ACIDS AND BASES – 2 IONIZATION OF WATER, pH

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IONIZATION OF WATER

Self-Ionization of Water

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

Base Acid Conj. Acid Conj. base

We call the equilibrium constant the ion-product constant, K_w.

$$K_w = [H_3O^+][OH^-] \text{ At } 25°C, K_w = 1.0 \times 10^{-14}$$

As temperature increases, the value of K_w increases.

Solutions can be characterized as

Acidic: $[H_3O^+] > 1.0 \times 10^{-7}\ M$

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

Basic: $[H_3O^+] < 1.0 \times 10^{-7}\ M$

EXAMPLE: CALCULATION OF CONCENTRATIONS

Calculate the hydronium and hydroxide ion concentration at 25°C in

a. 0.10 *M* HCl

b. 1.4×10^{-4} *M* $Mg(OH)_2$

Solution:

a. When HCl ionizes, it gives H^+ and Cl^-.
So $[H^+] = [Cl^-] = [HCl] = 0.10$ M.

b. When $Mg(OH)_2$ ionizes, it gives Mg^{2+} and 2 OH^-.
So $2[OH^-] = 2(1.4 \times 10^{-4}\ M) = 2.8 \times 10^{-4}$ M.

CALCULATING pH

$pH = -\log[H_3O^+]$

And inverse: $[H_3O^+] = 10^{-pH}$

- When pH = 7, the solution is neutral ($[H_3O^+] = 1.0 \times 10^{-7}$ M).
- When pH < 7.00, the solution is acidic ($[H_3O^+] > 1.0 \times 10^{-7}$ M).
- When pH > 7.00, the solution is basic ($[H_3O^+] < 1.0 \times 10^{-7}$ M).

TABLE 16.3 Benchmark pH Values for a Range of Hydronium Ion Concentrations at 25°C

$[H_3O^+]$ (M)	$-\log[H_3O^+]$	pH	
0.10	$-\log(1.0 \times 10^{-1})$	1.00	
0.010	$-\log(1.0 \times 10^{-2})$	2.00	
1.0×10^{-3}	$-\log(1.0 \times 10^{-3})$	3.00	
1.0×10^{-4}	$-\log(1.0 \times 10^{-4})$	4.00	
1.0×10^{-5}	$-\log(1.0 \times 10^{-5})$	5.00	
1.0×10^{-6}	$-\log(1.0 \times 10^{-6})$	6.00	Acidic
1.0×10^{-7}	$-\log(1.0 \times 10^{-7})$	7.00	Neutral
1.0×10^{-8}	$-\log(1.0 \times 10^{-8})$	8.00	Basic
1.0×10^{-9}	$-\log(1.0 \times 10^{-9})$	9.00	
1.0×10^{-10}	$-\log(1.0 \times 10^{-10})$	10.00	
1.0×10^{-11}	$-\log(1.0 \times 10^{-11})$	11.00	
1.0×10^{-12}	$-\log(1.0 \times 10^{-12})$	12.00	
1.0×10^{-13}	$-\log(1.0 \times 10^{-13})$	13.00	
1.0×10^{-14}	$-\log(1.0 \times 10^{-14})$	14.00	

TABLE 16.4 pH Values of Some Common Fluids

Fluid	pH	Fluid	pH
Stomach acid	1.0	Saliva	6.4–6.9
Lemon juice	2.0	Milk	6.5
Vinegar	3.0	Pure water	7.0
Grapefruit juice	3.2	Blood	7.35–7.45
Orange juice	3.5	Tears	7.4
Urine	4.8–7.5	Milk of magnesia	10.6
Rainwater (in clean air)	5.5	Household ammonia	11.5

EXAMPLE: CALCULATING pH

1) What is the pH of a solution that has a hydronium ion concentration of 6.5 x $10^{-5}M$?

$$pH = -\log[H_3O^+]$$

$$pH = -\log[6.54 \times 10^{-5}]$$

$$pH = 4.19$$

2) What is the hydronium ion concentration of a solution with pH 3.65?

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

$$[H_3O^+] = 10^{-3.65}$$

$$[H_3O^+] = 2.2x10^{-4}$$

pH ,pOH AND K_W

- pOH can be calculated just like pH.

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

$$[OH^-] = 10^{-pOH}$$

- K_W is the hydronium ion and hydroxide ion concentrations; so pK_W is the addition of pH and pOH.

$$K_w = [H_3O^+][OH^-]$$

$$pK_w = pH + pOH = 14.00$$

- Next slide has all relationships…

$$
\begin{tikzcd}[column sep=huge, row sep=huge]
{[H_3O^+]} \arrow[r, shift left, "{[OH^-]=\frac{1\times10^{-14}}{[H_3O^+]}}"] \arrow[d, shift left, "{pH=-\log[H_3O^+]}"] & {[OH^-]} \arrow[l, shift left, "{[H_3O^+]=\frac{1\times10^{-14}}{[OH^-]}}"] \arrow[d, shift left, "{pOH=-\log[OH^-]}"] \\
pH \arrow[u, shift left, "{[H_3O^+]=10^{-pH}}"] \arrow[r, shift left, "{pOH=14-pH}"] & pOH \arrow[l, shift left, "{pH=14-pOH}"] \arrow[u, shift left, "{[OH-]=10^{-pOH}}"]
\end{tikzcd}
$$

EXAMPLE - pOH

1) What is the pOH of a solution that has a hydroxide ion concentration of $4.3 \times 10^{-2} M$?

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

$$pOH = -\log[4.3 \times 10^{-2}]$$

$$pOH = 1.37$$

2) What is the hydroxide ion concentration of a solution with pOH 8.35?

$$[OH^-] = 10^{-pOH}$$

$$[OH^-] = 10^{-8.35}$$

$$[OH^-] = 4.5 x 10^{-9}$$

STRONG ACIDS AND BASES

- All strong acids and bases ionize completely; so their concentrations can be used to calculate the pH and pOH.
- For diprotic acids, e.g. sulfuric acid only the first ionization reactions is measured; on the contrary dihydroxide bases will ionize to give two hydroxides in solution.

Strong Acid	Ionization Reaction
Hydrochloric acid	$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$
Hydrobromic acid	$HBr(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Br^-(aq)$
Hydroiodic acid	$HI(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + I^-(aq)$
Nitric acid	$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$
Chloric acid	$HClO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + ClO_3^-(aq)$
Perchloric acid	$HClO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + ClO_4^-(aq)$
Sulfuric acid	$H_2SO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + HSO_4^-(aq)$

Group 1A hydroxides	
$LiOH(aq)$	$\longrightarrow Li^+(aq) + OH^-(aq)$
$NaOH(aq)$	$\longrightarrow Na^+(aq) + OH^-(aq)$
$KOH(aq)$	$\longrightarrow K^+(aq) + OH^-(aq)$
$RbOH(aq)$	$\longrightarrow Rb^+(aq) + OH^-(aq)$
$CsOH(aq)$	$\longrightarrow Cs^+(aq) + OH^-(aq)$

Group 2A hydroxides	
$Ca(OH)_2(aq)$	$\longrightarrow Ca^{2+}(aq) + 2OH^-(aq)$
$Sr(OH)_2(aq)$	$\longrightarrow Sr^{2+}(aq) + 2OH^-(aq)$
$Ba(OH)_2(aq)$	$\longrightarrow Ba^{2+}(aq) + 2OH^-(aq)$

EXAMPLE

- What is the pH of a 0.057 M solution of HBr?

Solution: HBr ionizes completely so the conc. of the acid can be used as H_3O^+.

$$HBr\,(aq) + H_2O\,(l) \longrightarrow H_3O^+(aq) + Br^-(aq)$$

$$pH = -\log[H_3O^+]$$

$$pH = -\log[0.057]$$

$$pH = 1.24$$ *Note: very low pH – strong acid.*

- What is the pOH of a solution of 0.034 *M* solution of $Ca(OH)_2$?

Solution: $Ca(OH)_2$ ionizes to give 2 mols of OH^-, so conc. has to be doubled.

$$Ca(OH)_2\,(aq) \longrightarrow Ca^{2+}(aq) + 2OH^-\,(aq)$$

$$[OH^-] = 0.034\ M\ Ca(OH)_2 x \frac{2\ mol\ OH^-}{1\ mol\ Ca(OH)_2}$$

$$[OH^-] = 0.068\ M$$

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

$$pOH = -\log[0.068]$$

$$pOH = 1.17$$

KEY CONCEPTS

- Calculation of pH and pOH
- Calculation of hydronium and hydroxide ion conc.