pH

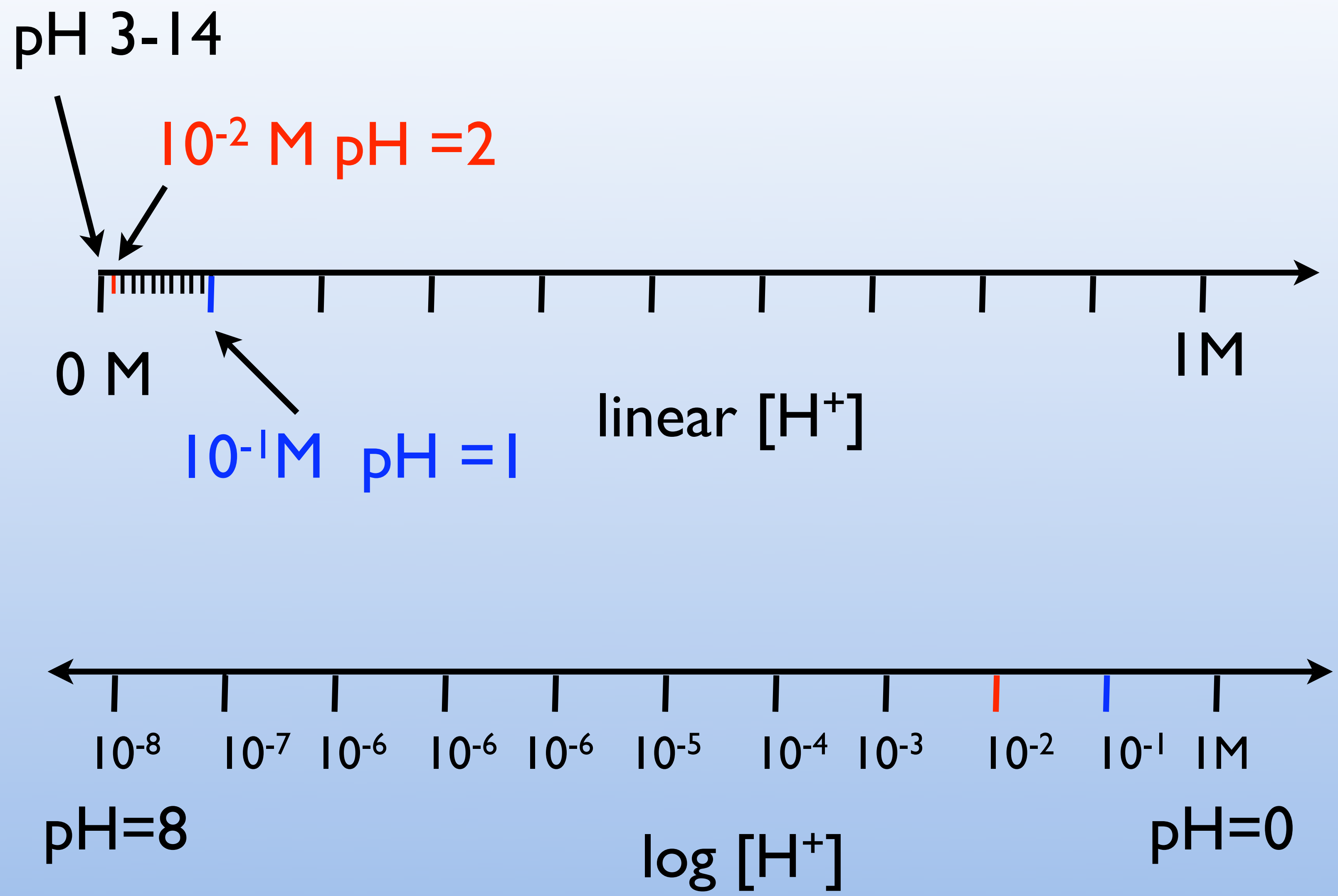
Log scale.
Useful when dealing with very small
or very large number (big ranges of numbers)
every "pH" unit is 10x larger or smaller $[H^+]$

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

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

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

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



pH 3-14
10-2 M pH =2
0 M
1M
10-1M pH =1
linear [H+]
10-8
10-7
10-6
10-6
10-6
10-5
10-4
10-3
10-2
10-1
1M
pH=8
log [H+]
pH=0

Strong Acids and Bases

"Strong" means one thing

The substance dissociates 100% in water

Strong Acid

$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$

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

Strong Electrolyte

$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$

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

Strong Acids

HCl	Hydrochloric
HBr	Hydrobromic
HI	Hydroiodic
$HClO_4$	Perchloric
$HClO_3$	Chloric
H_2SO_4	Sulfuric
HNO_3	Nitric

All Dissociate 100%

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

$$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$

0.1 M acid makes a solution with $[H^+] = 0.1M$

$$pH = -\log(0.1) = 1$$

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

A. 0.5

B. 1

C. 0.3

D. 0

E. 12

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

We can ignore the conjugate base of a strong acid

$$HCl(aq) \longrightarrow H^{+}(aq) + Cl^{-}(aq)$$

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

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?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

For this reaction which has a lower enthalpy?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

For this reaction which has a lower free energy?

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

A. the products

B. the reactants

C. they are the same

Liquid Water
will spontaneously dissociate to a small extent

$$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$$

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

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

Pure Water

	H^+	OH^-
I	0	0
C	+x	+x
E	+x	+x

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

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

pH of pure water at 25°C

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

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

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

$$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$$

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

$$HA(aq) \longrightarrow H^+(aq) + A^-(aq)$$

	HA	H^+	A^-
I	C	0 (really 10^{-7})	0
C	-x	+x	+x
E	C-x	+x	+x

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

assuming x << C

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

Weak Base

$$A^-(aq) + H_2O(l) \longrightarrow HA(aq) + OH^-(aq)$$

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

identical result as before (same assumptions)

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

Buffer Both HA and A-

$$HA(aq) \longrightarrow H^+(aq) + A^-(aq)$$

	HA	H^+	A^-
		really 10^{-7}	
I	C_{HA}	O	C_{A-}
C	-x	+x	+x
E	C_{HA}-x	+x	C_{A-} + x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(C_{A-}+x)}{C_{HA}-x} = \frac{(x)(C_A)}{C_{HA}}$$ assuming x << C