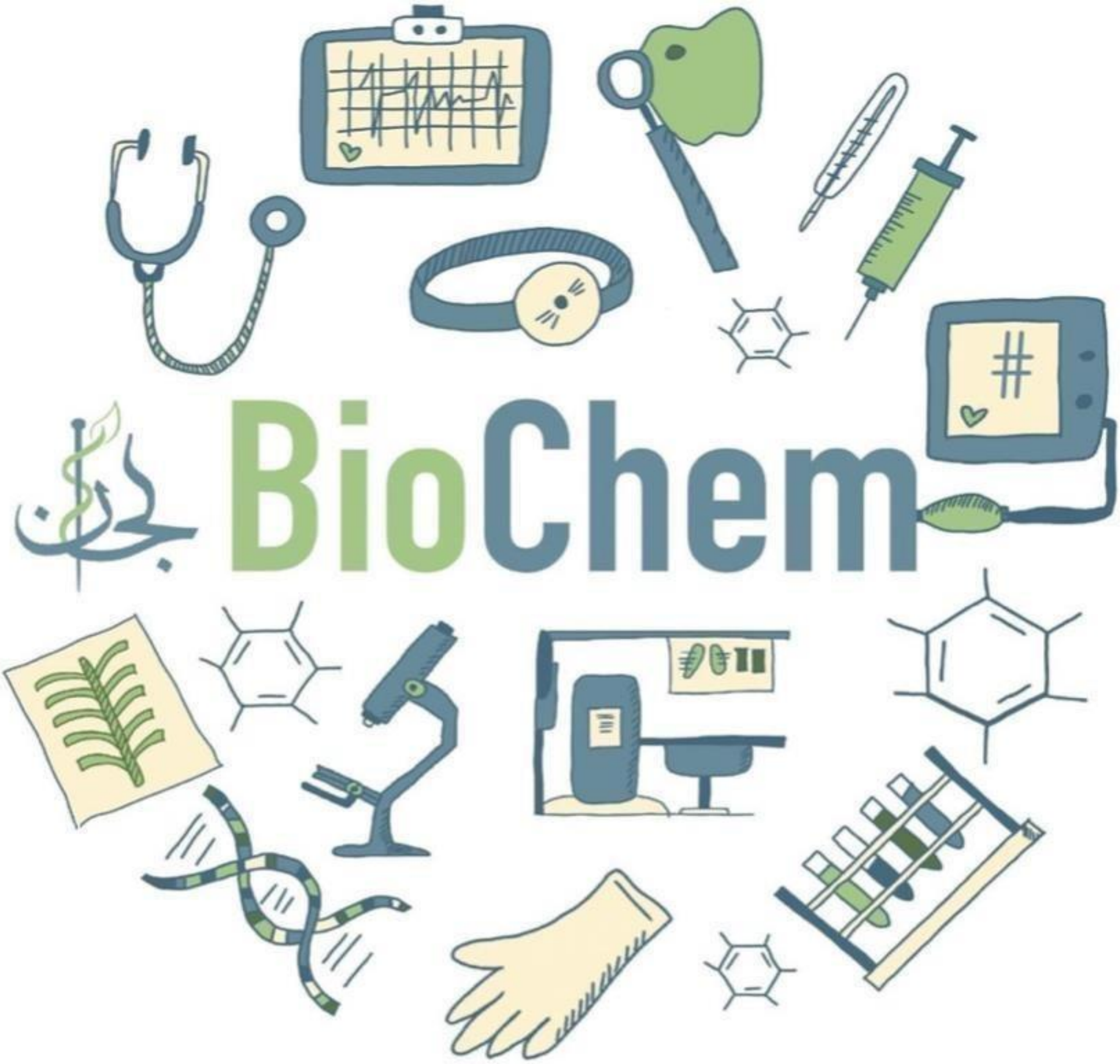

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Scientific Correction : Hala Zaghloul

Grammatical Correction: Hala Zaghloul

Final Correction : Hala Zaghloul

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Acids and Bases

We have different definitions given by the different scientists that studied them:

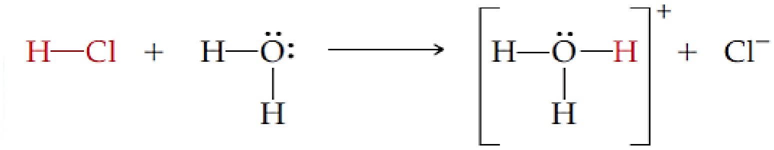
1- **Arrhenius** definition. 2- **Bronsted-Lowry** definition.

3-**Lewis** Definition.

- Arrhenius definition of acids and bases:

○ **Acid**: a substance that produces (H+) when dissolved in water

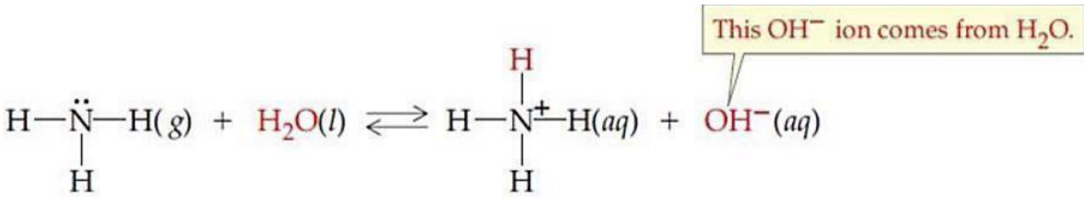
○ H+ Reacts with water producing hydronium ion (H3O+).



$$H{-}Cl + H{-}\ddot{O}{:}(H) \longrightarrow [H{-}\ddot{O}(H){-}H]^+ + Cl^-$$

Example: HCl donates (H+), H_2O accepts it and becomes hydronium ion whereas Cl- is released in the solution as a conjugate base.

○ **Base**: a substance that produces OH- (hydroxyl ion) when dissolved in water (in an aqueous solution).



This OH^- ion comes from H_2O.

$$H{-}\ddot{N}(H){-}H(g) + H_2O(l) \rightleftarrows H{-}\overset{+}{N}(H)(H){-}H(aq) + OH^-(aq)$$

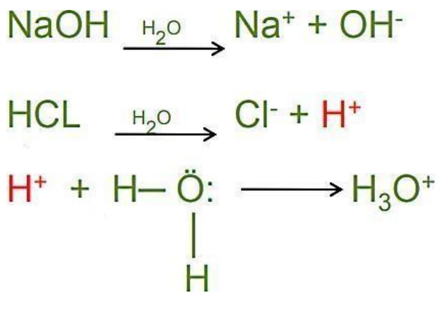
Example: NH_3 accepts protons from water molecules forming hydroxyl ions.

- Arrhenius definition of acids and bases:

○ Acids in H_2O are H^+ donors.

○ Bases in H_2O are OH^- donors.

Neutralization of acids and bases produces salt and water.



$$NaOH \xrightarrow{H_2O} Na^+ + OH^-$$

$$HCL \xrightarrow{H_2O} Cl^- + H^+$$

$$H^+ + H{-}\ddot{O}{:}(H) \longrightarrow H_3O^+$$

Drawbacks:

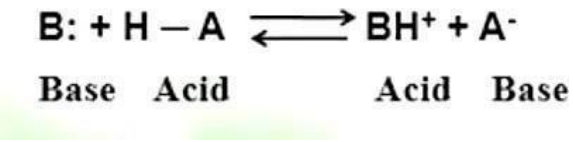
1. Reactions has to happen in aqueous solutions
2. H3O+ is released but not H+

- Bronsted-Lowry definition: (main focus is on the **proton** itself)

○ The Bronsted-Lowry acid **acid** (**proton donor**): any substance that's able to give a hydrogen ion (H+ -a proton) to another molecule.

○ Bronsted-Lowry base **base** (**proton acceptor**): any substance that accepts a

proton (H^+) from an acid. e.g: NaOH, NH_3, KOH

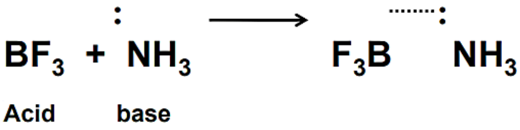


$$B: + H-A \rightleftharpoons BH^+ + A^-$$

The acid (HA) donates its proton and releases a conjugate base (A^-) while the base (B) accepts the proton and becomes a conjugate acid (BH^+).

- Lewis Definition of Acids and Bases:

○ Acids accept electrons.
○ Bases donate electrons (non-bonding pairs).



$$BF_3 + \overset{:}{N}H_3 \longrightarrow F_3B \cdots\!:NH_3$$

The base NH_3 has an unshared (non-bonding) pair of electrons that are accepted by the B in the acid BF_3.
*Note that electrons are partially shared between the acid and the base in this interaction.

Types of acids depending on the number of protons they contain:

o **Monoprotic** acid (1 proton): HCl, HNO_3, CH3COOH.
o **Polyprotic** acids (more than 1 proton):
A-Diprotic acid: H_2SO_4 B-Triprotic acid: H_3PO_3

Water is an amphoteric substance:

Water is neither considered an acid nor a base as it acts as an acid or a base depending on the solution or the substance that is found with it.
Substances that can act as an acid in one reaction and as a base in another are called **amphoteric substances**. (Example: water)

- With ammonia (NH_3), water acts as an acid because it donates a proton (hydrogen ion) to ammonia.

$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

- With hydrochloric acid, water acts as a base.

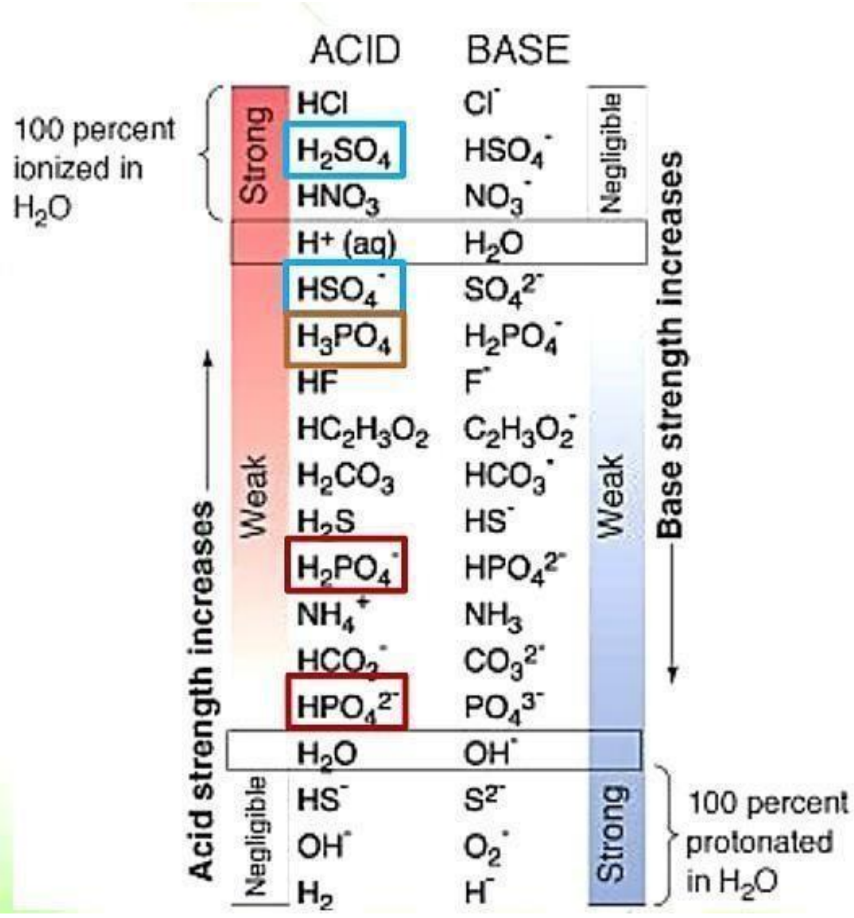
$$HCl + H_2O \rightarrow H3O^+ + Cl^-$$

Note: **amphoteric ≠ amphipathic**. Amphipathic means a substance with polar (hydrophilic) and nonpolar (hydrophobic) regions.

Acid/base strength:

- Acids differ in their ability to release protons.
  - ○ Strong acids dissociate 100% which means almost all of the molecules in the solution release their protons and become deprotonated. (you have to be very cautious when working with strong acids in the lab as they can be very harmful)
  - ○ Even strong acids relatively differ in their ability to donate protons and can be ranked according to their strength.
  - ○ Weak acids dissociate partially and release a very small number of protons.
  - ○ multi-protic acids (H_2SO_4, H_3PO_4) are going to give several protons, but each proton is donated at different strengths. The dissociation is gradual. (For example: H_2SO_4 doesn't lose its two protons in one step. First, it loses one proton and becomes HSO_4^-. Then, it loses the second proton and becomes SO_4^{2-}. Same thing for H_3PO_4 but in 3 steps).

    The first step is always the easiest, fastest, the step with the highest Ka and highest percentage of dissociation. Therefore, the first acid is the strongest, the second is weaker and if there is a third step it will be the weakest.

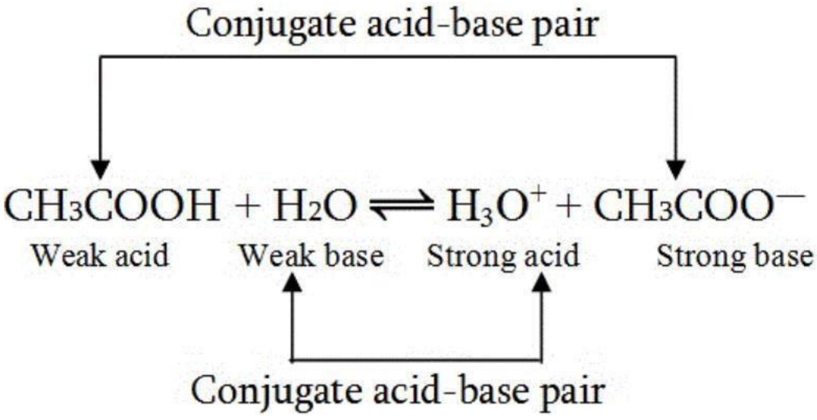


ACID	BASE
HCl	Cl^-
H_2SO_4	HSO_4^-
HNO_3	NO_3^-
H^+ (aq)	H_2O
HSO_4^-	SO_4^{2-}
H_3PO_4	$H_2PO_4^-$
HF	F^-
$HC_2H_3O_2$	$C_2H_3O_2^-$
H_2CO_3	HCO_3^-
H_2S	HS^-
$H_2PO_4^-$	HPO_4^{2-}
NH_4^+	NH_3
HCO_3^-	CO_3^{2-}
HPO_4^{2-}	PO_4^{3-}
H_2O	OH^-
HS^-	S^{2-}
OH^-	O_2^-
H_2	H^-

- Bases differ in their ability to accept protons. (Strong bases have strong affinity for protons).

Rule:

The stronger the acid, the weaker the conjugate base.



$$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$$

Strong vs. weak acids

- **Strong** acids and bases have **one-way** reactions because their disassociation is very high, it's almost 100%.

$$HCl \rightarrow H^+ + Cl^-$$
$$NaOH \rightarrow Na^+ + OH^-$$

- **Weak** acids and bases do not ionize completely **(two-way reactions)**.

$$HC_2H_3O_2 \leftrightarrow H^+ + C_2H_3O_2^-$$
$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

The double headed arrow indicates a two-way reaction.

whether the acid is strong (has complete disassociation) or weak (partial disassociation), it donates its proton to the solution, and the newly formed proton-lacking molecule would be the conjugate base.

Generally, there's an opposite relationship between the acid or the base and its conjugate partner. Acids produce conjugate bases and bases produce conjugate acids. Strong acids produce weak conjugate bases and weak acids produce strong conjugate bases and the same principle applies to bases and their conjugate acids.

Equilibrium constant and Acid dissociation constant Ka

Important for quantitative comparison rather than descriptive comparison

Acid/base solutions are at constant equilibrium.
We can write equilibrium constant (K_{eq}) for such reactions. (Which is equal to the concentration of the products multiplied by each other over the concentrations of the reactants multiplied by each other)

Equilibrium constant for acid dissociation K_a is equal to the concentration of the hydronium ion (or protons) multiplied by the concentration of the conjugate base over the concentration of the acid. (if there is water then it will be added to the denominator المقام. However, since H_2O concentration is constant 55.5 M we cross it out for simplicity)

$$\mathbf{HA \leftrightarrow H^+ + A^-}$$

$$K_a = \frac{[H_3O^+] \cdot [A^-]}{[HA]}$$

Note: H3O+ = H +

- Higher Ka values indicate higher proton concentrations and higher dissociation (strong acid) and vice versa.
- The value of the K_a indicates direction of reaction.

  - When K_a is greater than 1 the product side is favored. (the values in the numerator البسط are dominant).
  - When K_a is less than 1 the reactants are favored. (the value in the denominator is dominant).

In conclusion: Ka value indicates the strength of an acid and the direction of its reaction.

What is pK_a?

Notice how the values of K_a especially those of weaker acids contain negative powers that are difficult to deal with. So, to avoid this complexity it's better to convert those values into integers and that is by taking the negative log of these values and we end up with pK_a values.

$$pKa = -\log Ka$$

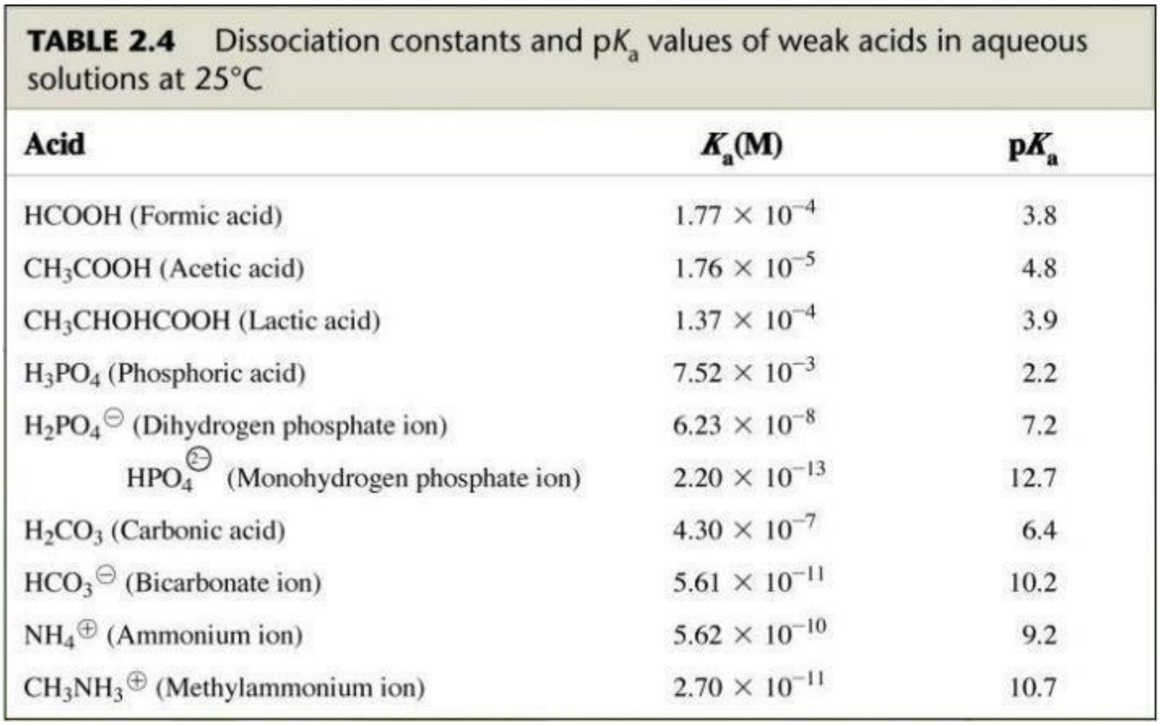


TABLE 2.4 Dissociation constants and pK_a values of weak acids in aqueous solutions at 25°C

Acid	K_a(M)	pK_a
HCOOH (Formic acid)	1.77×10^{-4}	3.8
CH_3COOH (Acetic acid)	1.76×10^{-5}	4.8
$CH_3CHOHCOOH$ (Lactic acid)	1.37×10^{-4}	3.9
H_3PO_4 (Phosphoric acid)	7.52×10^{-3}	2.2
$H_2PO_4^{\ominus}$ (Dihydrogen phosphate ion)	6.23×10^{-8}	7.2
$HPO_4^{2\ominus}$ (Monohydrogen phosphate ion)	2.20×10^{-13}	12.7
H_2CO_3 (Carbonic acid)	4.30×10^{-7}	6.4
$HCO_3^{\ominus}$ (Bicarbonate ion)	5.61×10^{-11}	10.2
$NH_4^{\oplus}$ (Ammonium ion)	5.62×10^{-10}	9.2
$CH_3NH_3^{\oplus}$ (Methylammonium ion)	2.70×10^{-11}	10.7

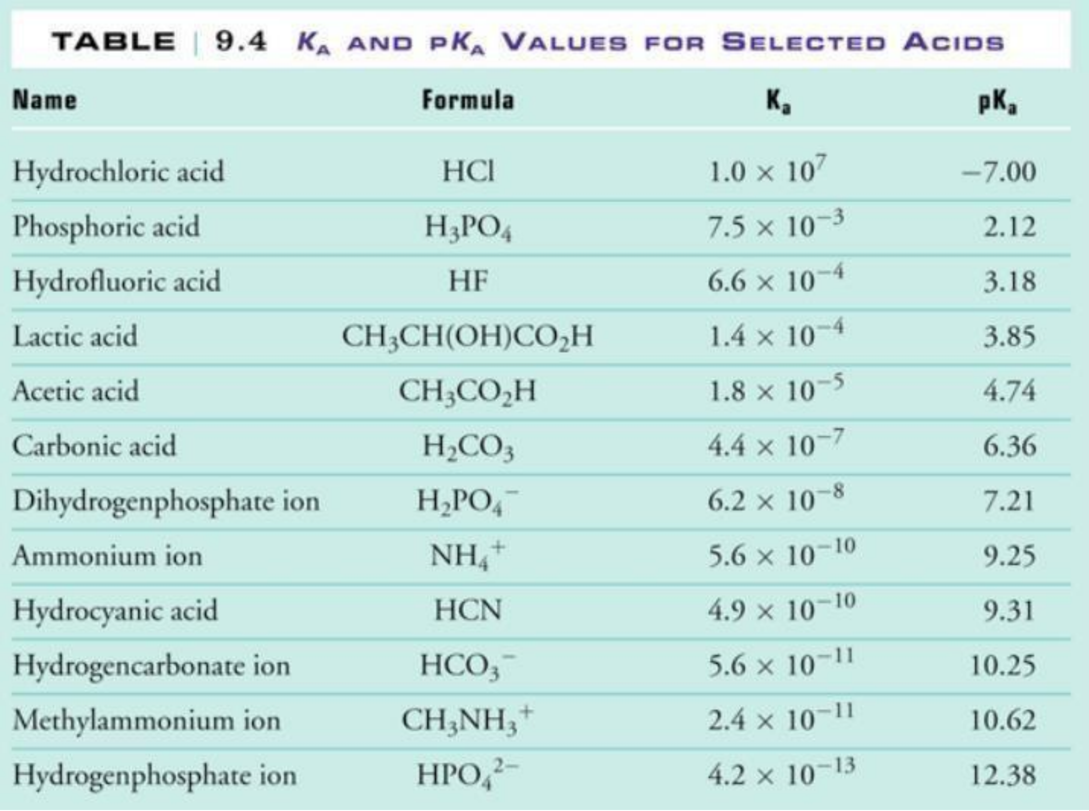
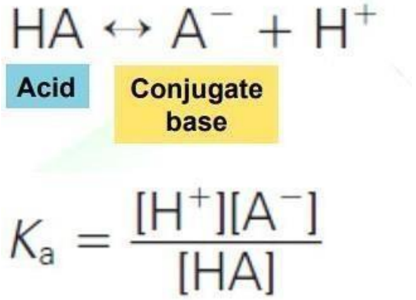


TABLE | 9.4 K_A AND pK_A VALUES FOR SELECTED ACIDS

Name	Formula	K_a	pK_a
Hydrochloric acid	HCl	1.0×10^{7}	−7.00
Phosphoric acid	H_3PO_4	7.5×10^{-3}	2.12
Hydrofluoric acid	HF	6.6×10^{-4}	3.18
Lactic acid	$CH_3CH(OH)CO_2H$	1.4×10^{-4}	3.85
Acetic acid	CH_3CO_2H	1.8×10^{-5}	4.74
Carbonic acid	H_2CO_3	4.4×10^{-7}	6.36
Dihydrogenphosphate ion	$H_2PO_4^{-}$	6.2×10^{-8}	7.21
Ammonium ion	NH_4^{+}	5.6×10^{-10}	9.25
Hydrocyanic acid	HCN	4.9×10^{-10}	9.31
Hydrogencarbonate ion	HCO_3^{-}	5.6×10^{-11}	10.25
Methylammonium ion	$CH_3NH_3^{+}$	2.4×10^{-11}	10.62
Hydrogenphosphate ion	HPO_4^{2-}	4.2×10^{-13}	12.38

the K_a values of strong acids are very high but their pK_a values are the smallest.
The values of K_a and pK_a are **inversely proportional**. The higher the Ka the lower the pK_a and vice versa.

To summarize the equilibrium constant, K_a:



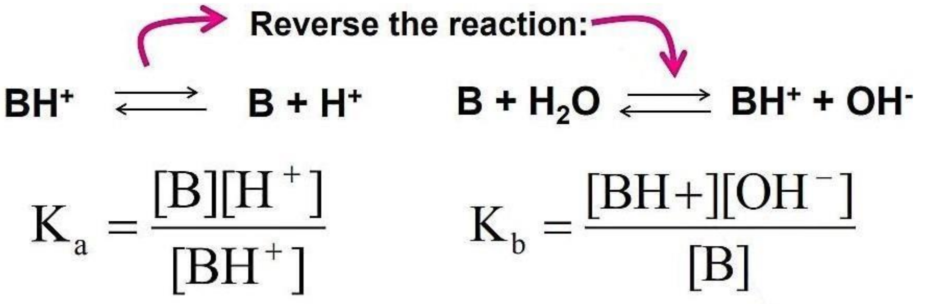
$$HA \leftrightarrow A^- + H^+$$

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

Larger K_a means: More dissociation, Smaller pK_a, Stronger acid.

Base dissociation constant (K_b):

Similarly, Equilibrium constant for base dissociation K_b is equal to the concentration of the hydroxyl ion multiplied by the concentration of the conjugate acid over the concentration of the base. (Again, if there is water then it will be added to the denominator المقام.
However, since H2O concentration is constant 55.5 M we cross it out for simplicity).



$$BH^+ \rightleftharpoons B + H^+ \qquad B + H_2O \rightleftharpoons BH^+ + OH^-$$

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

If you take the reversed reaction of the base dissociation:
$BH^+ \leftrightarrow B + H^+$
you would be calculating the K_a value for the base's conjugate acid (BH^+)

Common Weak Bases:

Weak bases have small K_b values with negative powers and strong bases have large values with positive powers.

TABLE 7.3 Values of K_b for Some Common Weak Bases

Name	Formula	Conjugate Acid	K_b
Ammonia	NH_3	NH_4^+	1.8×10^{-5}
Methylamine	CH_3NH_2	$CH_3NH_3^+$	4.38×10^{-4}
Ethylamine	$C_2H_5NH_2$	$C_2H_5NH_3^+$	5.6×10^{-4}
Aniline	$C_6H_5NH_2$	$C_6H_5NH_3^+$	3.8×10^{-10}
Pyridine	C_5H_5N	$C_5H_5NH^+$	1.7×10^{-9}

Measurement of concentration

Expression:

Solutions can be expressed in terms of its **concentration** or **molarity**.

Acids and bases can also be expressed in terms of their **normality (N)** or **equivalence (Eq).**

Molarity of solutions: (used in Ka and K_b calculations)

We know that moles of a solution are the amount in grams in relation to its molecular weight (MW or a.m.u.).

moles = grams / MW

Molarity is the number of moles over the volume of the solution.

M = moles / volume

A molar solution is one in which 1 liter of solution contains the number of grams equal to its molecular weight, which makes the moles **value** = 1

clarification: Moles = grams / MW
if grams = MW, then they would cancel out each other, the moles = 1.

Since (mol = grams / MW), you can calculate the grams of a chemical you need to dissolve in a known volume of water to obtain a certain concentration (M) using the following formula:

grams = M x vol x MW

Exercise: How many grams do you need to make 5M NaCl solution in 100 ml (MW 58.4)?

solution: grams = 58.4 x 5 M x 0.1 liter = 29.29 g

the volume must be in liters.

Equivalents (acids/bases and ions):

- When it comes to acids, bases and ions, it is useful to think of them as equivalents.
- **1 equivalent of a strong acid** contains **1 mol of H+ ions** (because strong acids dissociate completely and gives off a concentration of protons that is equal to the concentration of the acid and 1 eq is like 1 mol of acid that gives 1 mol of protons and here we're talking **specifically** about strong monoprotic acids), and 1 g-Eq of an acid is the mass in grams that contains 1 mol of H+ ions.
- Similarly, **1 equivalent of a strong base** contains **1 mol of OH^- ions** (because of the same reason stated in the previous point), and 1 g-Eq of a base is the mass in grams that contains 1 mol of OH^{--} ions.

- **For ions, a 1 g-Eq of any ion is defined as the molar mass of the ion divided by the ionic charge.**

Summary:

- equivalent for acids and bases = number of moles
- equivalent for ions = mw/charge

Note: we take the numerical value of the charge (its number), we're not concerned about its type (positive or negative)

Examples

- 1 mol HCl = 1 mol [H+] = 1 equivalent
- 1 mol H_2SO_4 = 2 mol [H+] = 2 equivalents
- One equivalent of Na^+ = 23.1 g
- One equivalent of Cl- = 35.5 g (we don't divide by the charge's negative value, just its number)
- One equivalent of Mg^{+2} = (24.3)/2 = 12.15 g

Remember: One equivalent of any acid neutralizes one equivalent of any base.

Exercise: Calculate milligrams of Ca^{+2} in blood if total concentration of Ca^{+2} is 5 mEq/L.

Solution: 1 Eq of Ca+2 = 40.1 g/2 =20.1 g

Grams of Ca+2 in blood =

= (5 mEq/L) x (1 Eq/1000 mEq) x (20.1 g/ 1 Eq)

= 0.1 g/L

=100 mg/L

to calculate the mass in grams:
we first need to convert mEq to Eq: by diving over 1000 and then multiply by the one gram Eq of Ca^{+2} we calculated above.

Normal solutions:

Normality (N) considers both the ***molarity*** of the solution and the ***equivalent*** content of the acid or base.

N= n x M (where n is an integer)

M: molarity n: no. of equivalence

For an acid solution, n is the number of H^+ provided by a formula unit of acid. Similarly, for a base solution is the number of OH^- a base can donate.

Examples:

- 3M H_2SO_4 solution is the same as a 6N H_2SO_4 solution

molarity (M) = 3 whereas normality (N) = 6 since it considers the number of equivalence and in this case, it equals the number of protons considering H_2SO_4 is an acid. So, N = M x no. of Eq = 3 x 2 = 6.

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

the number of equivalence is the number of OH- ions since Ca(OH)2 is a base.

N = M x no. of Eq = 1 x 2 = 2

Remember! The normality of a solution is NEVER less than the molarity: it's always larger or at least it equals M

Exercise: What is the normality of H2SO4 solution made by dissolving 6.5 g into 200 mL? (MW =98)

M = mol / MW or M = grams / (MW x vol)

6.5 =g / (98 x 0.2 L)

19.6 / 6.5 =

0.33 =

N = M x n = 0.33 x 2 = 0.66 N

Normality and equivalents:

For Neutralization:

$$Normality(N) = \frac{\textbf{Equivalent of acid or base}}{\textbf{Liters of solution}}$$

Based on the equation above, since no. of eq of an acid is neutralized by the same n eq of a base, then (N x liters) of an acid is neutralized by (N x liters) of a base.

However, this is not the way to calculate the equivalents of molecules but only to calculate how many equivalents of an acid or a base I need to be neutralized by the same amount of the opposite molecule.

Examples:

Note: titration is neutralization and will be covered in more details later on.

10.92 Titration of a 12.0 mL solution of HCl requires 22.4 mL of 0.12 M NaOH. What is the molarity of the HCl solution?

they need equal equivalent to neutralize each other

M1 x Vol1 = M2 x Vol2

0.12 x 22.4 = M2 x 12

M2 = (0.12 x 22.4) / 12

M2 = 0.224 M

10.93 What volume of 0.085 M HNO_3 is required to titrate 15.0 mL of 0.12 M $Ba(OH)_2$ solution?

Since there's a difference here as HNO_3 is monoprotic and $Ba(OH)_2$ produces 2 OH^- ions we **convert** the molarity to the normality.

convert M to N first
N of HNO_3 = 1 x 0.085 = 0.085 N
N of $Ba(OH)_2$ = 2 x 0.12 = 0.24
N1 x Vol1 = N2 x Vol2
0.24 x 15 = 0.085 x Vol2
Vol2 = (0.24 x 15) / 0.085
Vol2 = 42.35 mL

Hint: If number of (H^+) (OH^-) is different, convert M to N, find the answer in N, then convert N to M.

Apply same method for third question

10.94 Titration of a 10.0 mL solution of KOH requires 15.0 mL of 0.0250 M H_2SO_4 solution. What is the molarity of the KOH solution?

Convert M to N first:

N of H_2SO_4 = 2 × 0.025 = 0.05

N of KOH = M × 1 = M

N1 × VOL1 = N2 × VOL2

0.05 × 15 = M × 10

M = (0.05 × 15)/10

M= 0.075 M

Ionization of water:

- Water dissociates into both hydronium (H_3O^+) and hydroxyl (OH^-) ions (because it's not an acid nor a base and the reaction is in equilibrium since it's not strong).
- For simplicity, we refer to the hydronium ion as a hydrogen ion (H^+) and write the reaction equilibrium as:

$$\mathrm{H_2O(l)} \rightleftharpoons \mathrm{H^+(aq)} + \mathrm{HO^-(aq)}$$

Equilibrium constant of water:

H_2O dissociates to a slight extent to form hydrogen (H^+) and hydroxyl (OH^-) ions

$$\mathbf{H_2O \leftrightarrow H^+ + OH^-}$$

The equilibrium constant K_{eq} of the dissociation of water is:

$$K_{eq} = \frac{[H+][OH-]}{[H2O]}$$

We don't cross the concentration of water out since it's the only one here and the concentrations of H^+ and OH^- are equal.

The equilibrium constant for water ionization under standard conditions is 1.8×10^{-16} M.

Ion product of water:

Kw can be derived from multiplying Ka and Kb since water is neither an acid nor a base.

$$\mathrm{HA}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{A^-}(aq)$$

$$\mathrm{A^-}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{HA}(aq) + \mathrm{OH^-}(aq)$$

$$K_a \cdot K_b = \left(\frac{[\mathrm{H_3O^+}]\cancel{[\mathrm{A^-}]}}{\cancel{[\mathrm{HA}]}}\right)\left(\frac{\cancel{[\mathrm{HA}]}[\mathrm{OH^-}]}{\cancel{[\mathrm{A^-}]}}\right)$$

$$= [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

the concentration of H^+ times OH^- ions in any solution is a constant value, which is 10^{-14}

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

K_w is equal to k_{eq} of water times $[H_2O]$

Kw (ion product of water):

- Since there are 55.6 moles of water in 1 liter, the product of the hydrogen and hydroxide ion concentrations results in a value of 1 x 10-14 for:

$$K_{eq}\,(55.5\ M) = [H^{\oplus}]\,[OH^{\ominus}]$$

Kw determines the distribution of these ions in different solutions.

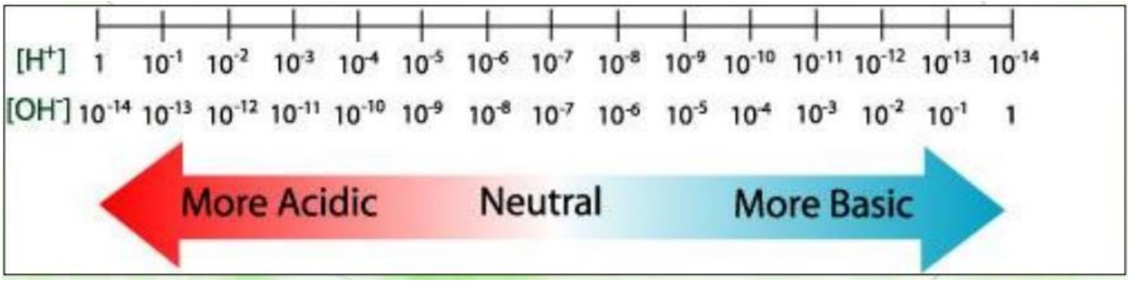
- This constant, Kw, is called the ion product for water.

$$K_w = [H^{\oplus}]\,[OH^{\ominus}] = 1.0 \times 10^{-14}\ M^2$$

[H+] of pure water is only 0.0000001 M

[H+] and [OH-]:

- For pure water, there are equal concentrations of $[H^+]$ and $[OH^-]$, each with a value of 1×10^{-7} M.
- Since Kw is a fixed value, the concentrations of $[H^+]$ and $[OH^-]$ are inversely changing.
- If the concentration of H^+ is high, then the concentration of OH^- must be low, and vice versa. For example, if $[H^+] = 10^{-2}$ M, then $[OH^-] = 10^{-12}$ M.



$[H^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1

Neutral solutions contain equal amounts of H+ and OH- ions. Solutions with higher OH- concentration are more basic than those with hgher H+ content.

GOOD LUCK!!!😊

Extra Practice Questions:

Q1) What **volume** of 0.15M $Ba(OH)_2$ is required to neutralize 45 ml of 0.29M HCl.

Q2) It takes 83 mL of a 0.45 M NaOH solution to neutralize 235 mL of an HCl solution. What is the **concentration** of the HCl solution?

Q3) It takes 38 mL of 0.75 M NaOH solution to completely neutralize 155 mL of a sulfuric acid solution (H_2SO_4). What is the **concentration** of the H_2SO_4 solution?

Q4) It takes 12.5 mL of 0.30 M HCl solution to neutralize 285 mL of NaOH solution. What is the **mass** in grams of the NaOH solution? (mw of NaOH = 40g/mol)

Q5) 33 mL of 3M Hydrochloric acid HCl is titrated with sodium hydroxide NaOH to form water and sodium chloride. How many **mmols** of sodium hydroxide are consumed in this rxn?

ANSWERS KEY	
QUESTION	ANSWER
Q1	43.5 ml* you can also convert to Liters
Q2	0.16 M
Q3	0.092 M
Q4	0.15 g
Q5	~100 mmol