Acids + Bases

- Theories of Acids/Bases
  - Operational Def'n
  - Arrhenius Definition

Acids $HCl \rightarrow H^+ + Cl^-$

Bases $NaOH \rightarrow Na + OH^-$

- Modified Arrhenius

Acids $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

- Bronstead-Lowry

Acids are Proton (H^+) Donor

Bases are Proton Acceptor

4. Which represents the reaction of a Brønsted-Lowry base?

(A) $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

(B) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

(C) $CH_3COOH(aq) + C_2H_5OH(\ell) \rightleftharpoons CH_3COOCH_2CH_3(\ell) + H_2O(\ell)$

(D) $H_2C_2O_4(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HC_2O_4^-(aq)$

Acid Strength/Base Strength

Acid Strength $\neq$ Concentration

ie: 12.0 M Acetic Acid is a Concentrated Weak Acid.

Strong + Weak Acids will behave differently.
0.1 M 0.1 M

15. If their concentrations are equal, which solution will show the highest electrical conductivity?

(A) H_3BO_3 (aq)
(B) HCl (aq)
(C) H_2CO_3 (aq)
(D) $H_2C_2O_4$ (aq)

$HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$

$H_2CO_3 + H_2O \overset{0.21\%}{\rightleftharpoons} H_3O^+ + HCO_3^-$

15. Which properties best describe a 0.10 mol/L solution with the highest pH?

	Reaction with active metal	Electrical conductivity	Litmus
(A)	moderate	moderate	red W.A
(B)	none	high	blue S.B
(C)	none	moderate	blue W.B
(D)	vigorous	high	red S.A

$NaOH_{(s)} \xrightarrow{100\%} Na^+ + OH^-$

$NH_3 + H_2O \overset{1\%}{\rightleftharpoons} NH_4^+ + OH^-$

Predicting B/L Reactions

Acids	Bases	Spectators
HCO_3^-	HCO_3^- S.B	Na^+
NH_4^+ SA	NO_2^-	
H_2O	H_2O	

Reactants Favored

$\underbrace{\underset{A}{NH_{4(aq)}^+} + \underset{B}{HCO_{3(aq)}^-}}_{weaker} \rightleftharpoons \underbrace{\underset{CB}{NH_{3(aq)}} + \underset{C.A.}{H_2CO_{3(aq)}}}_{Stronger}$

2% 52.(a) Determine the Brønsted-Lowry acid-base neutralization reaction that occurs between $NaHCO_3(aq)$ and $NH_4NO_2(aq)$.

pH Scale

- logarithmic

18. The pH of a sample of rainwater has changed from 5.4 to 4.4. What is true of the rainwater?

	$[H_3O^+]$	Factor
(A)	decrease	1
(B)	decrease	10
(C)	increase	1
(D)	increase	10

$$\boxed{\begin{aligned} pH &= 5.4 \\ [H_3O^+] &= 10^{-5.4} \\ &= 0.00000398M \end{aligned}}$$

$pH = 4.4$

$[H_3O^+] = 10^{-4.4}$

$0.0000398\ M$

- pH scale basedon K_w @ 25°C

Heat + $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

$$K_w = \frac{[H_3O^+][OH^-]}{(1)}$$

$K_w = 1.00 \times 10^{-14}$

Relevant Formulas

$pH = -\log[H_3O^+]$ (Provided)

$pOH = -\log[OH^-]$
$[H_3O^+] = 10^{-pH}$
$[OH^-] = 10^{-pOH}$
} not provided

$pH + pOH = 14$

$K_w = K_A \cdot K_B$

Strong Acids/Bases

*Always check if an acid is strong or weak!

For strong Acids: $[H_3O^+] = [Acid]$

(ionizes 100%)

Ex a)* Calculate the pH of 30.0ml of 0.025 M HBr.

b) If the solution were diluted to 300.0ml, find the new pH.

a) Since HBr is Strong $[HBr] = [H_3O^+]$

$pH = -\log(H_3O^+)$

$= -\log(0.025) = 1.60$

b) $C_iV_i = C_fV_f$ $\quad C_f = \dfrac{C_iV_i}{V_f}$

$= \dfrac{(0.025M)(0.0300L)}{(0.3000L)}$

$= 0.0025M$

$[H_3O^+] = 0.0025M$

$pH = -\log(0.0025M) = \underline{2.60}$

Strong Bases

Strong Bases react 100% to form OH^-, but:

$NaOH \rightarrow Na^+ + OH^-$ $\quad [NaOH] = [OH^-]$

whereas,

$Ba(OH)_2 \rightarrow Ba^{2+} + \underline{2}OH^-$ $\quad [Ba(OH)_2] \times 2 = [OH^-]$

.1M $\quad$ 0.2M

ex Find $[OH^-]$ and pOH of a 0.040M $Al(OH)_3$ solution.

$Al(OH)_3 \longrightarrow Al^{3+} + \underline{3}OH^-_{(aq)}$

0.040M $\quad$ 0.040M $\quad$ $3 \times 0.040M = 0.12\ M = [OH^-]$

$pOH = -\log[0.12M]$

$\boxed{= 0.92}$

Weak Acid/Base Problems

- Is it weak/strong?
- Acid or Base?
- Solve for K or solve for x? (pH, pOH, $[H_3O^+]$, $[OH^-]$)
- Rule of 500? (If solving x)

In General,

I	0.1	—	0	0
C	-x	–	+x	+x
E	0.1-x	–	x	x

% (c) Calculate the pH of a 0.25mol/L solution of HA(aq), if $K_a = 3.6 \times 10^{-7}$.

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons A^-_{(aq)} + H_3O^+$$

	$HA_{(aq)}$	$H_2O_{(l)}$	$A^-_{(aq)}$	H_3O^+
I	0.25 M	—	0	~0 M
C	-x		+x	+x
E	0.25M - x		x	x

$$K = \frac{x^2}{(0.25 - x)}$$

Rule of 500:

$$\frac{0.25M}{(3.6\times10^{-7})} > \text{Big} > \underline{500}$$

$$3.6\times10^{-7} = \frac{x^2}{(0.25)}$$

$$(3.6\times10^{-7})(0.25) = x^2$$

$$\sqrt{9.0\times10^{-8}} = \sqrt{x^2}$$

$$3.0\times10^{-4} = x = [H_3O^+]$$

$$pH = -\log[H_3O^+]$$

$$= -\log[3.0\times10^{-4}]$$

$$= 3.\boxed{52} \text{ 2sf}$$

If you need K_B:

$$K_W = K_A \cdot K_B$$

1.00×10^{-14}

ie Find K_B for NO_2^-

HNO_2 $K_A = 7.2\times10^{-4}$

$$K_B = \frac{K_W}{K_A} = \frac{(1.00\times10^{-14})}{(7.2\times10^{-4})} = 1.4\times10^{-11}$$

% (b) Calculate the concentration of a 10.00 mL CH_3COOH solution using the data provided from the standardization with a 0.1404 mol/L NaOH solution.

Burette	Trial 1	Trial 2	Trial 3
Final (mL)	16.90	32.02	47.18
Initial (mL)	1.35	16.90	32.02
Volume NaOH used (mL)	~~15.55~~	15.12	15.16

$$CH_3COOH + \cancel{Na}OH \rightarrow H_2O + CH_3COO^-$$

$V = 10.00 mL$
$c = ?$

$c = 0.1404M$

$$V = \frac{15.12 + 15.16}{2} = 15.14 mL$$

$$n_{NaOH} = c \times V$$

$$= \frac{0.1404 mol}{\cancel{L}} \times 0.01514 \cancel{L}$$

$$= 0.002125_5 \text{ mol}$$

$$n_{CH_3COOH} = 0.002125_5 \text{ mol NaOH} \times \frac{1 \text{ mol } CH_3COOH}{1 \text{ mol NaOH}}$$

$$= 0.002125_5 \text{ mol } CH_3COOH$$

$$[CH_3COOH] = \frac{0.002125_5 mol}{0.01000 L} = 0.2126 mol/L$$

Indicators

- Table provided
- Indicators are weak Acids

$$HIn_{(aq)} + H_2O_{(l)} \rightleftharpoons In^-_{(aq)} + \boxed{H_3O^+_{(aq)}} \uparrow$$

Orange Purple

Buffers

: Weak Acid + Conjugate Base

- Resist pH change
- Equilibrium is maintained by LCP.