

Acids

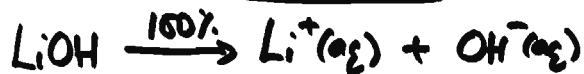
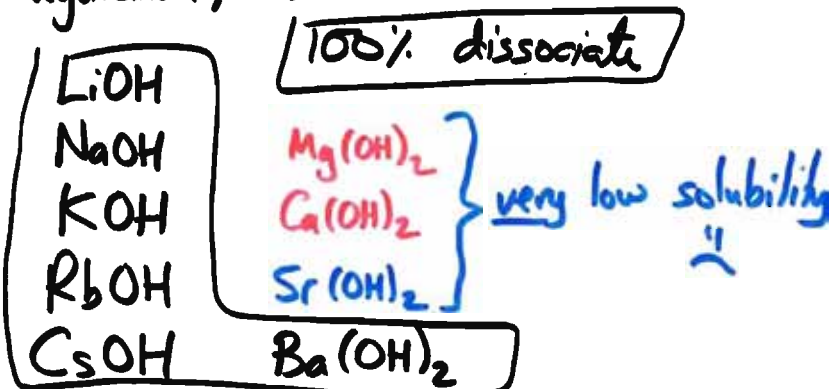
know 6 strong ones!
100% dissociate.

weak ones: all others!
<100% dissociation!

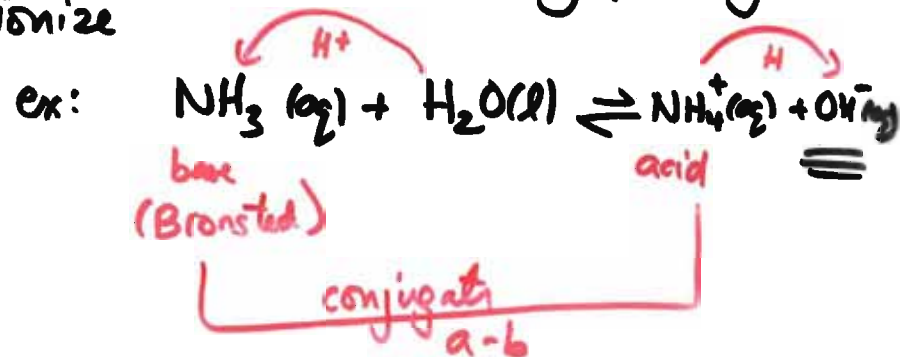


BASES

Strong bases... typically alkali metal hydroxides, as well as $\text{Ba}(\text{OH})_2$.

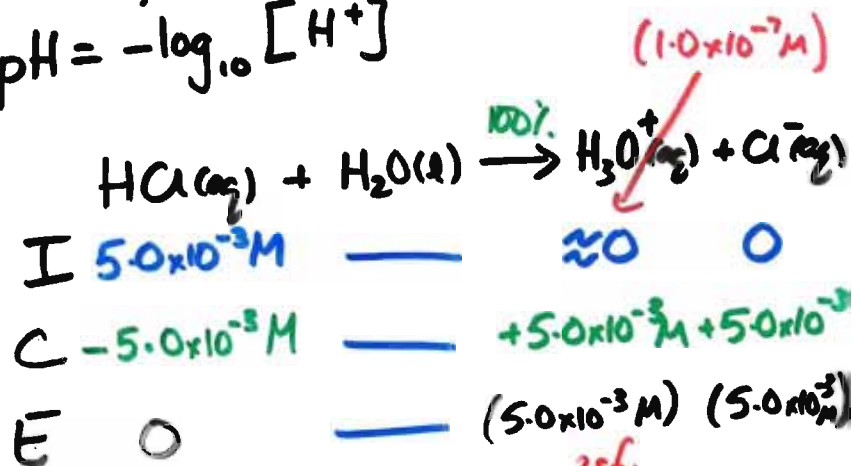


Most bases are weak... only partially ionize



What's pH of 5.0mM $\text{HCl}(\text{aq})$?

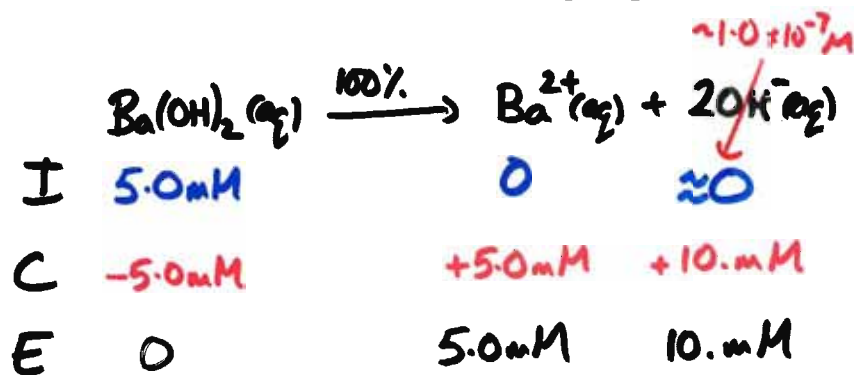
$\text{pH} = -\log_{10} [\text{H}^+]$



$\text{pH} = -\log [\text{H}^+] = -\log [5.0 \times 10^{-3}]$
 $= 2.30$

2sf.
2dp

ex: What's pH of 5.0mM Ba(OH)₂(aq)?
@25°C



$pH = -\log[H^+]$?

$pOH = -\log[OH^-] = -\log_{10}[10. \times 10^{-3}]$
= 2.00

$pH + pOH = 14.00$

$\Rightarrow pH = 12.00$

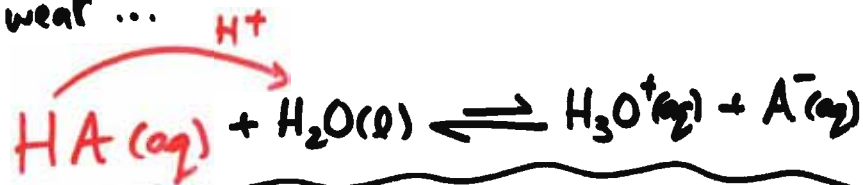
$[H^+] = 10^{-pH} = 10^{-12.00} = 1.0 \times 10^{-12} M$

Basic if $pH \gg (aq, 25^\circ C)$

Weak Acids and Ka

↳ Acid dissociation constant.

Most acids are weak ...



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

large K_a = 'stronger' acid
small K_a = weaker acid

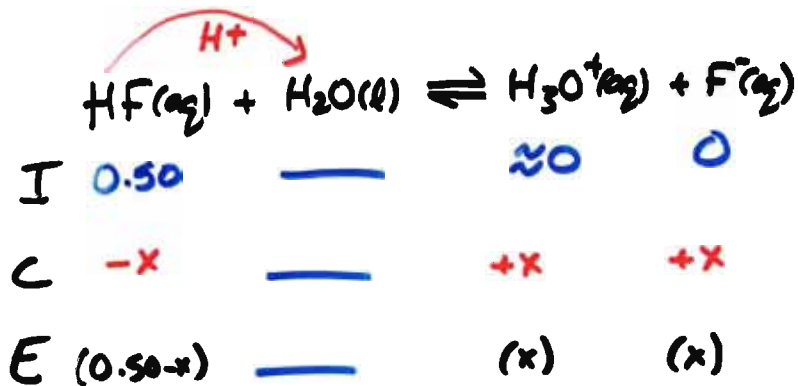
ex: HF
formic acid \rightarrow HCO₂H
acetic acid \rightarrow CH₃CO₂H
phenol or carboic acid \rightarrow C₆H₅OH

$K_a = 7.1 \times 10^{-4}$
 $= 1.7 \times 10^{-4}$
 $= 1.8 \times 10^{-5}$
 $= 1.3 \times 10^{-10}$

STRONGER
↓
WEAKER

Let's calculate pH of 0.50M HF(aq) @25°C.

$K_a = 7.1 \times 10^{-4}$ (weak acid).



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HA}]}$$

$$7.1 \times 10^{-4} = \frac{(x)(x)}{0.50-x}$$

Perhaps... $0.50 \gg x$ or $x \ll 0.50$

then... $7.1 \times 10^{-4} \approx \frac{x^2}{0.50}$

$$\sqrt{7.1 \times 10^{-4} \times 0.50} \approx \sqrt{x^2} = x$$

$$\Rightarrow x \approx 0.0188$$

$$\Rightarrow [\text{H}^+] = x = 0.0188\text{M}$$

small... so perhaps assumption that $x \ll 0.50$ was valid.

5% rule
approx. is OK if there's <5% disoc.

$$\frac{[\text{H}^+]_{\text{eq}}}{[\text{HA}]_0} \times 100 < 5\%$$

$$\Rightarrow \frac{0.0188\text{M}}{0.50\text{M}} \times 100 = 3.8\% \quad \checkmark$$

$$\text{pH} = -\log[\text{H}^+] = -\log[0.0188\text{M}] = 2.08$$