

ACIDS

- 1. are proton donors
- 2. are H^+ ions
- 3. are H^+ ions

$pH = -\log K_a$
 $pK_a = -\log K_a$

$K_a \times K_b = K_w = 1 \times 10^{-14}$ at 25°C

$pK_a + pK_b = pK_w = 14$

strong acids



strong bases



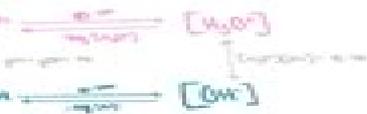
ACIDS	STRONG	WEAK
BASES	STRONG	WEAK

ACID-BASE



Kinetic Argument

A perturbation to equilibrium changes the relative values of RATE_{forward} and RATE_{reverse}.
 $K_a = K_b \rightarrow$ shift to R
 $K_a > K_b \rightarrow$ shift to R
 $K_a < K_b \rightarrow$ shift to L



$pH + pOH = 14$
 $pH = 14 + \log C_{OH^-}$

$pH = pOH = 7$
 $pH = 14 + \log C_{OH^-}$

STRONG - UNSTABLE / REACTIVE
WEAK - STABLE / UNREACTIVE
 REACTIONS ALWAYS PROCEED TO THE WEAKER ACID/BASE

↓ conjugate acid = stronger acid

ACID	$C_6H_5CO_2H$	CH_3CO_2H	HNO_3	HCl
pK_a	4.2	4.8	-1.3	-7
BASE	$C_6H_5CO_2^-$	$CH_3CO_2^-$	NO_3^-	Cl^-
pK_b	9.8	9.2	5.3	13

Thermodynamic Argument

A perturbation to equilibrium changes the reaction quotient $Q = [P][D]/[R][L]$.
 equilibrium constant
 $Q > K \rightarrow$ shift to L
 $Q < K \rightarrow$ shift to R

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

$pH = 14 + \log (C_{OH^-} \times K_b)$

$pOH = -\log C_{OH^-} + \log K_b$

$pH = pK_a + \log \left[\frac{[A^-]}{[HA]} \right]$

$pH = pK_a = 7$
 $pH > pK_a = 8$
 $pH < pK_a = 6$

$pH > 7 = \text{basic}$
 $pH < 7 = \text{acidic}$

$pH = pOH$
 if $pH = pK_a$
 $pH = pK_b$
 if $pH = 7$

Acid-Base Strength Effects
 The more basic the base, the weaker the acid. The more acidic the acid, the stronger the base. (Stronger acid has a weaker base than it.)

Acid-Base Strength Effects
 The larger the K_a , the better the acid. The smaller the K_b , the better the base. (Stronger acid has a weaker base than it.)

More acidic when having the negative charge. Thus, a more basic molecule than it is a weaker base than it.

the more electronegative, the more basic

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