Unit 10 Acids and Bases

pH is the log of the hydrogen ion concentration; it is a measure of the strength of acids and bases: $pH = -\log [H^+]$

acidic pH< 7 (1-7)	neutral = pH 7.0	basic pH>7 (alkaline) (7-14)

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

$2H_2O$	$\rightleftharpoons$ H_3O^+ hydronium ion (acid, proton donor)	+	OH^- hydroxide ion (base, proton acceptor)

pH is a ten times scale

pH 3 = 100X less acidic than pH 1
pH 3 = 10X less acidic than pH 2

pH 3 = 10X more acidic than pH 4
pH 3 = 100X more acidic than pH 5
pH 3 = 1000X more acidic than pH 6

Buffer - weak acid or base that minimizes changes in pH, e.g. carbonic acid:

H_2CO_3	$\rightleftharpoons$	HCO_3^-	+	H^+
H^+ donor (acid)		H^+ acceptor (base)		

buffers donate H^+ when the solution is basic and accept H^+ when the solution is acidic.

Ionization constant for pure water

$K_W = [H^+]\,[OH^-] = (1\times10^{-7}\,M)(1\times10^{-7}\,M) = 1\times10^{-14}\,M$

Therefore, pH + pOH = 14

Monoprotic acid- contain 1 ionizable hydrogen, e.g. HNO_3 , nitric acid.

Diprotic acid- contain 2 ionizable hydrogens, e.g. H_2SO_4 , sulfuric acid.
Triprotic acid- contain 3 ionizable hydrogens, e.g. H_3PO_4 , phosphoric acid.
Amphoteric- substance that can act as both an acid and a base, e.g. water.

Acid -Base Theories
Svante Arrhenius- acids yield H^+ ions and bases yield OH^- ions.

Johannes Brønsted - Thomas Lowry - acids are proton donors and bases are proton acceptors.

$$NH_3\,(aq) + H_2O\,(l) \rightarrow NH_4^+\,(aq) + OH^-\,(aq)$$

ammonia	water	ammonium	hydroxide
base; hydrogen ion acceptor	*acid; hydrogen ion donor*	ion	ion

Gilbert Lewis- Acids accept electrons and bases donate electrons to form a covalent bond.
Acids accept e^-/donate p^+ & bases donate e^-/accept p^+

Concentration vs strength
Concentration is the amount of a substance dissolved in solution, while strength refers to how much is ionized and, therefore, able to react. Strong acids and bases ionize completely whereas weak acids and bases ionize only slightly.

Strong acids
HCl, HNO_3, H_2SO_4 hydrochloric, nitric, sulfuric acids

Weak acids
CH_3COOH, H_2CO_3, H_3BO_3 ethanoic, carbonic, boric acids

Strong bases
KOH, $NaOH$, $Ca(OH)_2$ potassium, sodium, calcium hydroxides

Weak bases

CH_3NH_2, NH_3, NaCN methylamine, ammonia, sodium cyanide

Neutralization reaction- an acid and base react to form a salt.

$$\boldsymbol{A}\text{OH} + \text{H}\boldsymbol{B} \rightarrow \boldsymbol{AB} + H_2O$$

base acid salt water

$$NaOH + HCl \rightarrow NaCl + H_2O$$

base acid salt water

Conjugate acid- formed when a base gains a hydrogen ion.
Conjugate base- what is formed when an acid donates a hydrogen ion.
Conjugate acid-base pair- 2 substances related by the loss or gain of a single electron pair, e.g. H_2O and OH^-.

$$NH_3\ (aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$$

ammonia *base; hydrogen ion acceptor*	water *acid; hydrogen ion donor*	ammonium ion*; conjugate acid*	hydroxide ion; *conjugate base*

Examples From the Flexbook

<u>Sample Problem 21.1: Using K_w in an Aqueous Solution</u>

Hydrochloric acid (HCl) is a strong acid, meaning that it is essentially 100% ionized in solution. What are the values of $[H^+]$ and $[OH^-]$ in a 2.0×10^{-3} M solution of HCl?

Step 1: List the known values and plan the problem.

Known

$[HCl] = 2.0 \times 10^{-3}$ M

$K_w = 1.0 \times 10^{-14}$

Unknown

$[H^+] = ?$ M

$[OH^-] = ?$ M

Because HCl is 100% ionized, the concentration of H^+ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one H^+ ion and one Cl^- ion. The concentration of OH^- can then be determined from $[H^+]$ and K_w .

Step 2: Solve.
$[H^+] = 2.0 \times 10^{-3}$ M
$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$
$[OH^-] = K_w/[H^+] = 1.0 \times 10^{-14} / 2.0 \times 10^{-3} = 5.0 \times 10^{-12}$ M

Step 3: Think about your result.
$[H^+]$ is much higher than $[OH^-]$ because the solution is acidic.

<u>Sample Problem 21.2: The pH of a Base</u>
Sodium hydroxide is a strong base. Find the pH of a solution prepared by dissolving 1.0 g of NaOH into enough water to make 1.0 L of solution.

Step 1: List the known values and plan the problem.
Known
mass of NaOH = 1.0 g
molar mass of NaOH = 40.00 g/mol
volume of solution = 1.0 L
$K_w = 1.0 \times 10^{-14}$

Unknown
pH of solution = ?

First, convert the mass of NaOH to moles. Second, calculate the molarity of the NaOH solution. Because NaOH is a strong base and is soluble in water, all of the dissolved NaOH will be dissociated, so $[OH^-]$ will be equal to the calculated concentration of the NaOH. Third, use K_w to calculate the $[H^+]$ in the solution. Lastly, calculate the pH.

Step 2: Solve.
1.0g NaOH x 1 mol NaOH / 40.00g NaOH = 0.025 mol NaOH

Molarity = 0.025 mol NaOH / 1.0 L = 0.025 M NaOH = 0.025 M OH^-

$[H^+] = K_w / [OH^-] = 1.0 \times 10^{-14} / 0.025M = 4.0 \times 10^{-13}M$

pH= $-\log[H^+] = -\log (4.0 \times 10^{-13}) = 12.40$

Step 3: Think about your result.

The solution is basic, so its pH is greater than 7. The reported pH is rounded to two decimal places because the original mass and volume each have two significant figures.

<u>Sample Problem 21.3: Using pOH</u>
Find the hydroxide concentration of a solution with a pH of 4.42.

Step 1: List the known values and plan the problem.

Known
pH = 4.42
pH + pOH = 14

Unknown
$[OH^-]$ = ? M

First, the pOH is calculated, followed by the $[OH^-]$.

Step 2: Solve.
pOH = 14 - pH = 14 - 4.42 = 9.58
$[OH^-] = 10^{-pOH} = 10^{-9.58} = 2.6 \times 10^{-10}$ M

use '-9.58' and press 'shift + log' or '10^x' depending on the calculator design

Step 3: Think about your result.
The pH is that of an acidic solution, and the resulting hydroxide-ion concentration is less than 1×10^{-7} M. The answer has two significant figures because the given pH has two decimal places.

K_a (acid ionization constant)- equilibrium constant for the ionization of an acid.

$$HA\ (aq) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + A^-\ (aq)$$

$$HA\ (aq) \rightarrow H^+\ (aq) + A^-\ (aq)$$

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

the more it ionizes, [product] > [reactant],

the stronger the acid: K_a > 1

K_b (base ionization constant) - equilibrium constant for the ionization of a base.

$$NH_3\ (aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

the more it ionizes, [product] > [reactant],

the stronger the base: $K_b > 1$

Titration - experiment in which a solution, whose concentration is known, is gradually added to a measured volume of another solution in order to determine its concentration. Titration calculations use $V_aM_a=V_bM_b$

All neutralization reactions between a strong acid and a strong base simplify to a **net ionic equation** - an equation that shows only the soluble, strong electrolytes reacting (ions) and do not include the spectator ions (unchanged in the reaction).

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-\ (aq) \rightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O\ (l)$$

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net ionic equation $H^+(aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$

Molality (m) number of moles of solute per kilogram of solvent.

Normality (N) - amount of solute to the total volume of solution; usually defined as the number of equivalents per liter of solution:

normality = number of equivalents / 1 L of solution

There is a very simple relationship between normality and molarity:

N = n × M (where n is an integer)

For a solution, **n** is the number of H^+ or OH^- provided by a formula unit of an acid or base.

Examples:

1M H_2SO_4 solution is the same as a 2N H_2SO_4 solution.

1M $Ca(OH)_2$ solution is the same as a 2N $Ca(OH)_2$ solution.