

pH

Log scale.
Useful when dealing with very small
or very large number (big ranges of numbers)
every "pH" unit is 10x larger or smaller $[H^+]$

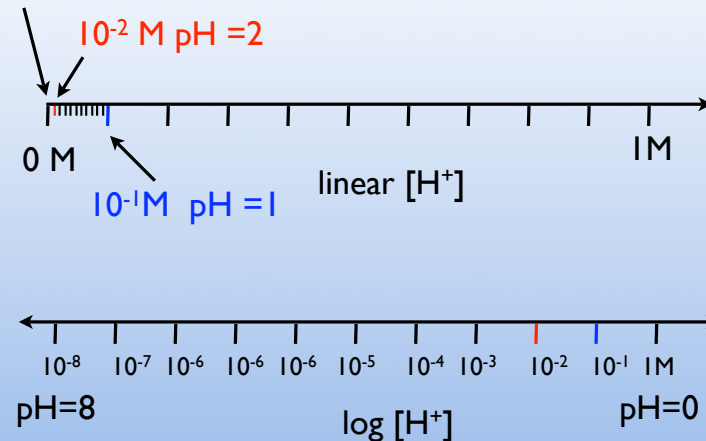
$$pH = -\log[H^+]$$

$$pH = 13 \\ [H^+] = 10^{-13}$$

$$pH = 7 \\ [H^+] = 10^{-7}$$

$$pH = 2 \\ [H^+] = 10^{-2}$$

pH 3-14

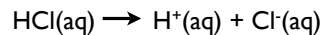


Strong Acids and Bases

"Strong" means one thing

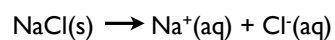
The substance dissociates 100% in water

Strong Acid



$$K_a = \frac{[H^+][Cl^-]}{[HCl]} \approx \infty$$

Strong Electrolyte



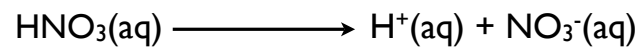
$$K_{sp} = [Na^+][Cl^-] \approx \infty$$

Strong Acids

HCl	Hydrochloric
HBr	Hydrobromic
HI	Hydroiodic
HClO ₄	Perchloric
HClO ₃	Chloric
H ₂ SO ₄	Sulfuric
HNO ₃	Nitric

All Dissociate 100%

What is the pH of a 0.1 M solution of Nitric Acid



0.1 M acid makes a solution with $[\text{H}^+] = 0.1\text{M}$

$$\text{pH} = -\log(0.1) = 1$$

What is the pH of a 0.5M solution of HBr?

A. 0.5

B. 1

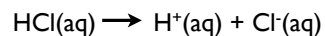
C. 0.3

D. 0

E. 12

$$[\text{H}^+] = 0.5 \quad 0 < \text{pH} < 1$$

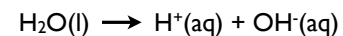
We can ignore the conjugate base of a strong acid



$$K_a = \frac{[\text{H}^+][\text{Cl}^-]}{[\text{HCl}]} \approx \infty$$

equilibrium constant is so large,
even if we add Cl^- the shift back to
 HCl will be negligible

For this reaction which has a higher entropy?

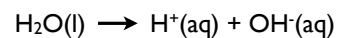


A. the products

B. the reactants

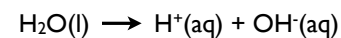
C. they are the same

For this reaction which has a lower enthalpy?



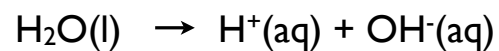
- A. the products
- B. the reactants
- C. they are the same

For this reaction which has a lower free energy?



- A. the products
- B. the reactants
- C. they are the same

Liquid Water
will spontaneously dissociate to a small extent



$$K = \frac{[\text{H}^+][\text{OH}^-]}{1}$$

$$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$$

Pure Water

	H ⁺	OH ⁻
I	0	0
C	+x	+x
E	+x	+x

$$K_w = 10^{-14} = [\text{H}^+][\text{OH}^-] = (x)(x)$$

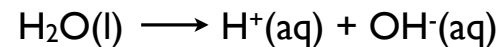
$$x = 10^{-7} \quad [\text{H}^+] = [\text{OH}^-] = 10^{-7}$$

pH of pure water at 25°C

$$x = 10^{-7} \quad [H^+] = [OH^-] = 10^{-7}$$

$$pH = -\log[H^+] = -\log(10^{-7}) = 7$$

Neutral	Acidic	Basic
$[H^+] = [OH^-]$	$[H^+] > [OH^-]$	$[H^+] < [OH^-]$
at 25°C	at 25°C	at 25°C
pH = 7 pOH = 7	pH < 7 pOH > 7	pH > 7 pOH > 7



This reaction is endothermic.
Given that information what do you think
the pH is for pure water at 60°C?

- A. 6.5 ←
- B. 7
- C. 7.5

If pure water has a pH = 6.5 at 60°C is it Acidic?

A. Yes

B. No ←
[H⁺] = [OH⁻] its neutral $K_w = 9 \times 10^{-14}$

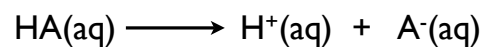
TABLE 7.1 Various Ways to Describe Acid Strength

Property	Strong Acid	Weak Acid
K_a value	K_a is large	K_a is small
Position of the dissociation equilibrium	Far to the right	Far to the left
Equilibrium concentration of H^+ compared with original concentration of HA	$[H^+] \approx [HA]_0$	$[H^+] \ll [HA]_0$
Strength of conjugate base compared with that of water	A^- much weaker base than H_2O	A^- much stronger base than H_2O

Let's look at three possible solutions

Weak Acid
Conjugate Base of the Weak Acid
Weak Acid + Conjugate Base

Weak Acid



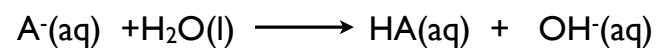
	HA	H ⁺	A ⁻
I	C	0 <small>← really 10⁻⁷</small>	0
C	-x	+x	+x
E	C-x	+x	+x

$$K_a = \frac{[\text{H}^{\text{+}}][\text{A}^{-}]}{[\text{HA}]} = \frac{(x)(x)}{C-x}$$

assuming $x \ll C$

$$x \sim \sqrt{K_a C}$$

Weak Base

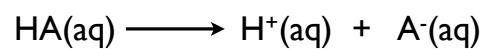


$$K_b = \frac{[\text{HA}][\text{OH}^{-}]}{[\text{A}^{-}]}$$

identical result as before (same assumptions)

$$[\text{OH}^{-}] = \sqrt{K_b C}$$

Buffer Both HA and A⁻



	HA	H ⁺	A ⁻
I	C _{HA}	0 <small>← really 10⁻⁷</small>	C _{A⁻}
C	-x	+x	+x
E	C _{HA} -x	+x	C _{A⁻} + x

$$K_a = \frac{[\text{H}^{\text{+}}][\text{A}^{-}]}{[\text{HA}]} = \frac{(x)(C_{\text{A}^{-}} + x)}{C_{\text{HA}} - x} = \frac{(x)(C_{\text{A}^{-}})}{C_{\text{HA}}}$$

assuming $x \ll C$