pH for the acidic solution

In order to find the numeric value of the level of acidity or basicity of a substance, the pH scale (where in pH stands for ‘potential of hydrogen’) can be used.

The pH scale is the most common and trusted way to measure how acidic or basic a substance. A pH scale measure can vary from 0 to 14, where 0 is the most acidic and 14 is the most basic a substance can be.

Another way to check if a substance is acidic or basic is to use litmus paper. There are two types of litmus paper available that can be used to identify acids and bases – red litmus paper and blue litmus paper. Blue litmus paper turns red under acidic conditions and red litmus paper turns blue under basic or alkaline conditions.

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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 from7 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:

$$\mathbf{pH} = -\log\,[H^+]$$

$$\mathbf{pH} = -\log\,[OH^-]$$

Both pH and pOH are related to each other. pH is inversely proportional to pOH; pH increases with decreasing pOH.

Relation between p $[H^+]$

$$\mathbf{PH} \propto \frac{1}{POH}$$

and p [〖OH〗 :

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

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pH – A Measure of Acidity

$$\boxed{pH = -\log [H^+]}$$

Solution Is		**At 25⁰C**	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1.0 \times 10^{-7}$	$pH = 7$
acidic	$[H^+] > [OH^-]$	$[H^+] > 1.0 \times 10^{-7}$	$pH < 7$
basic	$[H^+] < [OH^-]$	$[H^+] < 1.0 \times 10^{-7}$	$pH > 7$

$pH \uparrow$ $\quad [H^+] \downarrow$

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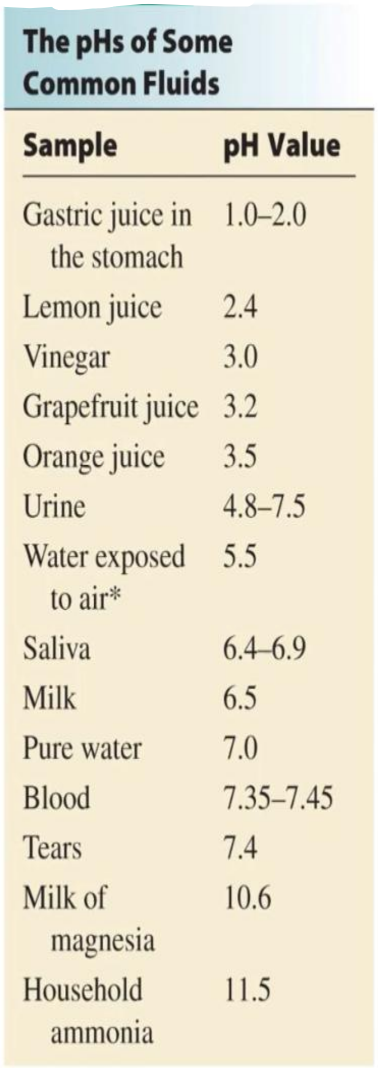

The pHs of Some Common Fluids

Sample	pH Value
Gastric juice in the stomach	1.0–2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Urine	4.8–7.5
Water exposed to air*	5.5
Saliva	6.4–6.9
Milk	6.5
Pure water	7.0
Blood	7.35–7.45
Tears	7.4
Milk of magnesia	10.6
Household ammonia	11.5

Other important relationships

$$\mathbf{pOH = -\log [OH^-]}$$

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

$$-\log [H^+] - \log [OH^-] = 14.00$$

$$\mathbf{pH + pOH = 14.00}$$

pH Meter

Example :

The concentration of H^+ ions in a solution is **3.2×10^{-4} *M*.** The solution was left for a while, it was found that the hydrogen ion concentration equal to **1.0×10^{-3}** *M*. **Calculate the pH** of the solution on these two occasions.

$pH = -\log [H^+]$ $\quad\quad$ $[H^+] = 3.2 \times 10^{-4}\ M$

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

$= -\log (3.2 \times 10^{-4}) = \mathbf{3.49}$

On the second occasion, $[H^+] = 1.0 \times 10^{-3}\ M$,

so that

$pH = -\log (1.0 \times 10^{-3}) = \mathbf{3.00}$

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Example :

In a NaOH solution $[OH^-]$ is 2.9×10^{-4} *M*. Calculate the pH of the solution.

$$\mathbf{pOH = -\log [OH^-]}$$

$$= -\log (2.9 \times 10^{-4})$$

$$= 3.54$$

$$\mathbf{pH + pOH = 14.00}$$

$$pH = 14.00 - pOH$$

$$= 14.00 - 3.54 = \mathbf{10.46}$$

Alternatively, we can use $K_w = [H^+][OH^-]$ to calculate $[H^+]$, and then we can calculate the pH from the $[H^+]$. *Try it*.

$$[H^+] = \frac{K_w}{[OH^-]} \Longrightarrow pH = -\log [H^+]$$

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Strength of acids and bases

Strong Electrolyte – 100% dissociation

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

Weak Electrolyte – not completely dissociated

$$CH_3COOH \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$$

Strong Acids are strong electrolytes

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

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

$$HClO_4\ (aq) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + ClO_4^-\ (aq)$$

$$H_2SO_4\ (aq) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + HSO_4^-\ (aq)$$

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Weak Acids are weak electrolytes

$$HF\ (aq) + H_2O\ (l) \boxed{\rightleftharpoons} H_3O^+\ (aq) + F^-\ (aq)$$

$$HNO_2\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + NO_2^-\ (aq)$$

$$HSO_4^-\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + SO_4^{2-}\ (aq)$$

$$H_2O\ (l) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$$

Strong Bases are strong electrolytes

$$NaOH\ (s) \boxed{\xrightarrow{H_2O}} Na^+\ (aq) + OH^-\ (aq)$$

$$KOH\ (s) \xrightarrow{H_2O} K^+\ (aq) + OH^-\ (aq)$$

$$Ba(OH)_2\ (s) \xrightarrow{H_2O} Ba^{2+}\ (aq) + 2OH^-\ (aq)$$

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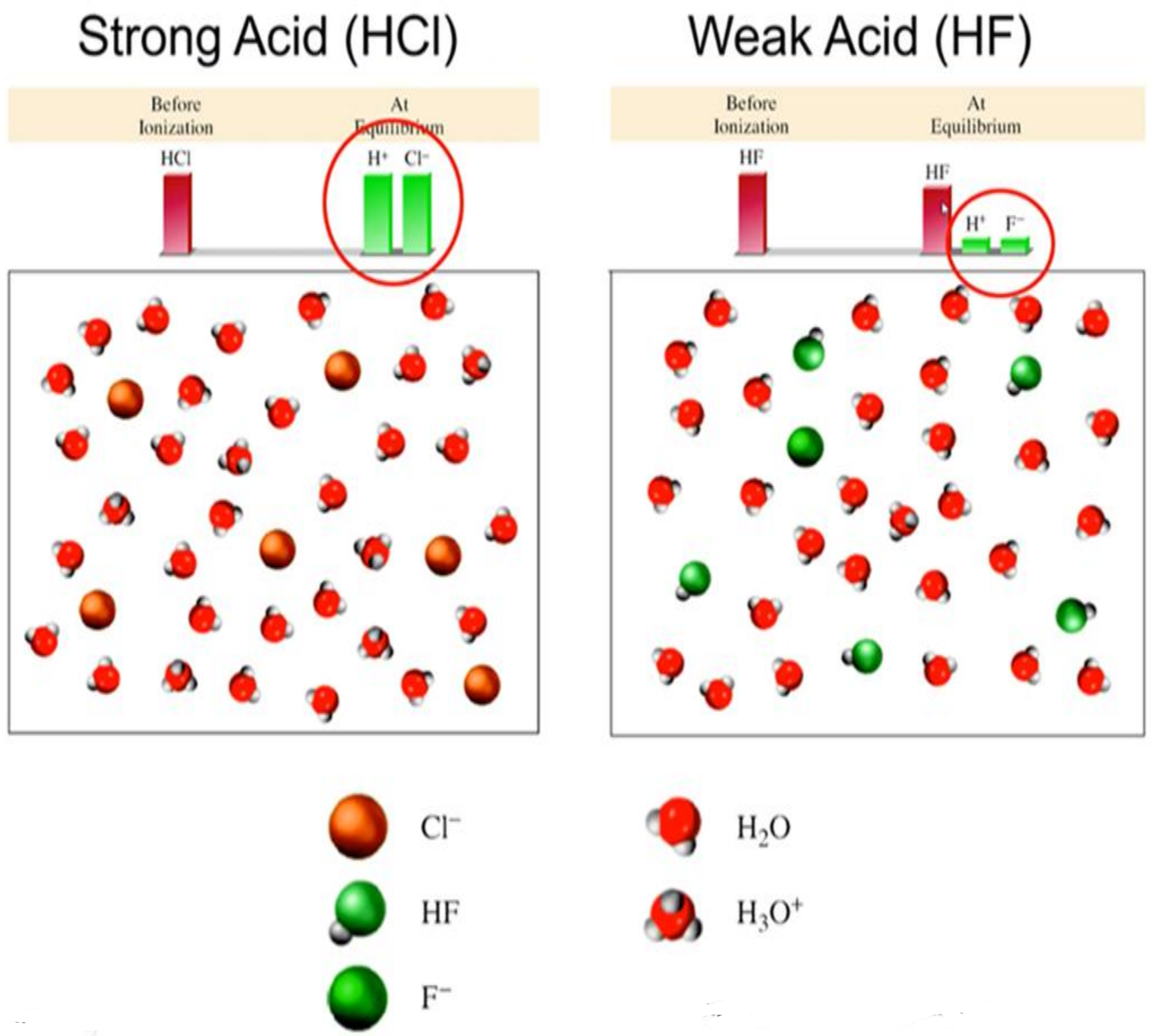

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Weak Bases are weak electrolytes

$$F^- (aq) + H_2O (l) \rightleftharpoons OH^- (aq) + HF (aq)$$

$$NO_2^- (aq) + H_2O (l) \rightleftharpoons OH^- (aq) + HNO_2 (aq)$$

Conjugate acid-base pairs:

- The conjugate base of a strong acid has no measurable strength.
- H_3O^+ is the strongest acid that can exist in aqueous solution.
- The OH^- ion is the strongest base that can exist in aqueous solution.

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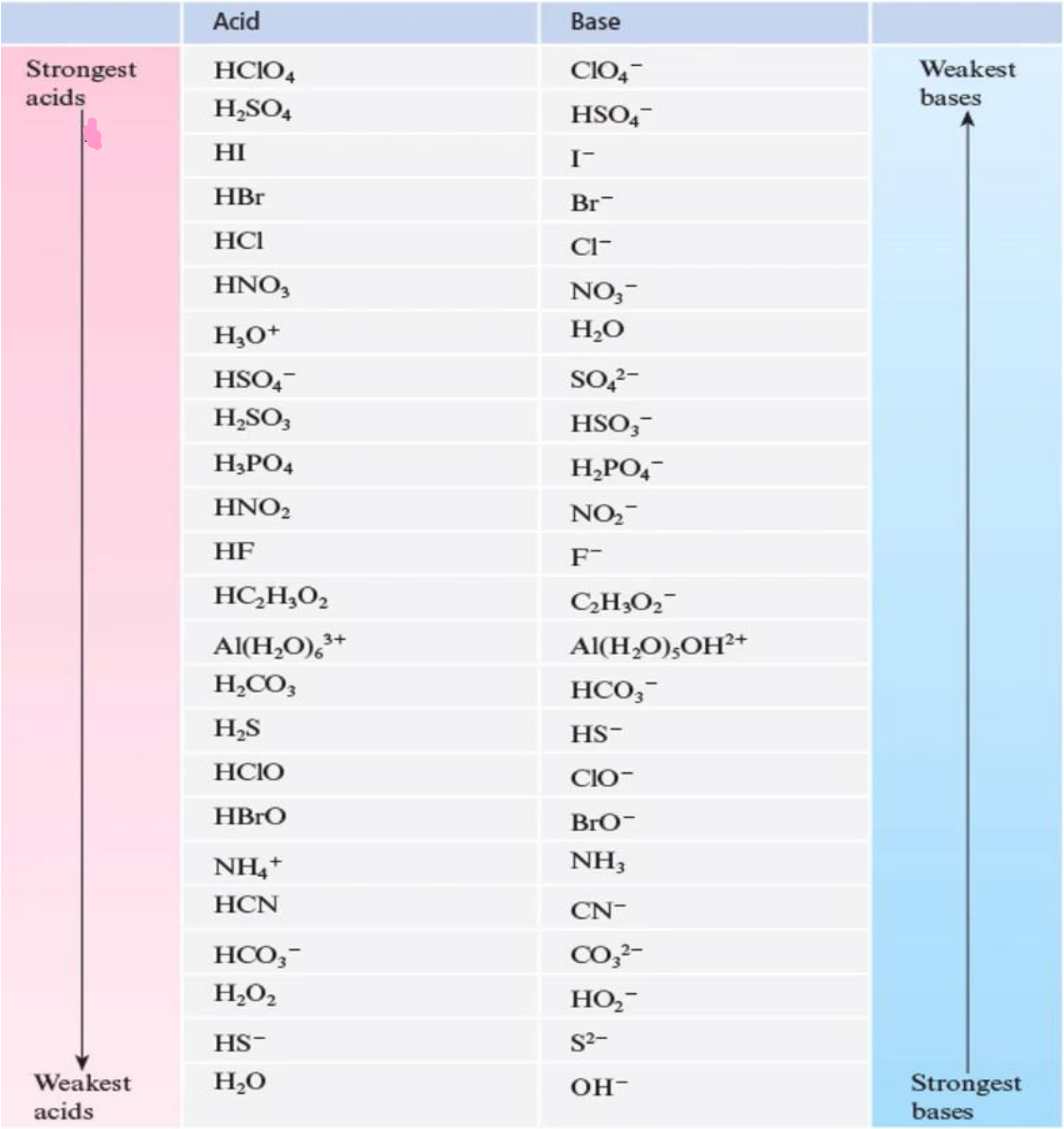
Relative strength of acids and bases

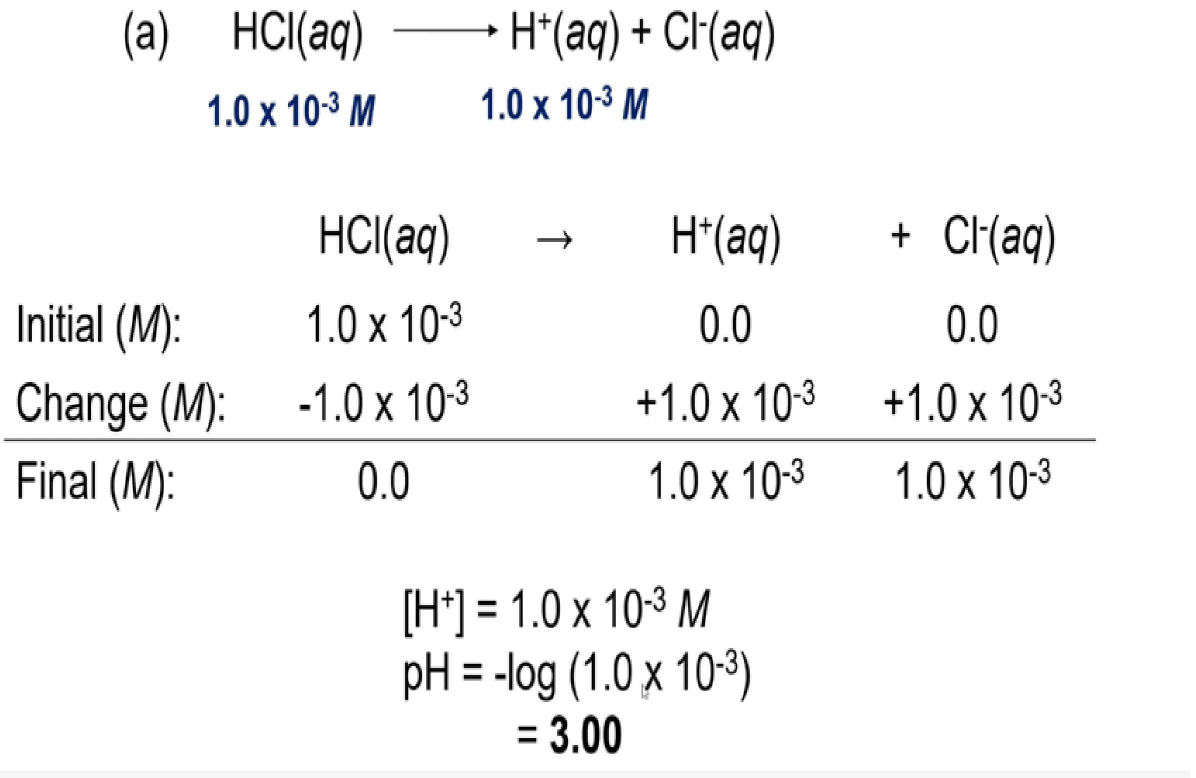


	Acid	Base	
Strongest acids	$HClO_4$	ClO_4^-	Weakest bases
	H_2SO_4	HSO_4^-	
	HI	I^-	
	HBr	Br^-	
	HCl	Cl^-	
	HNO_3	NO_3^-	
	H_3O^+	H_2O	
	HSO_4^-	SO_4^{2-}	
	H_2SO_3	HSO_3^-	
	H_3PO_4	$H_2PO_4^-$	
	HNO_2	NO_2^-	
	HF	F^-	
	$HC_2H_3O_2$	$C_2H_3O_2^-$	
	$Al(H_2O)_6^{3+}$	$Al(H_2O)_5OH^{2+}$	
	H_2CO_3	HCO_3^-	
	H_2S	HS^-	
	$HClO$	ClO^-	
	$HBrO$	BrO^-	
	NH_4^+	NH_3	
	HCN	CN^-	
	HCO_3^-	CO_3^{2-}	
	H_2O_2	HO_2^-	
	HS^-	S^{2-}	
Weakest acids	H_2O	OH^-	Strongest bases

Example :

Calculate the pH of a :

(a) 1.0×10^{-3} *M* **HCl** solution (b) 0.020 *M* **$Ba(OH)_2$** solution



(a) $HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$

1.0×10^{-3} *M* **1.0×10^{-3} *M***

	$HCl(aq)$	→	$H^+(aq)$	+ $Cl^-(aq)$
Initial (*M*):	1.0×10^{-3}		0.0	0.0
Change (*M*):	-1.0×10^{-3}		$+1.0 \times 10^{-3}$	$+1.0 \times 10^{-3}$
Final (*M*):	0.0		1.0×10^{-3}	1.0×10^{-3}

$[H^+] = 1.0 \times 10^{-3}$ *M*
$pH = -\log(1.0 \times 10^{-3})$
$= \mathbf{3.00}$

(b) $Ba(OH)_2$ is a strong base

$$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$$

0.020 *M* **2 X 0.020 *M* = *0.040 M***

$$[OH^-] = 0.040\ M$$

$$pOH = -\log 0.040 = 1.40$$

pH = 14.00 - pOH

= 14.00 - 1.40

= **12.60**

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