Equilibrio ácido-base

- Teoría de Arrhenius. Una sustancia es un ácido si, en solución acuosa, es capaz ceder H^+. Una sustancia es un base si, en solución acuosa, es capaz ceder OH^-.

  Ello implica que las especies ácida o básicas, deben contener en su estructura H y OH para liberar.

ácido

$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(ac)} + Cl^-_{(ac)}$$

$$HC_2H_3O_{2(ac)} \xleftrightarrow{H_2O} H^+_{(ac)} + C_2H_3O^-_{2(ac)}$$

base

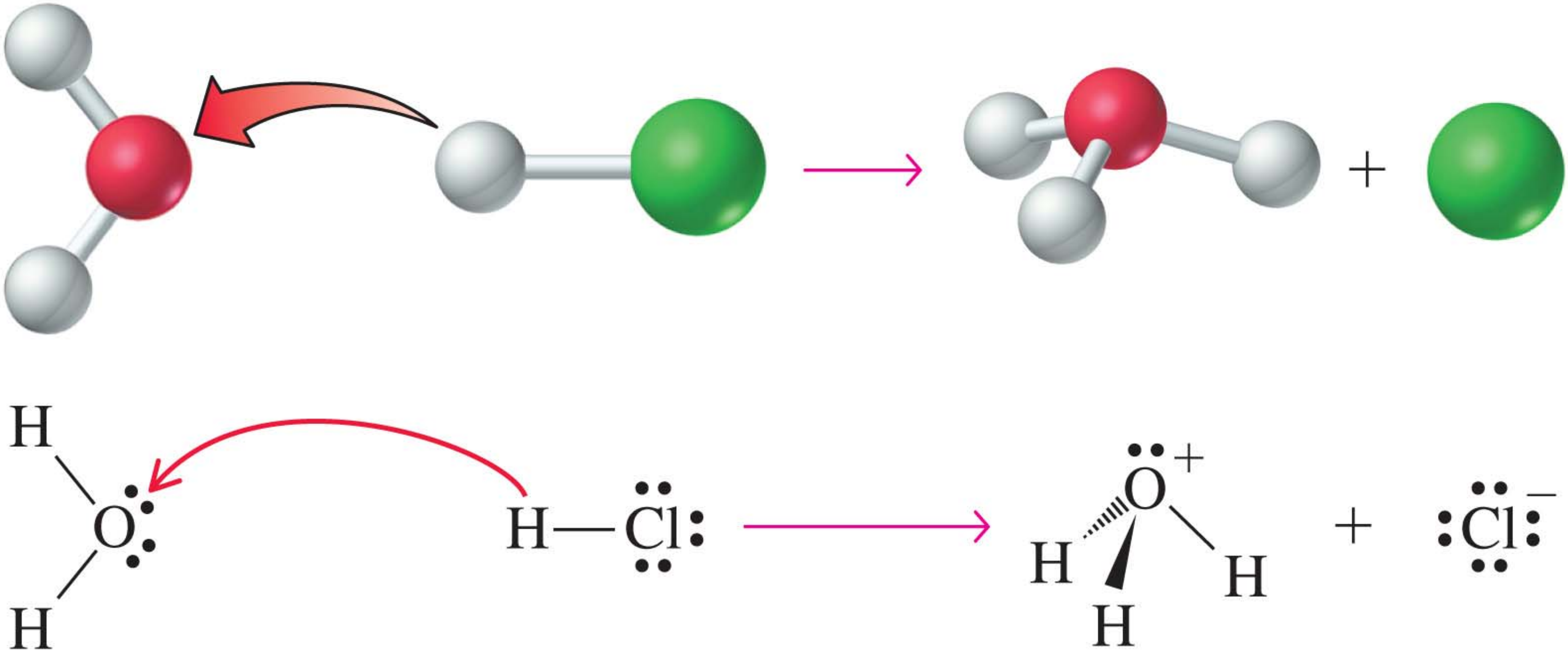
$$NaOH_{(s)} \xrightarrow{H_2O} OH^-_{(ac)} + Na^+_{(ac)}$$

$$NH_2OH_{(ac)} \xleftrightarrow{H_2O} NH^+_{2(ac)} + OH^-_{(ac)}$$

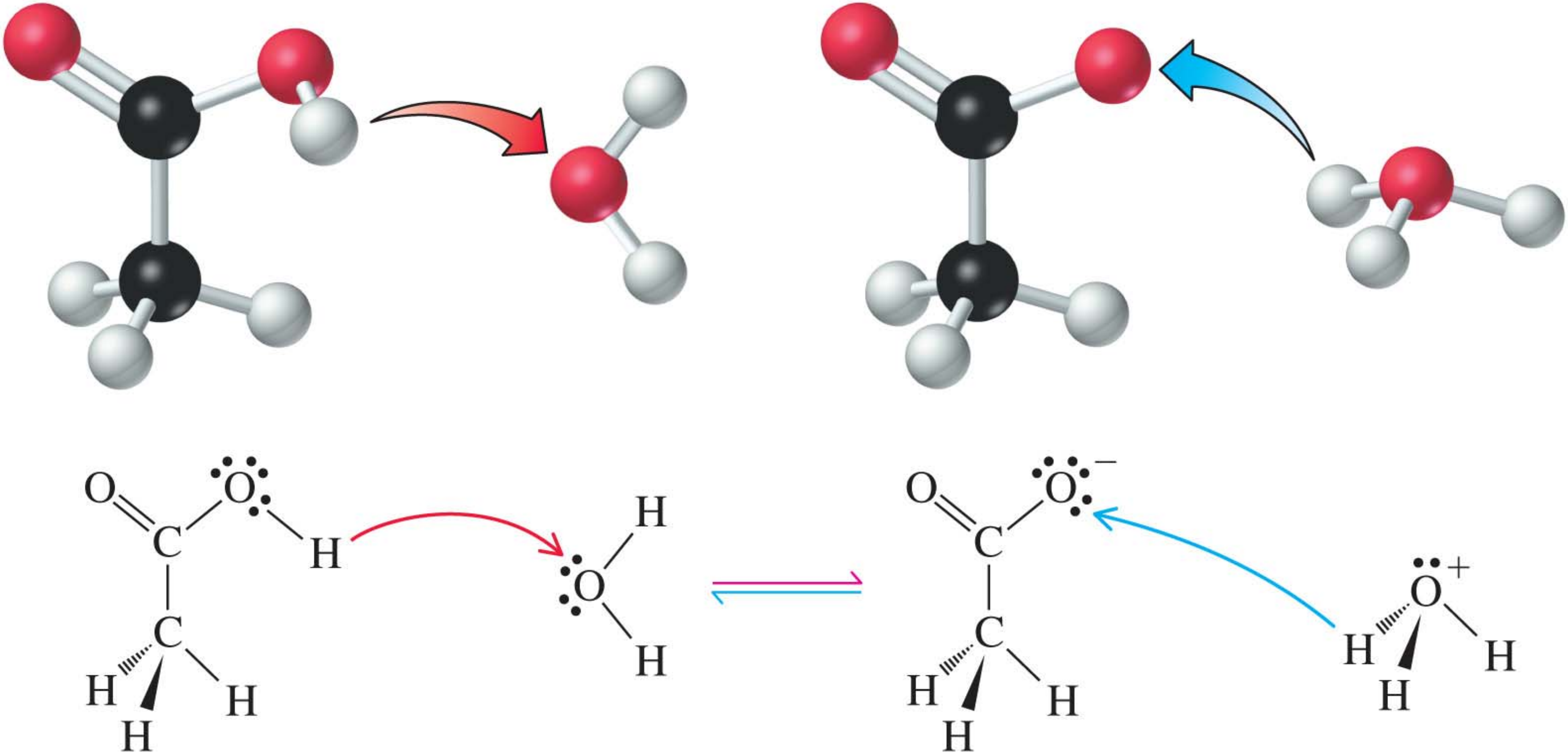
Teoría de Bronsted-Lowry

- Dador/Aceptor. Transferencia de protones.
- No supeditada a un solvente

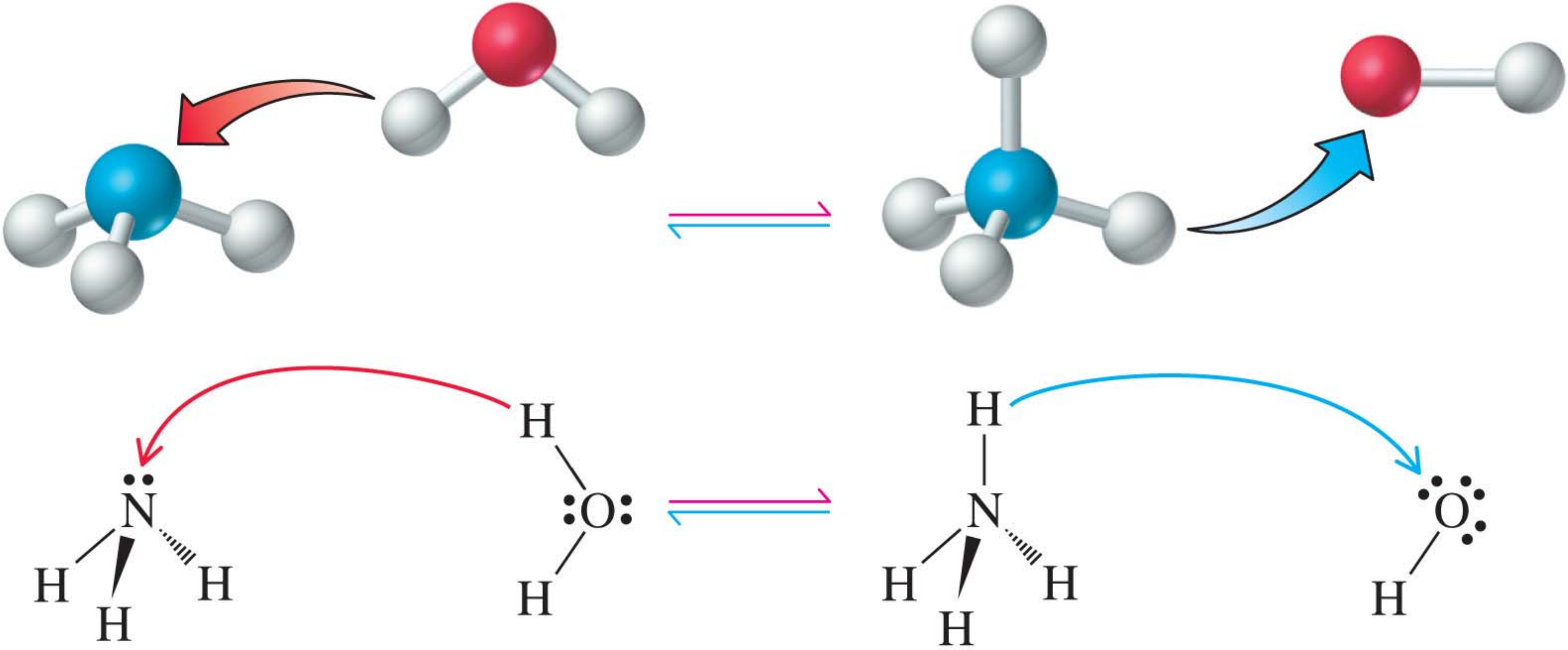
Según Bronsted un ***ácido es un donador de protones (H$^+$)*** *y una* ***base es un aceptor de protones (H$^+$)****.*



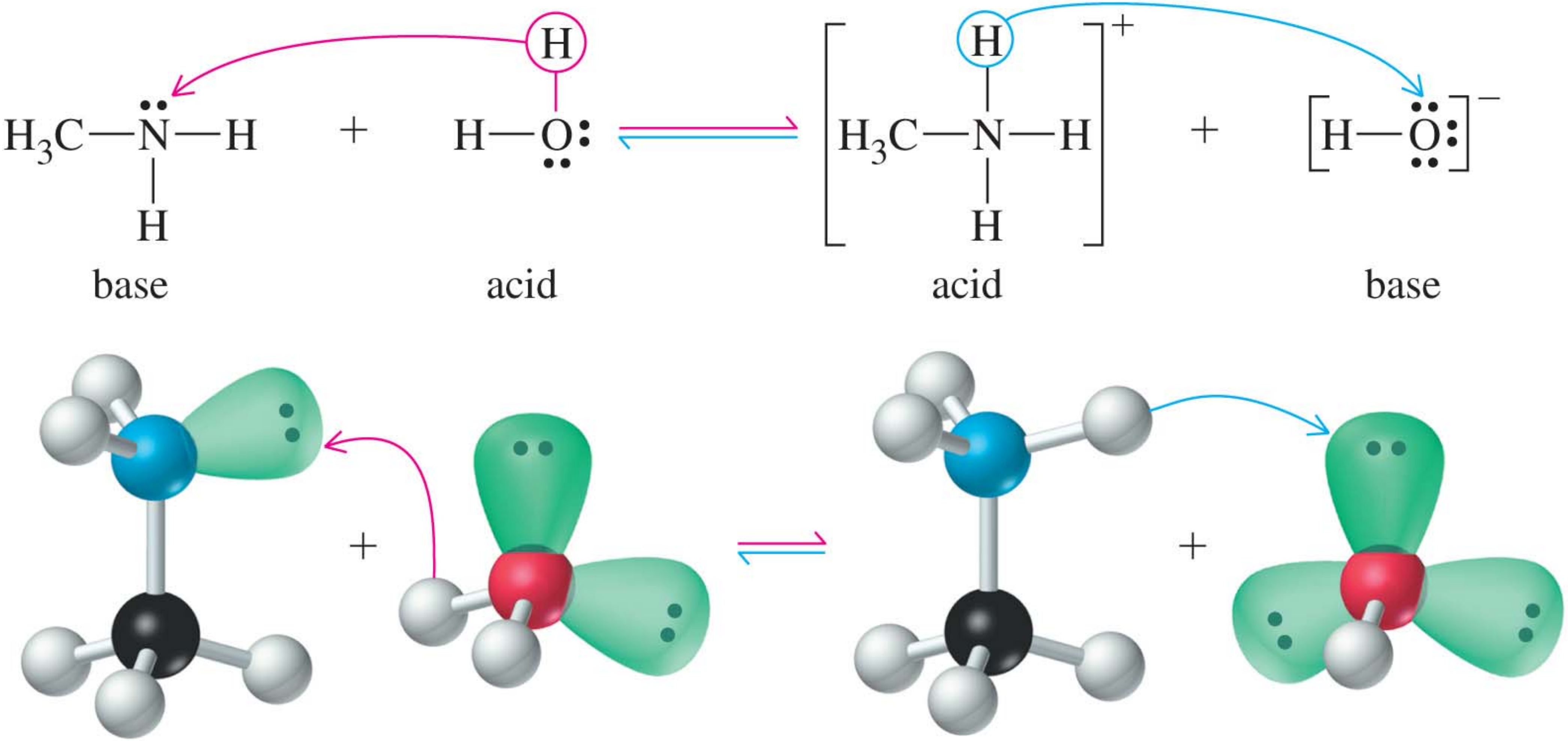
$$\underset{\text{base(1)}}{H_2O} + \underset{\text{acid(2)}}{HCl} \longrightarrow \underset{\text{acid(1)}}{H_3O^+} + \underset{\text{base(2)}}{Cl^-}$$



$$\underset{\text{acid(1)}}{HC_2H_3O_2} + \underset{\text{base(2)}}{H_2O} \rightleftharpoons \underset{\text{base(1)}}{C_2H_3O_2^-} + \underset{\text{acid(2)}}{H_3O^+}$$



$$\underset{\text{base(1)}}{NH_3} + \underset{\text{acid(2)}}{H_2O} \rightleftharpoons \underset{\text{acid(1)}}{NH_4^+} + \underset{\text{base(2)}}{OH^-}$$


base
acid
acid
base

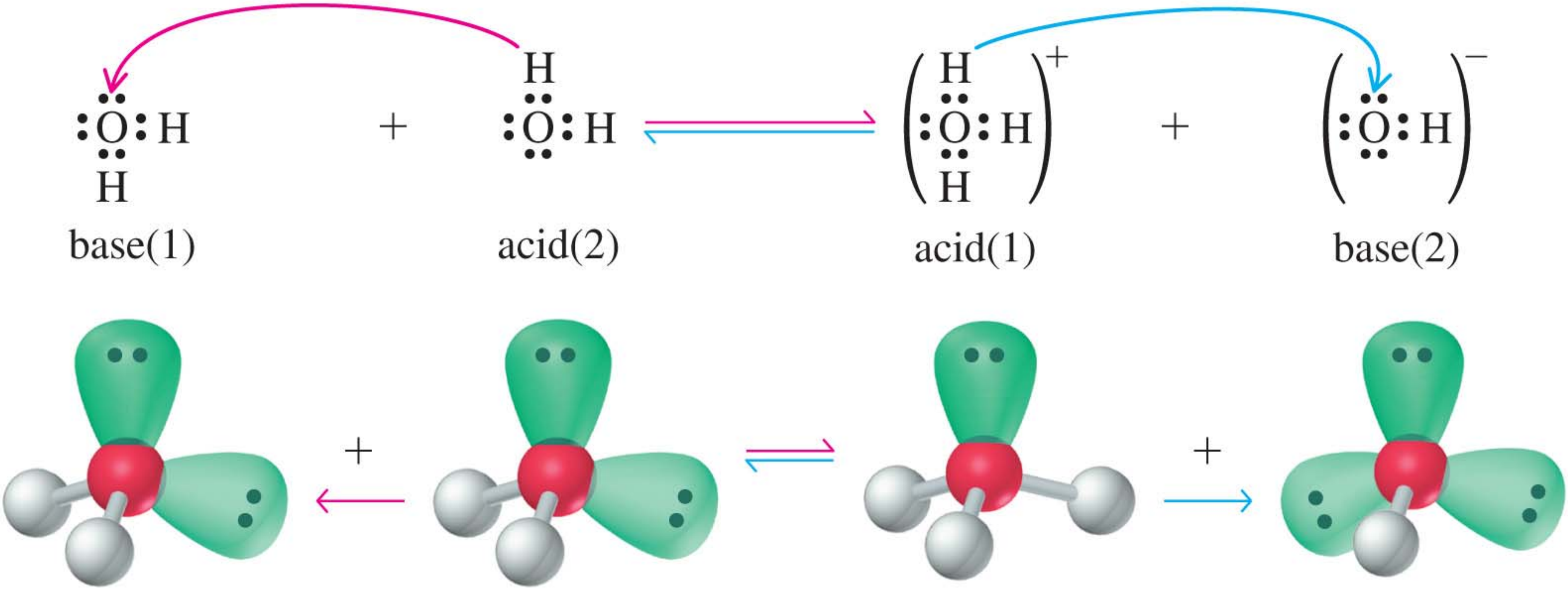
La reacción de neutralización es una reacción entre un ácido y una base

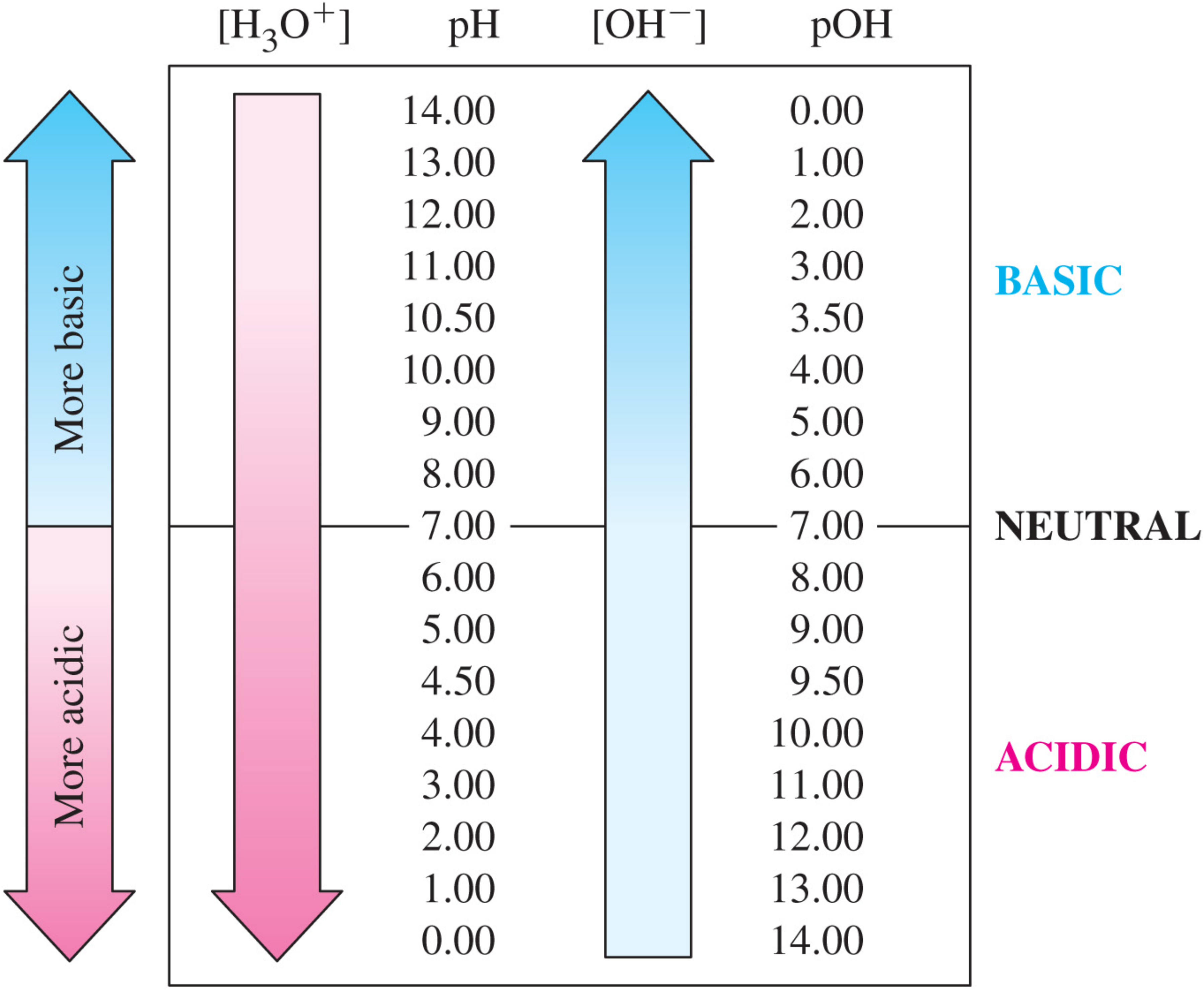
$$HCl_{(ac)} + NaOH_{(ac)} \rightarrow NaCl + H_2O \; (ec. molecular)$$

$$H^{+}_{(ac)} + Cl^{-}_{(ac)} + Na^{+}_{(ac)} + OH^{-}_{(ac)} \rightarrow Na^{+}_{(ac)} + Cl^{-}_{(ac)} + H_2O_{(l)} \; (ec. ionica)$$

$$H^{+}_{(ac)} + OH^{-}_{(ac)} \rightarrow +H_2O_{(l)}$$

Na^+ y Cl^- son iones espectadores


base(1)
+
acid(2)
acid(1)
+
base(2)



$[H_3O^+]$	pH	$[OH^-]$	pOH
	14.00		0.00
	13.00		1.00
	12.00		2.00
	11.00		3.00
	10.50		3.50
	10.00		4.00
	9.00		5.00
	8.00		6.00
	7.00		7.00
	6.00		8.00
	5.00		9.00
	4.50		9.50
	4.00		10.00
	3.00		11.00
	2.00		12.00
	1.00		13.00
	0.00		14.00

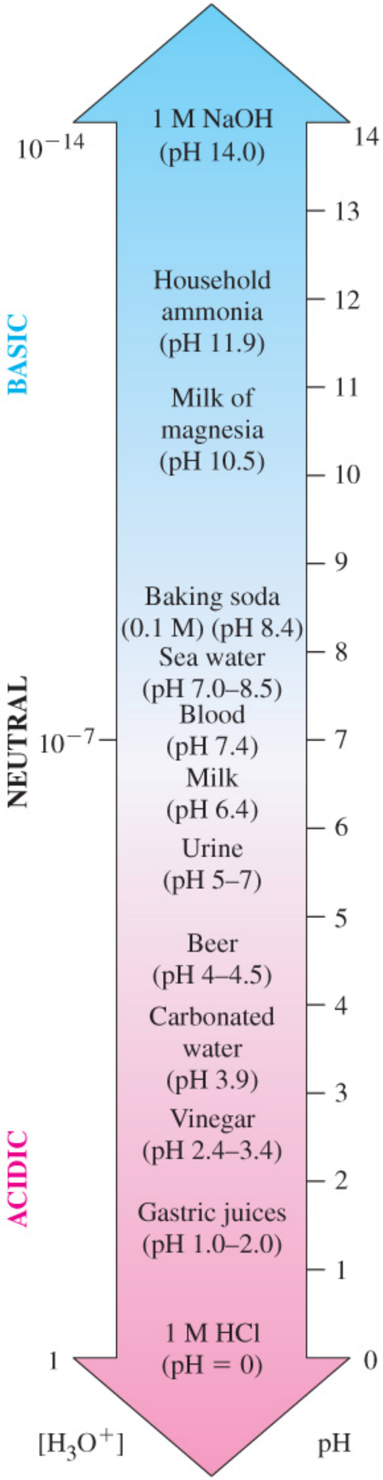

1 M NaOH
(pH 14.0)
Household ammonia
(pH 11.9)
Milk of magnesia
(pH 10.5)
Baking soda
(0.1 M) (pH 8.4)
Sea water
(pH 7.0–8.5)
Blood
(pH 7.4)
Milk
(pH 6.4)
Urine
(pH 5–7)
Beer
(pH 4–4.5)
Carbonated water
(pH 3.9)
Vinegar
(pH 2.4–3.4)
Gastric juices
(pH 1.0–2.0)
1 M HCl
(pH = 0)
BASIC
NEUTRAL
ACIDIC
10^{-14}
10^{-7}
1
$[H_3O^+]$
14
13
12
11
10
9
8
7
6
5
4
3
2
1
0
pH

TABLE 16.2
The Common Strong Acids and Strong Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
$HClO_4$	RbOH
HNO_3	CsOH
H_2SO_4[a]	$Mg(OH)_2$
	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

---

[a] H_2SO_4 ionizes in two distinct steps. It is a strong acid only in its first ionization

	Ionization Equilibrium	Ionization Constant K	pK
Acid		K_a =	pK_a =
Iodic acid	$HIO_3 + H_2O \rightleftharpoons H_3O^+ + IO_3^-$	1.6×10^{-1}	0.80
Chlorous acid	$HClO_2 + H_2O \rightleftharpoons H_3O^+ + ClO_2^-$	1.1×10^{-2}	1.96
Chloroacetic acid	$HC_2H_2ClO_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_2ClO_2^-$	1.4×10^{-3}	2.85
Nitrous acid	$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$	7.2×10^{-4}	3.14
Hydrofluoric acid	$HF + H_2O \rightleftharpoons H_3O^+ + F^-$	6.6×10^{-4}	3.18
Formic acid	$HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^-$	1.8×10^{-4}	3.74
Benzoic acid	$HC_7H_5O_2 + H_2O \rightleftharpoons H_3O^+ + C_7H_5O_2^-$	6.3×10^{-5}	4.20
Hydrazoic acid	$HN_3 + H_2O \rightleftharpoons H_3O^+ + N_3^-$	1.9×10^{-5}	4.72
Acetic acid	$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}	4.74
Hypochlorous acid	$HOCl + H_2O \rightleftharpoons H_3O^+ + OCl^-$	2.9×10^{-8}	7.54
Hydrocyanic acid	$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$	6.2×10^{-10}	9.21
Phenol	$HOC_6H_5 + H_2O \rightleftharpoons H_3O^+ + C_6H_5O^-$	1.0×10^{-10}	10.00
Hydrogen peroxide	$H_2O_2 + H_2O \rightleftharpoons H_3O^+ + HO_2^-$	1.8×10^{-12}	11.74
Base		K_b =	pK_b =
Diethylamine	$(C_2H_5)_2NH + H_2O \rightleftharpoons (C_2H_5)_2NH_2^+ + OH^-$	6.9×10^{-4}	3.16
Ethylamine	$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$	4.3×10^{-4}	3.37
Ammonia	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	1.8×10^{-5}	4.74
Hydroxylamine	$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$	9.1×10^{-9}	8.04
Pyridine	$C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH^-$	1.5×10^{-9}	8.82
Aniline	$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$	7.4×10^{-10}	9.13

TABLE 16.1 Relative Strengths of Some Common Brønsted–Lowry Acids and Bases

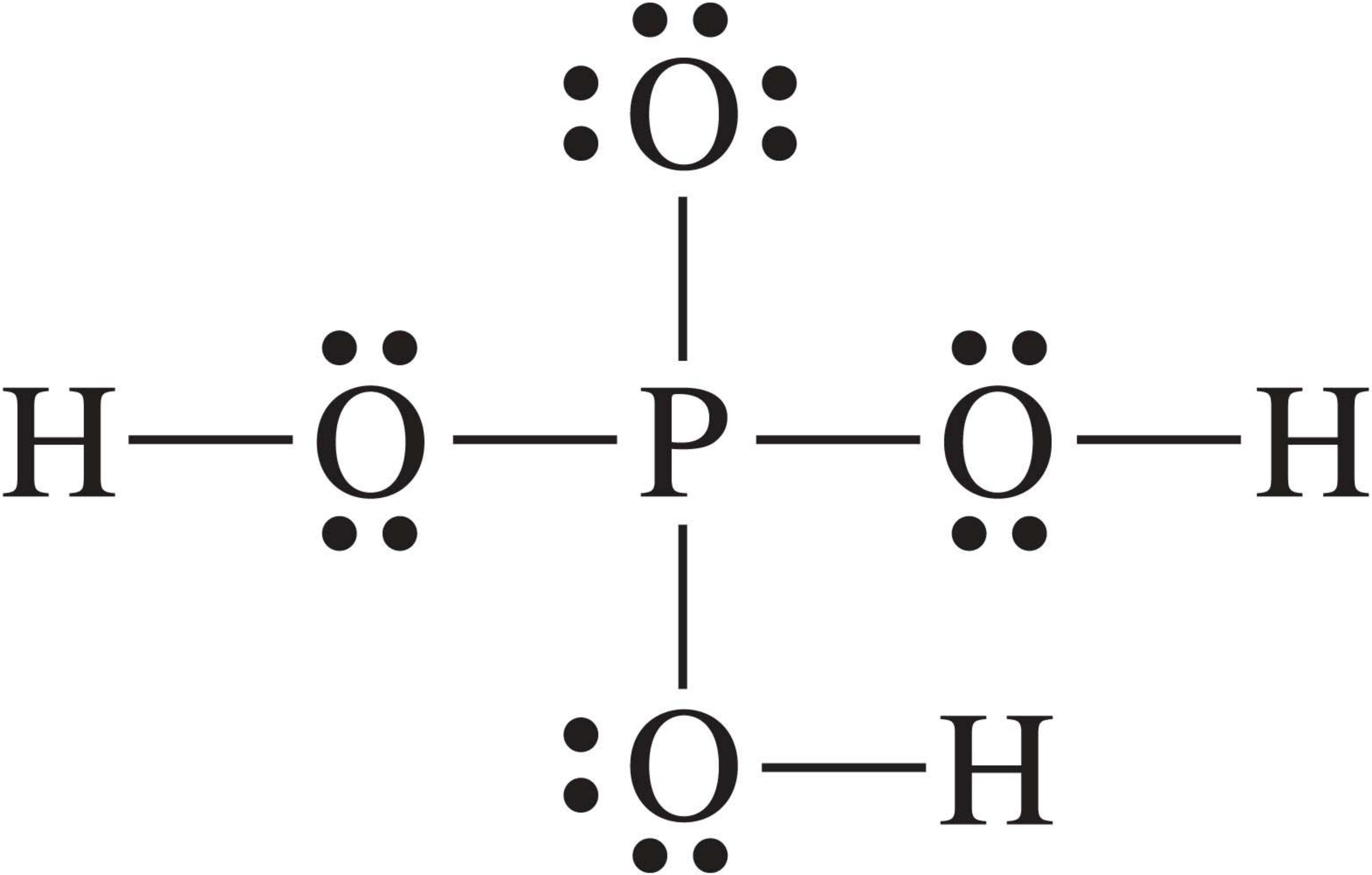
Acid		Conjugate Base	
Perchloric acid	$HClO_4$	Perchlorate ion	ClO_4^-
Hydroiodic acid	HI	Iodide ion	I^-
Hydrobromic acid	HBr	Bromide ion	Br^-
Hydrochloric acid	HCl	Chloride ion	Cl^-
Sulfuric acid	H_2SO_4	Hydrogen sulfate ion	HSO_4^-
Nitric acid	HNO_3	Nitrate ion	NO_3^-
Hydronium ion[a]	H_3O^+	Water[a]	H_2O
Hydrogen sulfate ion	HSO_4^-	Sulfate ion	SO_4^{2-}
Nitrous acid	HNO_2	Nitrite ion	NO_2^-
Acetic acid	$HC_2H_3O_2$	Acetate ion	$C_2H_3O_2^-$
Carbonic acid	H_2CO_3	Hydrogen carbonate ion	HCO_3^-
Ammonium ion	NH_4^+	Ammonia	NH_3
Hydrogen carbonate ion	HCO_3^-	Carbonate ion	CO_3^{2-}
Water	H_2O	Hydroxide ion	OH^-
Methanol	CH_3OH	Methoxide ion	CH_3O^-
Ammonia	NH_3	Amide ion	NH_2^-

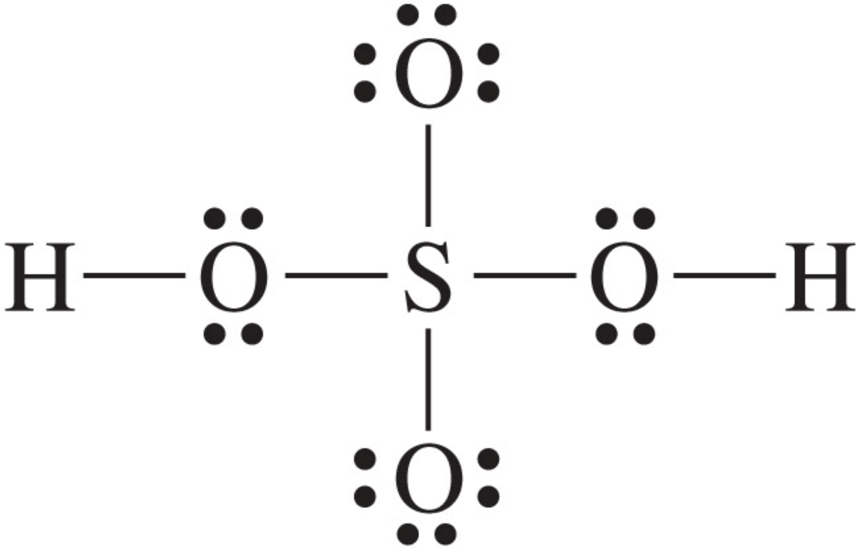
Increasing acid strength ↑

Increasing base strength ↓

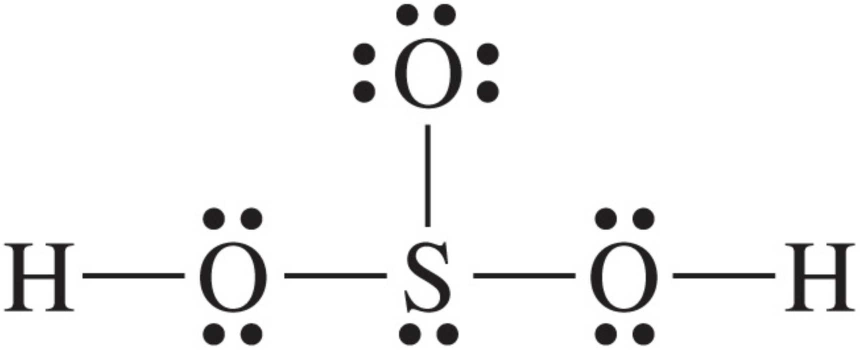
[a] The hydronium ion–water combination refers to the ease with which a proton is passed from one water molecule to another; that is, $H_3O^+ + H_2O \rightleftharpoons H_2O + H_3O^+$

Acid	**Ionization Equilibria**	**Ionization Constants, K**	**pK**
Hydrosulfuric[a]	$H_2S + H_2O \rightleftharpoons H_3O^+ + HS^-$	$K_{a_1} = 1.0 \times 10^{-7}$	$pK_{a_1} = 7.00$
	$HS^- + H_2O \rightleftharpoons H_3O^+ + S^{2-}$	$K_{a_2} = 1 \times 10^{-19}$	$pK_{a_2} = 19.0$
Carbonic[b]	$H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^-$	$K_{a_1} = 4.4 \times 10^{-7}$	$pK_{a_1} = 6.36$
	$HCO_3^- + H_2O \rightleftharpoons H_3O^+ + CO_3^{2-}$	$K_{a_2} = 4.7 \times 10^{-11}$	$pK_{a_2} = 10.33$
Phosphoric	$H_3PO_4 + H_2O \rightleftharpoons H_3O^+ + H_2PO_4^-$	$K_{a_1} = 7.1 \times 10^{-3}$	$pK_{a_1} = 2.15$
	$H_2PO_4^- + H_2O \rightleftharpoons H_3O^+ + HPO_4^{2-}$	$K_{a_2} = 6.3 \times 10^{-8}$	$pK_{a_2} = 7.20$
	$HPO_4^{2-} + H_2O \rightleftharpoons H_3O^+ + PO_4^{3-}$	$K_{a_3} = 4.2 \times 10^{-13}$	$pK_{a_3} = 12.38$
Sulfurous[c]	$H_2SO_3 + H_2O \rightleftharpoons H_3O^+ + HSO_3^-$	$K_{a_1} = 1.3 \times 10^{-2}$	$pK_{a_1} = 1.89$
	$HSO_3^- + H_2O \rightleftharpoons H_3O^+ + SO_3^{2-}$	$K_{a_2} = 6.2 \times 10^{-8}$	$pK_{a_2} = 7.21$
Sulfuric[d]	$H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-$	K_{a_1} = very large	$pK_{a_1} < 0$
	$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$	$K_{a_2} = 1.1 \times 10^{-2}$	$pK_{a_2} = 1.96$

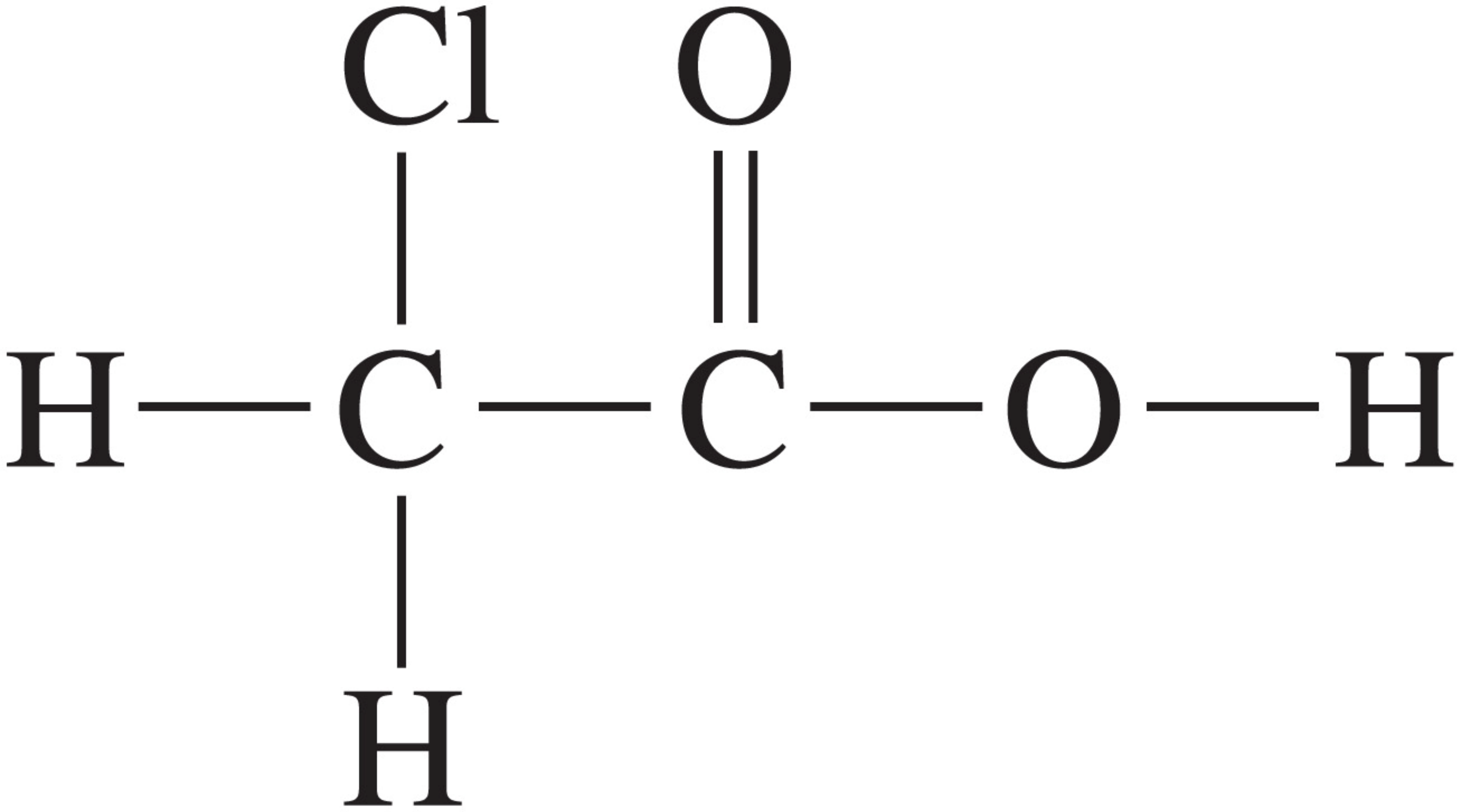




$K_{a_1} \approx 10^3$

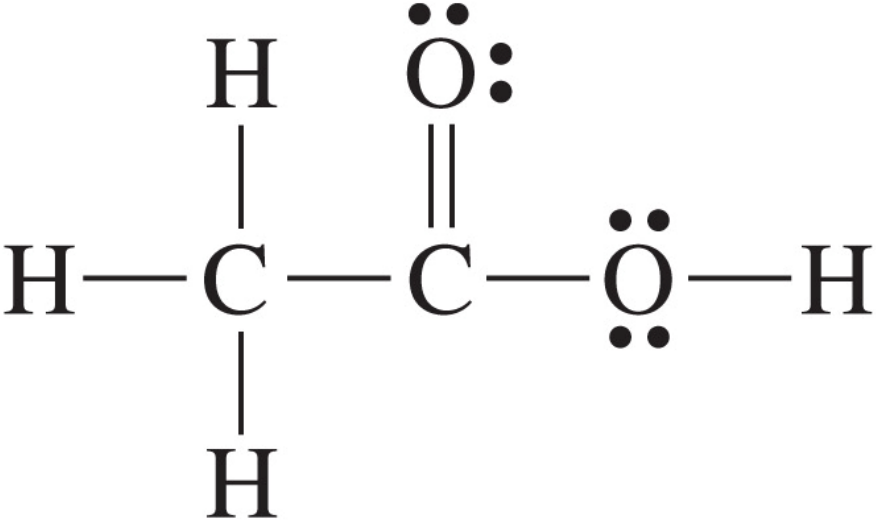


$K_{a_1} = 1.3 \times 10^{-2}$



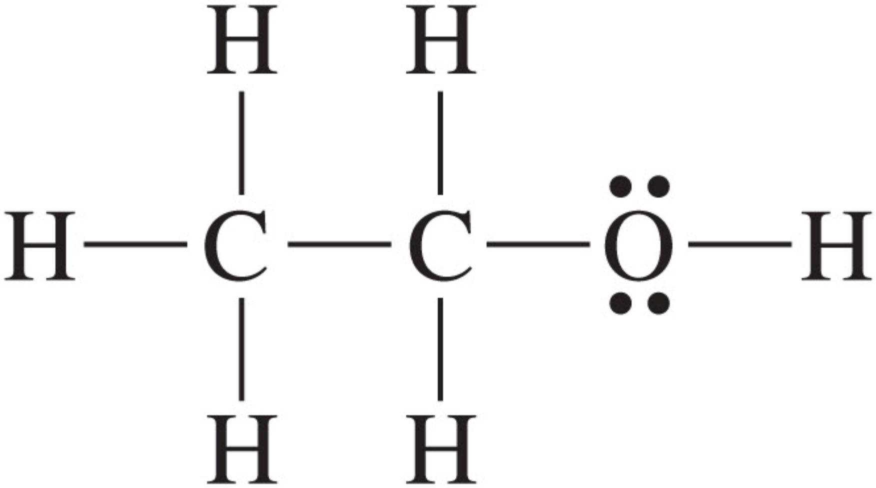
chloroacetic acid

$K_a = 1.4 \times 10^{-3}$



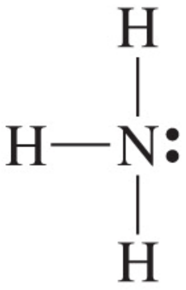
acetic acid

$K_a = 1.8 \times 10^{-5}$



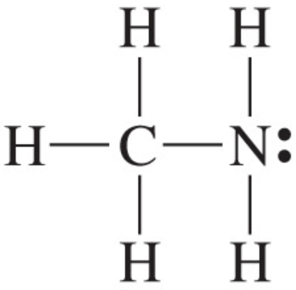
ethanol

$K_a = 1.3 \times 10^{-16}$



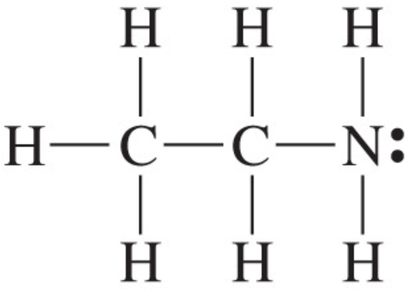
ammonia

pK_b = 4.74



methylamine

CH_3NH_2, pK_b = 3.38



ethylamine

$CH_3CH_2NH_2$, pK_b = 3.37

Grado de Disociación

$$HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$$

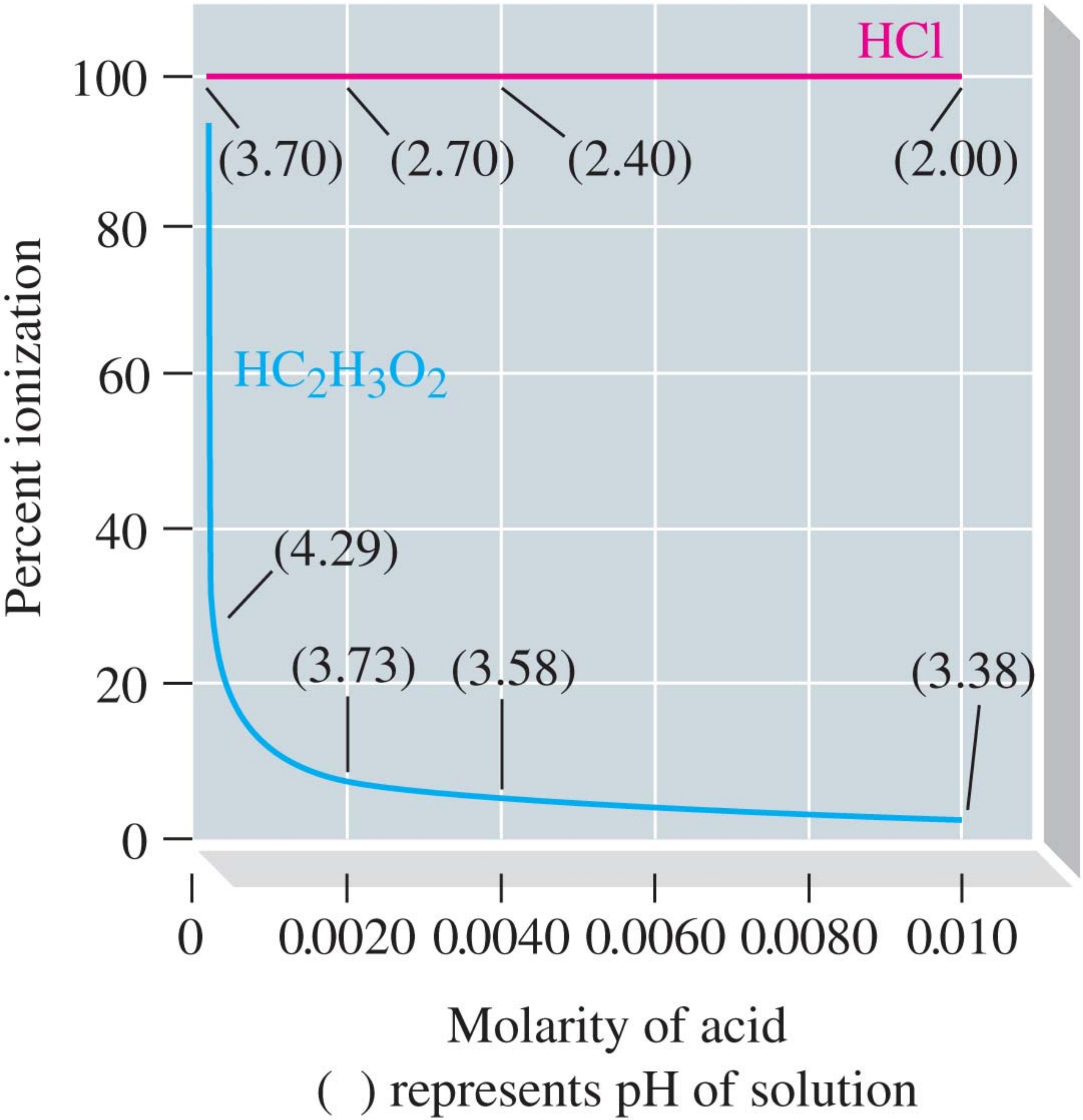
$$C_0 - x \qquad\qquad\qquad x \qquad\qquad x$$

$$K_a = \frac{\left[H_3O^+\right]\left[C_2H_3O_2^-\right]}{\left[HC_2H_3O_2\right]} = 1{,}8*10^{-5}$$

$$1{,}8*10^{-5} = \frac{x^2}{C_0 - x}$$

$$\frac{\left[C_{a,b}\right]}{K_{a,b}} > 100 \quad \Rightarrow \textit{se desprecia } x$$

a,b= ácido o base según corresponda

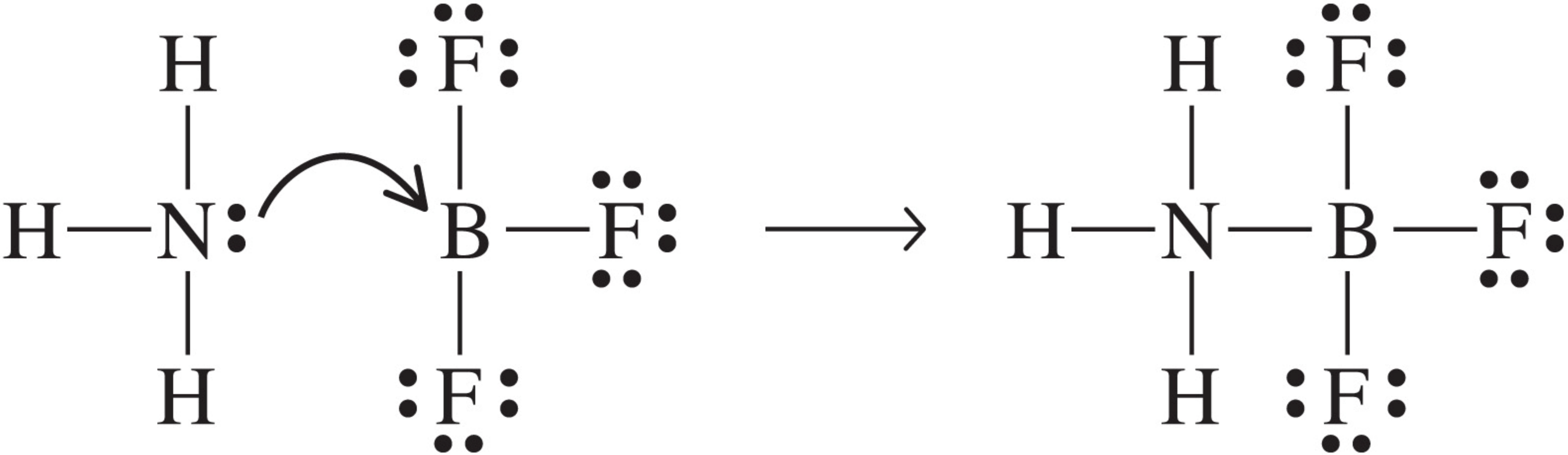

HCl
(3.70)
(2.70)
(2.40)
(2.00)
$HC_2H_3O_2$
(4.29)
(3.73)
(3.58)
(3.38)
Percent ionization
100
80
60
40
20
0
0
0.0020
0.0040
0.0060
0.0080
0.010
Molarity of acid
( ) represents pH of solution

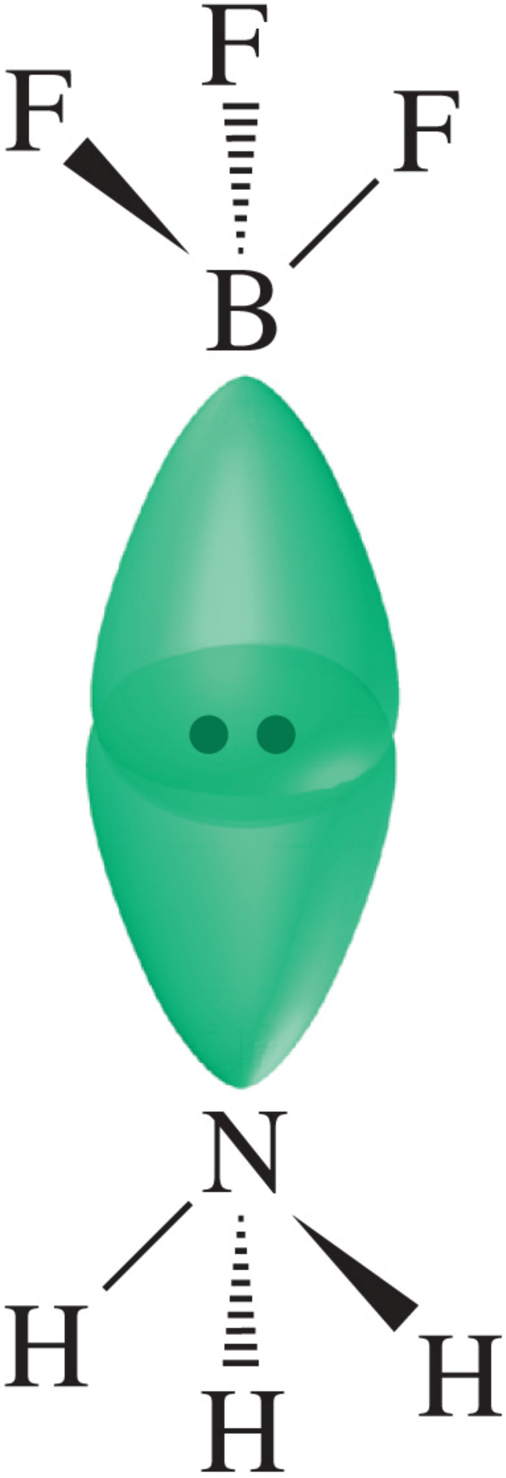
Teoría ácido-base de Lewis

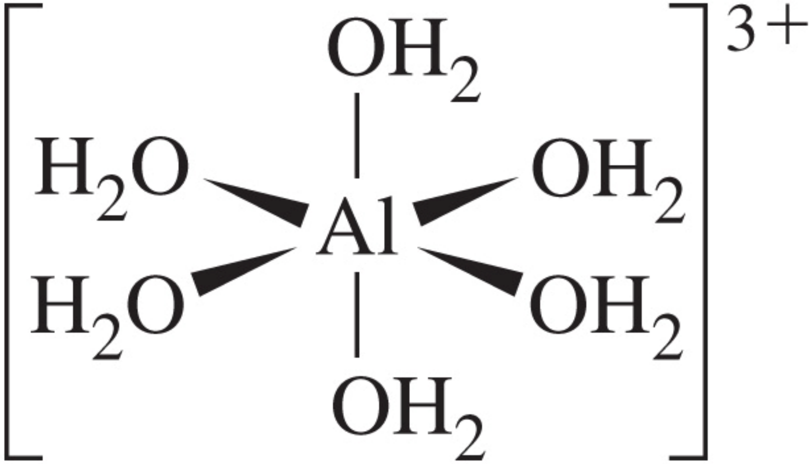
- Un ácido de Lewis es una especie (átomo, ion o molécula) que es un aceptor de pares de electrones.

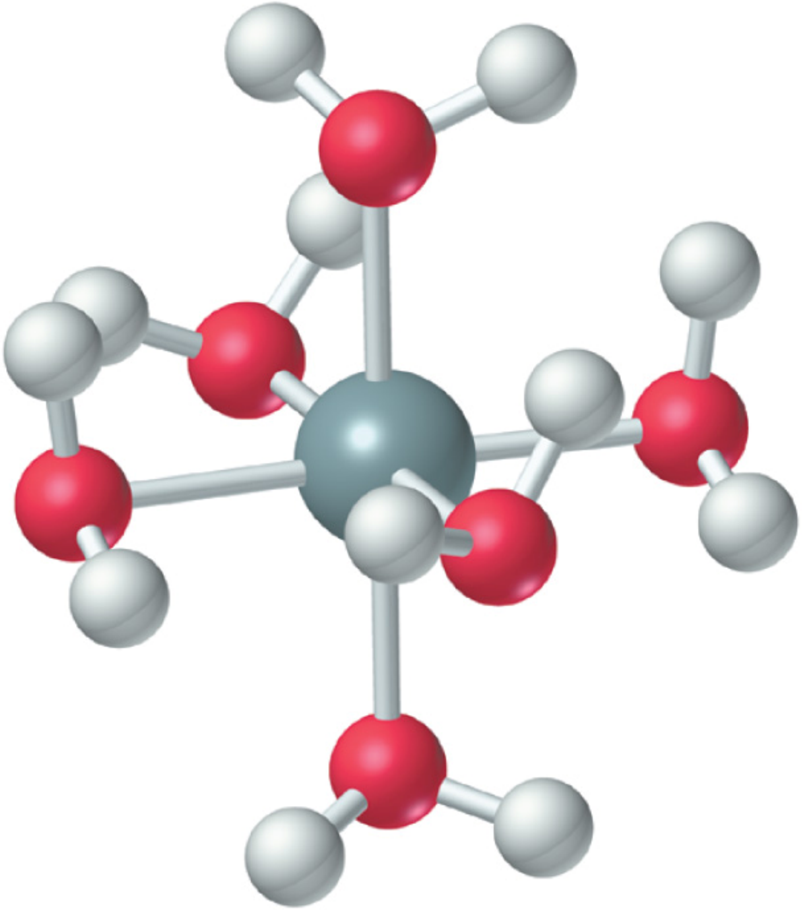
- Una base de Lewis es una especie que es dadora de pares de electrones.

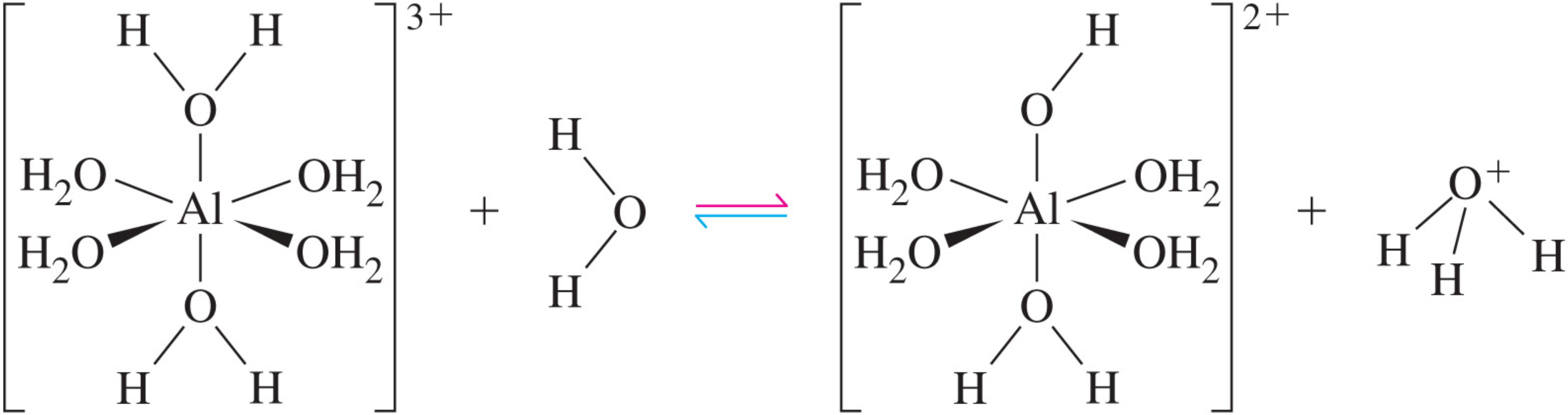
- La reacción entre ambos genera una especie unida por enlace covalente llamada aducto




F
F
F
B
N
H
H
H


3+
OH2
H2O
Al
OH2
H2O
OH2
OH2





$$Al(H_2O)_6^{3+} + H_2O \rightleftharpoons [Al(OH)(H_2O)_5]^{2+} + H_3O^+$$