5/15/20

Arrhenius definition

acid - dissociates to form H^+

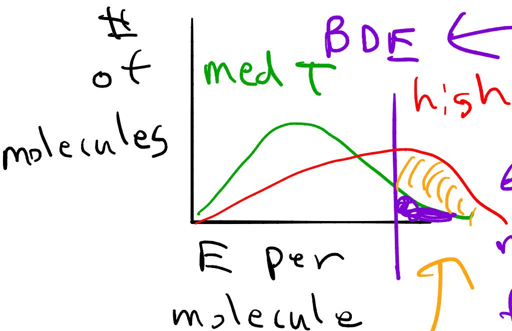
base - dissociates to form ^-OH

Auto-ionization of water

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

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

@ 25°C $\quad = 1.0 \times 10^{-14}$



BDE ← Bond dissociation energy

At any temperature, some fraction of molecules will be able to dissociate.

At higher T, more molecules can dissociate, so K_w increases.

$$pH \equiv -\log_{10}[H^+]$$

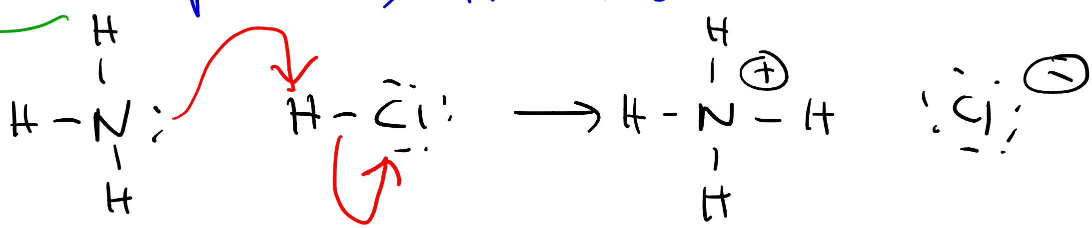
Brønsted-Lowry

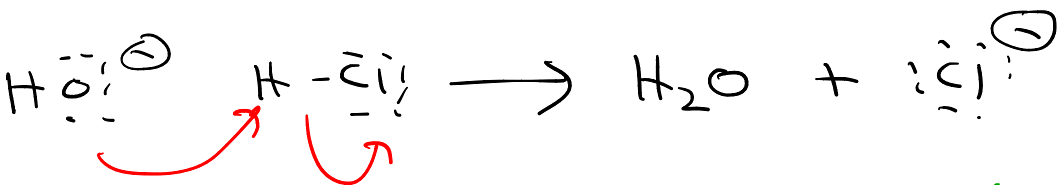
acid - dissociates to produce H^+
(same as an Arrhenius acid)
proton donor

base - reacts to accept H^+
proton acceptor

${}^{1}_{1}H \xrightarrow{-e^-} {}^{1}_{1}p^{+}$ proton

Example → Ammonia





Since these substances react with acids in the same way, that's why under the BL definition they are both considered bases.

The K_a of acetic acid is 1.76×10^{-5}. What is the pH of a 0.25 M aqueous solution of acetic acid?

L3

Auto-ionization of an acid

$$1\,HA_{(aq)} \rightleftharpoons 1\,H^+_{(aq)} + 1\,A^-_{(aq)}$$

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

acid dissociation constant

$K_a >> 1$ Strong acid (extensive dissociation)

$K_a << 1$ weak acid (minimal dissociation)

ICE problem – Start off assuming the acid has not dissociated

$$Q = \frac{0.0}{0.25} = 0$$

	[HA]	[H+]	[A-]
Initial ($Q<K$) concentration	0.25	0	0
		not 0, but small enough it can usually be ignored	no rxn yet
Change ($Q<K$)	$-x$	$+x$	$+x$
End	$0.25-x$	x	x

	HA	H^+	A^-
I	0.25	0	0
C	$-x$	$+x$	$+x$
E	$0.25-x$	x	x

$$K_a = \frac{[H^+][A^-]}{[HA]} = 1.76\times10^{-5} = \frac{(x)(x)}{0.25-x}$$

$$1.76\times10^{-5} = \frac{x^2}{0.25-x}$$

$0.25 - x \approx 0.25$

$$x^2 + 1.76\times10^{-5}x - 4.4\times10^{-6} = 0$$

Since acetic acid is a weak acid, is x small enough it can be ignored? ⟶ Try it and see!

$$1.76\times10^{-5} = \frac{x^2}{0.25}$$

$$x^2 = 4.4\times10^{-6}$$

$$x = 2.10\times10^{-3}\ M = [H^+]$$

Less than 1% of HA concentration ⟶ valid to ignore x

$$pH = -\log_{10}[2.10\times10^{-3}] = 2.68$$