3-do not react with metals
4- turns blue litmus pink	4- turns red litmus blue
5- reacts with base to make salt water	5- reacts with acid to make salt water
6- having the pH <7	6- having the pH >7

Arrhenius was the first person to suggest a reason why substances are in A or B due their ionization in water

Arrhenius theory

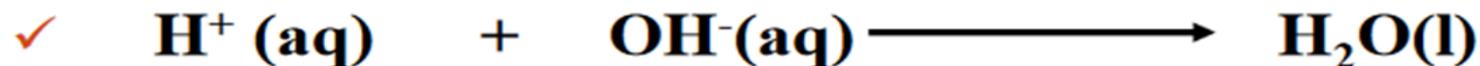
- ✓ Proposed in 1884 by Svante Arrhenius
- ✓ **ACIDS**: An acid is a substance that contains hydrogen and produces H^+ in aqueous Solution.



- ✓ **BASES**: A base is a solution that contains the OH group and produces Hydroxide ions OH^- , in aqueous solution.



- ✓ Neutralization is defined as the combination of H^+ ions with OH^- ions to form H_2O molecules.



limitation of Arrhenius theory

- 1- water is essential
- 2- not explain acidity or Basicity of no aqueous solvent
eg: benzene .
- 3- Basicity of ammonia (No OH^- ion) is not explain .
- 4- acidity of $\text{BF}_3, \text{AlCl}_3$ (NO H^- ion) is not explain

bronsted and lowry theory

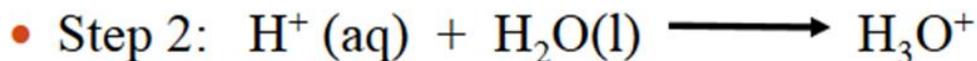
Proposed in 1923 by N. Bronsted and T.M.Lowry

Acid – any substance donating a proton, H^+

Base – any substance accepting a proton.

- Acids - H^+ donors
- Bases - H^+ acceptors

- An acid-base reaction is the transfer of a proton from an acid to a base.
- The complete ionization of hydrogen chloride, HCl in water is an acid-base reaction.



Lewis theory

- ✓ **Proposed by G. N. Lewis**
- ✓ An acid is any species that can accept a share in an electron pair. A base is any species that can make available or donate a share in an electron pair.
- ✓ **Lewis base - an electron pair donor**
- ✓ **Lewis acid - an electron pair acceptor**
- ✓ Lewis acid + Lewis base \longleftrightarrow Adduct

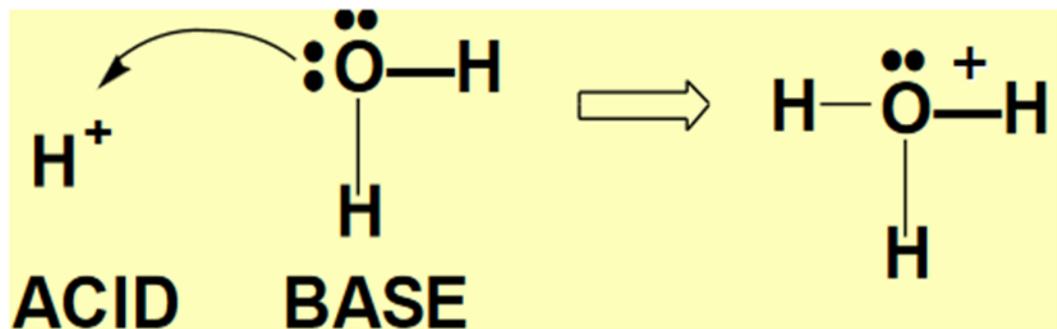
- **Bases - Electron Pair Donors**

species with O,N, halogen frequently have lone pairs of electrons to share ...
Lewis bases

Lewis theory

A Lewis acid and base can interact by sharing an electron pair.

Formation of **hydronium ion** is an excellent example.



Strong vs. Weak Acids and Bases

Acid and base strength is based on the extent of substance is dissolved . ionization that occurs when the in water

Strong Acids:

- strong electrolytes - completely ionized in solution.
- there are (6) strong acids – know them: HCl, HBr, HI, HNO₃, (diprotic).HClO₄, H₂SO₄

Weak Acids

- weak electrolytes - partially ionized (typically < 5%) in aqueous solution.
- any acid that is not a strong acid is a weak acid some examples: HF, H₂CO₃, H₃PO₄, HNO₂, HBrO₄

Strong vs. Weak Acids and Bases

Strong Bases:

- strong electrolytes - completely ionized in solution.
- the strong bases are the hydroxides of the alkali metals & hydroxides of most alkaline earth metals; KNOW THEM:
LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂.

Weak Bases:

- weak electrolytes - partially ionized (typically < 5%) in aqueous solution.
- weak bases: tend to be organic compounds that contain nitrogen; ammonia and substituted amines some examples:
NH₃, (CH₃)NH₂, (CH₃)₃N, C₅H₅N, N₂H₄, NH₂OH.

Strong vs. Weak Acids and Bases

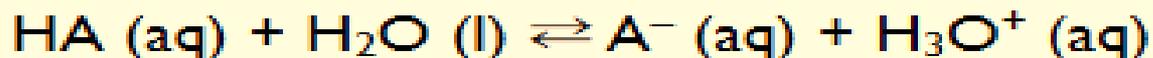
Reversible H⁺ Transfer Reactions

- we defined weak acids and weak bases as weak electrolytes (only partially ionized in aqueous solution).
- Now we can talk about their behaviour in terms of an equilibrium that exists in solution:



- These are heterogeneous equilibria
- We will discuss/define equilibrium constants, K_a & K_b .

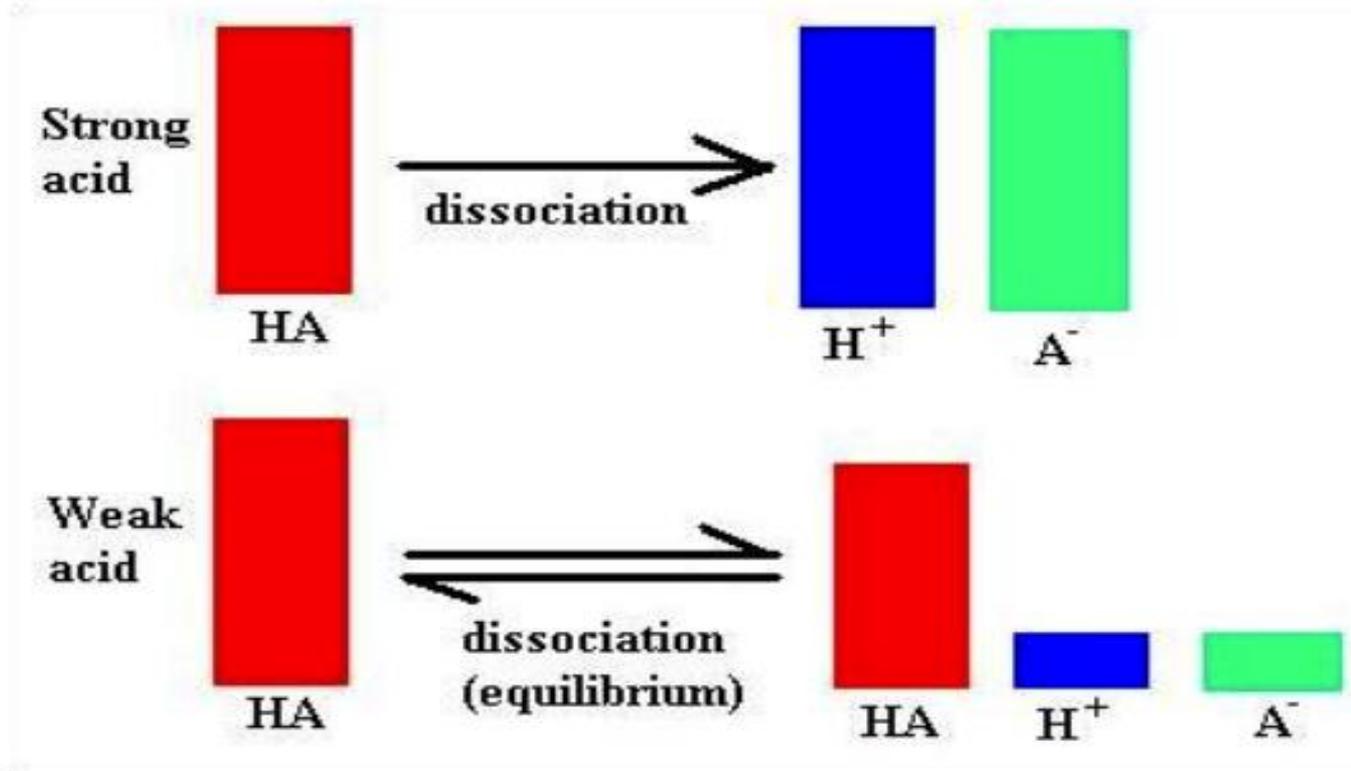
Weak Acids and Acid Ionization Constant, K_a



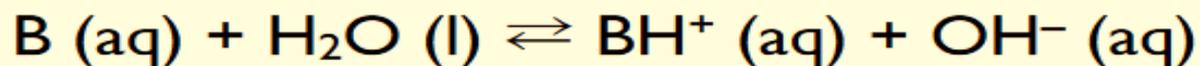
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

- ◆ K_a is the acid ionization constant
- ◆ the larger the value of K_a ...
 - the equilibrium position lies farther to the right
 - higher $[\text{H}_3\text{O}^+]$
 - greater extent of ionization
 - stronger acid***

Weak Acids and Acid Ionization Constant, K_a



Weak Bases and Bases Ionization Constant, K_b



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

- ◆ K_b is the base ionization constant
- ◆ the larger the value of K_b ...
 - the equilibrium position lies farther to the right
 - higher $[OH^-]$
 - greater extent of ionization
 - stronger base***

Acid- Ionization Constant at 25° C

Table 16.1 Acid-Ionization Constants at 25°C*

Substance	Formula	K_a
Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$	1.7×10^{-5}
Benzoic acid	$\text{HC}_7\text{H}_5\text{O}_2$	6.3×10^{-5}
Boric acid	H_3BO_3	5.9×10^{-10}
Carbonic acid	H_2CO_3	4.3×10^{-7}
	HCO_3^-	4.8×10^{-11}
Cyanic acid	HOCN	3.5×10^{-4}
Formic acid	HCHO_2	1.7×10^{-4}
Hydrocyanic acid	HCN	4.9×10^{-10}
Hydrofluoric acid	HF	6.8×10^{-4}
Hydrogen sulfate ion	HSO_4^-	1.1×10^{-2}
Hydrogen sulfide	H_2S	8.9×10^{-8}
	HS^-	$1.2 \times 10^{-13}\ddagger$

Bases- Ionization Constant at 25° C

Table 16.2 Base-Ionization Constants at 25°C

Substance	Formula	K_b
Ammonia	NH_3	1.8×10^{-5}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2$	4.2×10^{-10}
Dimethylamine	$(\text{CH}_3)_2\text{NH}$	5.1×10^{-4}
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2$	4.7×10^{-4}
Hydrazine	N_2H_4	1.7×10^{-6}
Hydroxylamine	NH_2OH	1.1×10^{-8}
Methylamine	CH_3NH_2	4.4×10^{-4}
Pyridine	$\text{C}_5\text{H}_5\text{N}$	1.4×10^{-9}
Urea	NH_2CONH_2	1.5×10^{-14}

Acidic, Basic & Neutral Aqueous Solutions

- distinguish between acidic, basic and neutral solutions based on the relative $[H_3O^+]$ & $[OH^-]$

if $[H_3O^+] > [OH^-]$, solution is acidic

if $[OH^-] > [H_3O^+]$, solution is basic

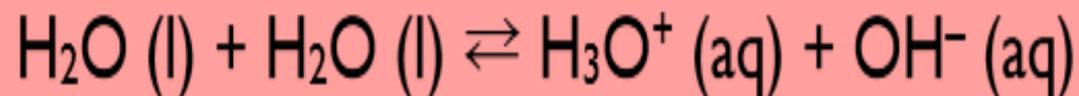
if $[H_3O^+] = [OH^-]$, solution is neutral

- ♦ for a neutral solution at 25°C:

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$$

Auto- Ionization of water and KW

- recall that water is amphoteric - can act as an acid or a base.
- now consider a reaction between 2 water molecules:



- ♦ this is called the auto-ionization of water heterogeneous equilibrium

$$K_{\text{W}} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$\text{at } 25^\circ\text{C}, K_{\text{W}} = 1.0 \times 10^{-14}$$

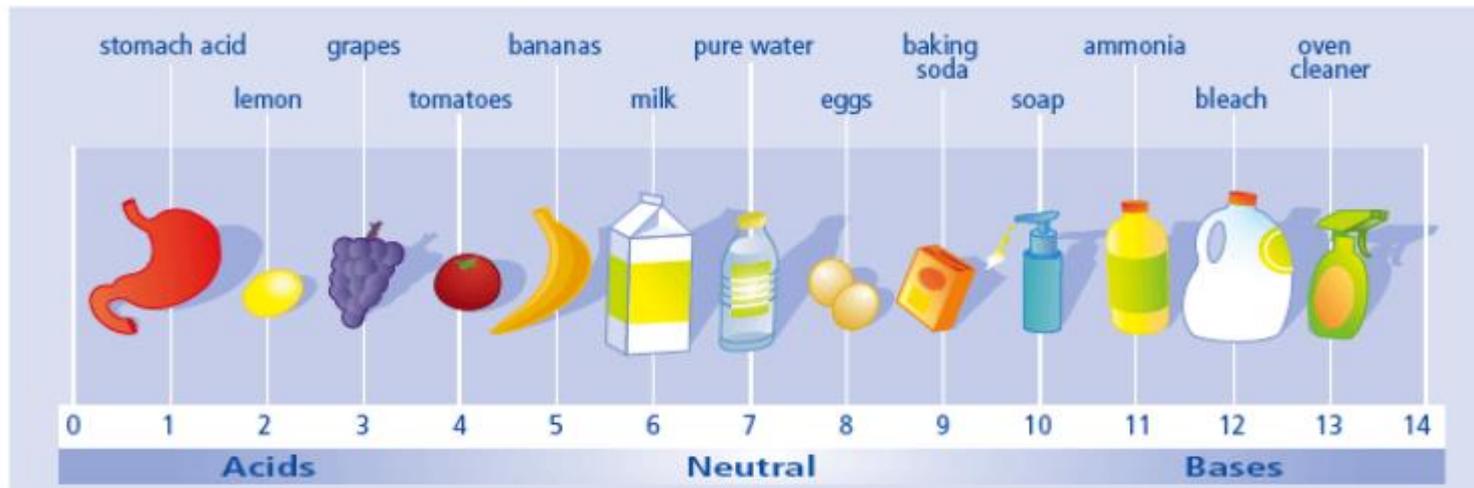
In any aqueous solution at 25°C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_{\text{W}} = 1.0 \times 10^{-14}$$

The PH Scale

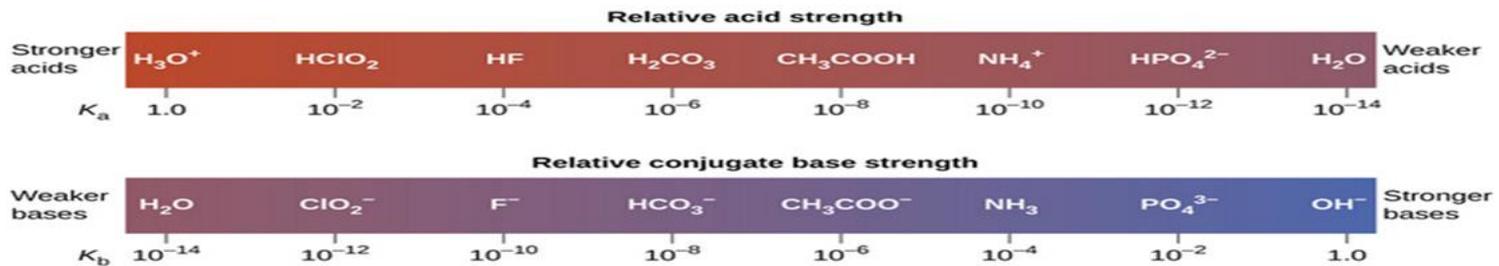
logarithmic scale of $[H_3O^+]$ in solution

$$pH = -\log[H_3O^+]; \quad [H_3O^+] = 10^{-pH}$$



PH

- pH is a unit of measure which describes the degree of acidity or alkalinity (basic) of a solution.
- It is measured on a scale of 0 to 14.
- The formal definition of pH is the negative logarithm of the hydrogen ion activity.
- $\text{pH} = -\log[\text{H}^+]$



pH value

- The pH value of a substance is directly related to the ratio of the hydrogen ion and hydroxyl ion concentrations.
- If the H^+ concentration is higher than OH^- the material is acidic.
- If the OH^- concentration is higher than H^+ the material is basic.
- 7 is neutral, $<$ is acidic, >7 is basic

PH calculation

Relative acidity and basicity of solution

- ♦ in any aqueous solution at 25°C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$$

- ♦ $\text{pH} = -\log[\text{H}_3\text{O}^+]$; $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$

- ♦ higher $[\text{H}^+]$ \Rightarrow more acidic solution \Rightarrow
lower pH

- ♦ higher $[\text{OH}^-]$ \Rightarrow lower $[\text{H}^+]$ \Rightarrow
more basic solution \Rightarrow higher pH

POH calculation

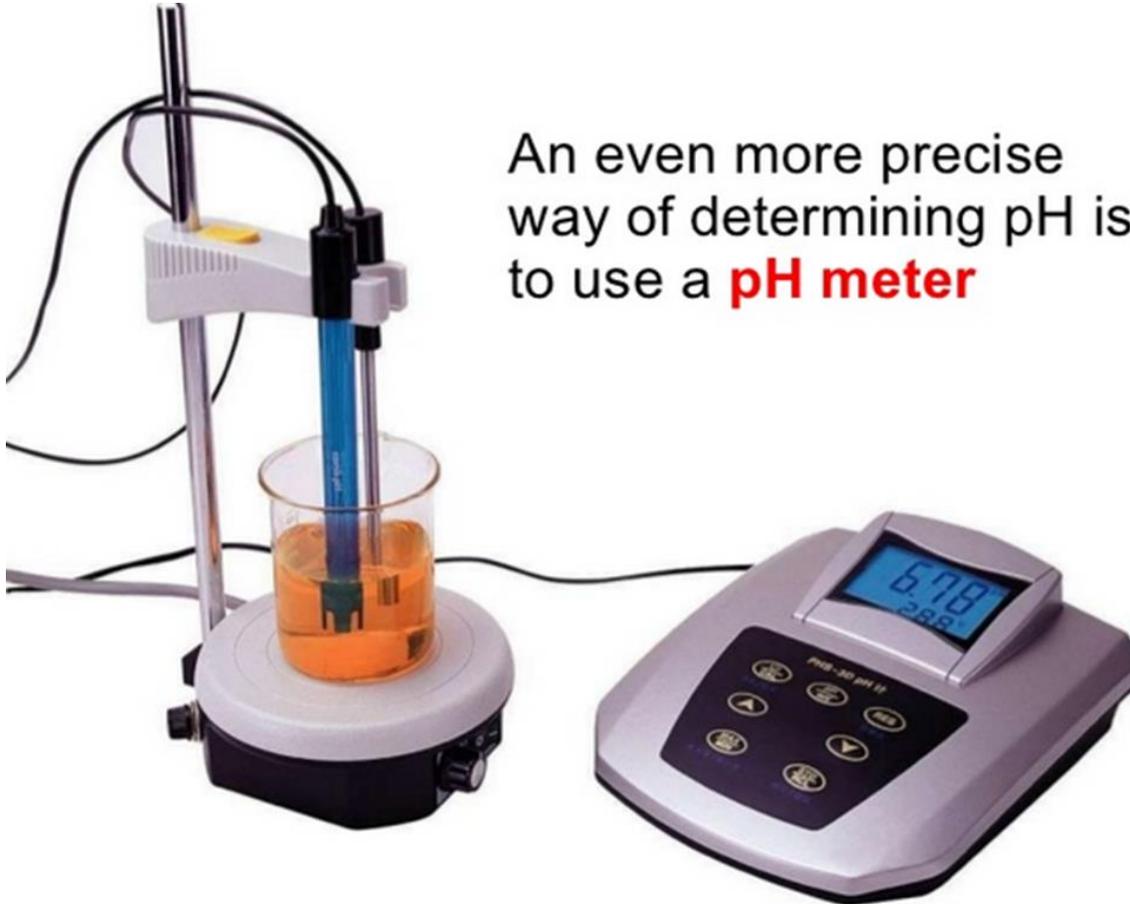
- ♦ $\text{pOH} = -\log [\text{OH}^-]$ $[\text{OH}^-] = 10^{-\text{pOH}}$
the higher the $[\text{OH}^-]$, the lower the pOH
as $[\text{OH}^-]$ changes by factor of 10, the pOH
changes by 1 unit

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

at 25 °C

PH meter

An even more precise way of determining pH is to use a **pH meter**



Example 1: What is the pH of the strong base 0.1M NaOH Assume complete dissociation ?

SOL:



$$K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1 \times 10^{-14}$$

$$[\text{H}^+(\text{aq})] = k_w / [\text{OH}^-(\text{aq})] = 1 \times 10^{-14} / 0.1 = 1 \times 10^{-13} \text{ M}$$

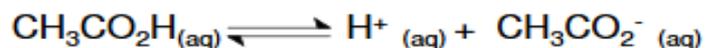
$$\text{pH} = -\log[1 \times 10^{-13}] = 13.00$$

Example 2:

What is the concentration of HCl with a pH of 1.35?

$$[\text{H}^+] = 1 \times 10^{-1.35} = 0.045\text{M}$$

Example 3 What is the pH of a solution of 0.01M ethanoic acid (ka is $1.7 \times 10^{-5} \text{ mol cm}^{-3}$)?



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{CH}_3\text{CO}_2^-(\text{aq})]}{[\text{CH}_3\text{CO}_2\text{H}_{(\text{aq})}]} \rightarrow K_a = \frac{[\text{H}^+(\text{aq})]^2}{[\text{CH}_3\text{CO}_2\text{H}_{(\text{aq})}]} \rightarrow 1.7 \times 10^{-5} = \frac{[\text{H}^+(\text{aq})]^2}{0.01}$$

$$[\text{H}^+(\text{aq})]^2 = 1.7 \times 10^{-5} \times 0.01$$

$$[\text{H}^+(\text{aq})] = \sqrt{1.7 \times 10^{-7}} = 4.12 \times 10^{-4}$$

$$\text{pH} = -\log [\text{H}^+] = -\log (4.12 \times 10^{-4})$$

$$\text{pH} = 3.38$$

Ex:

Calculate pH for 0.060M NaOH:

NaOH is strong base and completely ionizes. Therefore, $[\text{OH}^{-1}] = 0.060\text{M}$

$$[\text{H}^{+1}][\text{OH}^{-1}] = 1.0 \times 10^{-14}$$

$$[\text{H}^{+1}][0.060] = 1.0 \times 10^{-14}$$

$$[\text{H}^{+1}] = 1.7 \times 10^{-13}$$

$$\text{pH} = -\log[\text{H}^{+1}] = -\log[1.7 \times 10^{-13}] = 12.77$$

Calculate pH for 0.00025M $\text{Ca}(\text{OH})_2$:

$\text{Ca}(\text{OH})_2$ strong base and completely ionizes. Therefore, $[\text{OH}^{-1}] = 0.00050\text{M}$

$$[\text{H}^{+1}][\text{OH}^{-1}] = 1.0 \times 10^{-14}$$

$$[\text{H}^{+1}][0.00050] = 1.0 \times 10^{-14}$$

$$[\text{H}^{+1}] = 2.0 \times 10^{-11}$$

$$\text{pH} = -\log[\text{H}^{+1}] = -\log[2.0 \times 10^{-11}] = 10.70$$

Calculate $[\text{OH}^{-1}]$ when $\text{pH} = 11.75$

$$\text{pOH} = 14 - 11.75 = 2.25$$

$$[\text{OH}^{-1}] = \text{antilog}(-\text{pOH}) = \text{antilog}(-2.25) = 0.00562\text{M}$$

Thank you