

Acids + Bases

- Theories of Acids/Bases
  - operational Def'n
  - Arrhenius Definition
    - Acids  $HCl \rightarrow H^+ + Cl^-$
    - Bases  $NaOH \rightarrow Na + OH^-$
  - Mod. Fied 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 Bronsted-Lowry base?
- (A)  $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$   
 (B)  $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$   
 (C)  $CH_3COOH(aq) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$   
 (D)  $H_2C_2O_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HC_2O_4^-(aq)$

Acid Strength/Base Strength

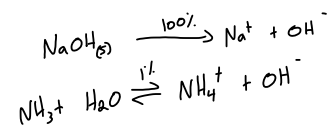
Acid Strength  $\neq$  Concentration  
 12.0M Acetic Acid is a Concentrated Weak Acid.

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

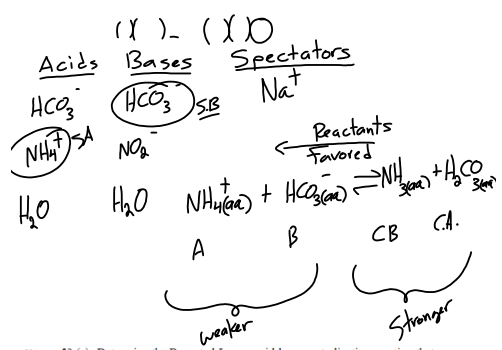
15. If their concentrations are equal, which solution will show the highest electrical conductivity?
- (A)  $H_3BO_3(aq)$   
 (B)  $HCl(aq)$   
 (C)  $H_2CO_3(aq)$   
 (D)  $H_2C_2O_4(aq)$
- $HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$   
 $H_2CO_3 + H_2O \rightleftharpoons^{0.2\%} H_3O^+ + HCO_3^-$

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
<input checked="" type="radio"/> (B)	none	high	blue S.B
(C)	none	moderate	blue W.B
(D)	vigorous	high	red S.A



Predicting B/L Reactions



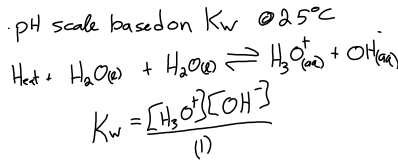
26 52. (a) Determine the Bronsted-Lowry acid-base neutralization reaction that occurs between  $NaHCO_3(aq)$  and  $NH_4NO_2(aq)$ .

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 <sub>3</sub> O <sup>+</sup> ]	Factor
(A)	decrease	1
(B)	decrease	10
(C)	increase	1
(D)	increase	10

$pH = 5.4 \rightarrow -5.4$   
 $[H_3O^+] = 10^{-5.4} = 0.0000398 M$   
 $pH = 4.4 \rightarrow -4.4$   
 $[H_3O^+] = 10^{-4.4} = 0.0000398 M$



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

Relevant Formulas  
 $pH = -\log[H_3O^+]$  (Provided)  
 $pOH = -\log[OH^-]$  } not provided  
 $[H_3O^+] = 10^{-pH}$   
 $[OH^-] = 10^{-pOH}$   
 $pH + pOH = 14$   
 $K_w = K_A \cdot K_B$

Strong Acids/Bases

\*Always check if an acid is strong or weak!

For strong Acids:  $[H_3O^+] = [Acid]$   
 (ionizes 100%)

Ex: Calculate the pH of 30.0ml of 0.025 M HBr.  
 IF the solution were diluted to 300.0ml, find the new pH.

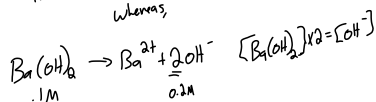
Since HBr is strong  $[HBr] = [H_3O^+]$   
 $pH = -\log(H_3O^+)$   
 $= -\log(0.025) = 1.60$

$C_i V_i = C_f V_f \quad C_f = \frac{C_i V_i}{V_f}$   
 $= \frac{(0.025M)(0.0300L)}{(0.3000L)}$   
 $= 0.0025M$

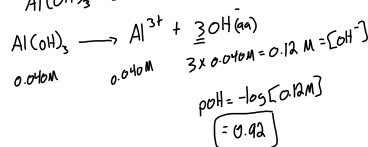
$[H_3O^+] = 0.0025M$   
 $pH = -\log(0.0025M) = 2.60$

Strong Bases

Strong Bases react 100% to form OH<sup>-</sup>,  
 but:  
 $NaOH \rightarrow Na^+ + OH^- \quad [NaOH] = [OH^-]$   
 whereas,



ex Find [OH<sup>-</sup>] and pOH of a 0.040M Al(OH)<sub>3</sub> solution.



Weak Acid/Base Problems

- Is it weak/strong?
- Acid or Base?
- Solve for K or solve for x?
- Rule of 500? (if solving x)  $\left( \frac{[H_3O^+]}{[OH^-]} \right)$

In General,

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

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

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(l)$$

I	0.25 M	-	0	~0 M
C	-x		+x	+x
E	0.25 M - x		x	x

$$K = \frac{x^2}{(0.25-x)} \xrightarrow{\text{Rule of 500:}} \frac{0.25 M}{(3.6 \times 10^{-7})} > 500$$

$$3.6 \times 10^{-7} = \frac{x^2}{(0.25)}$$

$$(3.6 \times 10^{-7})(0.25) = x^2$$

$$\sqrt{9.0 \times 10^{-8}} = \sqrt{x^2}$$

$$3.0 \times 10^{-4} = x = [H_3O^+]$$

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

$$= -\log[3.0 \times 10^{-4}]$$

$$= 3.52 \text{ approx}$$

If you need  $K_B$ :

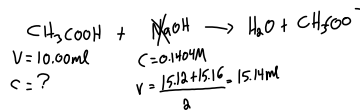
$$K_w = K_A \cdot K_B \quad \text{Find } K_B \text{ for } NO_2^-$$

$1.0 \times 10^{-14}$   $\swarrow$   $HNO_2 \quad K_A = 7.2 \times 10^{-4}$

$$K_B = \frac{K_w}{K_A} = \frac{(1.00 \times 10^{-14})}{(7.2 \times 10^{-4})} = 1.4 \times 10^{-11}$$

(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)	<del>15.55</del>	15.12	15.16



$$n_{NaOH} = C \cdot V$$

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

$$= 0.002125 \text{ mol}$$

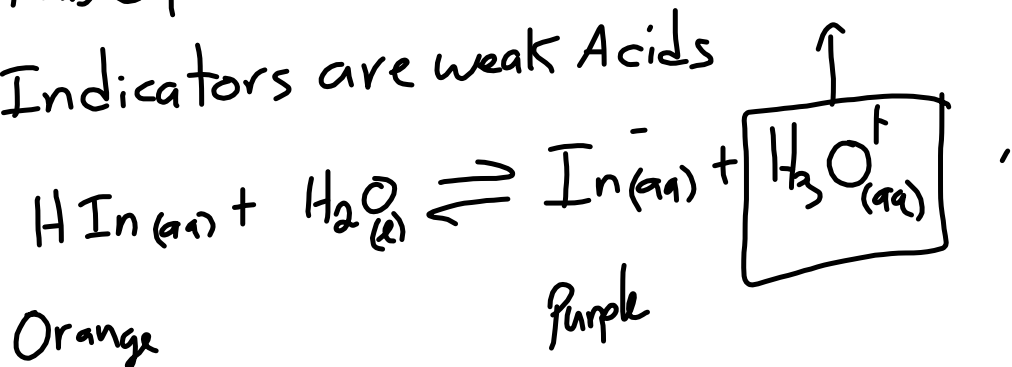
$$n_{CH_3COOH} = 0.002125 \text{ mol NaOH} \times \frac{1 \text{ mol } CH_3COOH}{1 \text{ mol NaOH}}$$

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

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

## Indicators

- Table provided
- Indicators are weak Acids



Buffers : Weak Acid + Conjugate Base

- Resist pH change
- Equilibrium is maintained by LCP.

