South Pasadena • AP Chemistry Name Groves Period ___

Ch 17 • Acids & Bases

STATION 1—CONJUGATE ACIDS & BASES

Identify the conjugate acid-base pairs in the following equation:

$$HC_2H_3O_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$$

The conjugate acid of SO_4^{2-} is HSO_4^-

The conjugate base of HCO_3^- is CO_3^{2-}

The conjugate acid of NH_3 is NH_4^+

Which substance(s) are amphiprotic / amphoteric? (H_2O) (OH^-) (NH_3) (HCO_3^-) SO_4^{2-}

(conj acid: H_3O^+, H_2O, NH_4^+, H_2CO_3, HSO_4^- — conj base: OH^-, O^{2-}, NH_2^-, CO_3^{2-}, —)

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STATION 2—pH OF ACID SOLUTIONS

Calculate the pH of a 0.150 M solution of HCl. (K_a = very large)

STRONG ACID... $[H^+] = [HCl] = .150\ M$

$pH = -\log[H^+] = -\log(.150) = 0.824$

Calculate the pH of a 0.150 M solution of HF. ($K_a = 7.2 \times 10^{-4}$)

ICE BOX PROCESS

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$

	HF	H_2O	H_3O^+	F^-
i	.150		0	0
C	−x		+x	+x
E	.150 − x		x	x

assume $x \ll .150$ ∴ $(.150 - x) \approx (.150)$

$K_a = \frac{[H^+][F^-]}{[HF]} = \frac{x^2}{.150} = 7.2 \times 10^{-4}$

$x = .0104 = [H^+]$

$pH = -\log[H^+] = -\log(.0104) = 1.98$

What is the % dissociation of HF in a 0.150 M solution?

$\frac{[H^+]}{[HF]} \times 100 = \frac{.0104}{.150} \times 100 = 6.9\%$

(I'm using 2 sig figs because $K_a = 7.2 \times 10^{-4}$)

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STATION 3—EQUILIBRIA

EQUIL. FAVORS WEAK ACID & WEAK BASE

The following equation is written to represent relative concentrations in solution:

$$HC_2H_3O_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$$

WEAKER ACID — WEAKER BASE

Which species is the stronger acid: $HC_2H_3O_2$ or H_3O^+

H_3O^+ is the BETTER DONOR we see it in the equation AFTER IT HAS DONATED... H_2O

Knowing that H_2S is a stronger acid than HCN, re-write the following equilibrium to show which species are more concentrated in solution. HCN is the WEAKER ACID

$$HCN(aq) + HS^-(aq) \rightleftharpoons CN^-(aq) + H_2S(aq)$$

$HCN_{(aq)} + HS^-_{(aq)} \rightleftharpoons CN^-_{(aq)} + H_2S_{(aq)}$

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STATION 4—pH & pOH

[H+]	[OH−]	pH	pOH	acidic or basic?
5.6×10^{-4}	1.8×10^{-11}	3.25	10.75	ACIDIC
2.5×10^{-2}	4.0×10^{-13}	1.60	12.40	ACIDIC
1.4×10^{-10}	7.1×10^{-5}	9.85	4.15	BASIC

$[H^+] = 10^{-pH}$ $\quad$ $[OH^-] = 10^{-pOH}$

$pH = -\log[H^+]$ $\quad$ $pOH = -\log[OH^-]$

$[H^+][OH^-] = 1 \times 10^{-14}$ $\quad$ $pH + pOH = 14$

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STATION 5—ICE BOX

A 0.0300 M solution of the weak acid, HA, has a pH of 4.25.

a. Write the dissociation equation for this weak acid.

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

b. Calculate the $[H^+]$ for a 0.0300 M solution of this weak acid.

USE THE pH $pH = 4.25$ $[H^+] = 10^{-pH} = \boxed{5.6 \times 10^{-5}\ M}$

c. Calculate the K_a of this weak acid.

	HA	$+ H_2O \rightleftharpoons$	H_3O^+	$+ A^-$
I	.0300 M		0	0
C	-5.6×10^{-5}		$+5.6\times10^{-5}$	$+5.6\times10^{-5}$
E	$.0300 - 5.6\times10^{-5}$		5.6×10^{-5}	5.6×10^{-5}

NOTE: DON'T USE "X" .. YOU KNOW THE VALUES.

$K_a = \dfrac{[H_3O^+][A^-]}{[HA]} = \dfrac{(5.6\times10^{-5})^2}{(.0300 - 5.6\times10^{-5})} = 1.056\times10^{-7}$

$K_a = \boxed{1.1\times10^{-7}}$

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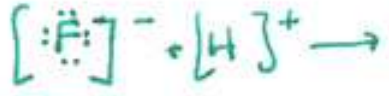
STATION 6—LEWIS & BRONSTED-LOWRY DEFINITIONS

Consider the reaction: $F^- + H_2O \rightleftharpoons HF + OH^-$



F- would be a (circle all answers that apply):

Brønsted-Lowry acid — (Brønsted-Lowry base) — Lewis acid — (Lewis base)

F⁻ accepts an H⁺

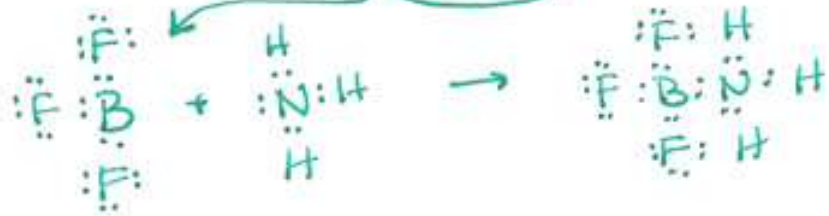
F⁻ DONATES e⁻ pair to make BOND

Consider the reaction: $BF_3 + NH_3 \rightarrow BF_3NH_3$

BF_3 would be a (circle all answers that apply):

e⁻ pair acceptor

Brønsted-Lowry acid — Brønsted-Lowry base — (Lewis acid) — Lewis base



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STATION 7—DIPROTIC ACIDS

Consider the diprotic acid, sulfurous acid: H_2SO_3 $K_{a1} = 1.2 \times 10^{-2}$ $K_{a2} = 6.6 \times 10^{-8}$

a. Write the equations for the step-wise dissociation of sulfurous acid.

$H_2SO_3 + H_2O \rightleftharpoons H_3O^+ + HSO_3^-$ K_{a1}

$HSO_3^- + H_2O \rightleftharpoons H_3O^+ + SO_3^{2-}$ K_{a2}

In a solution of H_2SO_3, all of the H_3O^+ is from the 1st step.

b. Calculate the pH of a 0.0200 M solution of H_2SO_3.

$H_2SO_3 + H_2O \rightleftharpoons H_3O^+ + HSO_3^-$

	H_2SO_3	H_2O	H_3O^+	HSO_3^-
I	.0200		0	0
C	$-x$		$+x$	$+x$
E	$.0200 - x$		x	x

(b) ICE BOX using K_{a1}

OOPS! You must use QUADRATIC

pH = 1.97

c. What is the $[SO_3^{2-}]$ in a 0.0125 M solution of H_2SO_3?

$K_{a2} = \dfrac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]} = 6.6 \times 10^{-8}$

BECAUSE OF STEP 1, $[H_3O^+] \approx [HSO_3^-]$

SO $\dfrac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]} = 6.6 \times 10^{-8}$ $[SO_3^{2-}] = K_{a2}$

MEMORIZE!

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STATION 8—SALT SOLUTIONS

For each solution, state whether it would be **Acidic**, **Basic**, or **Neutral**.

A	NH_4Cl (NH_3)	N	Na_2SO_4	B	$KC_2H_3O_2$ ($HC_2H_3O_2$)
B	CaF_2 (HF)	B	$MgSO_3$ (H_2SO_3)	N	KI

NOTE: THE ONLY conjugate acid we usually see is NH_4^+ (NH_3) or $CH_3NH_3^+$ (CH_3NH_2)

HCN is a weak acid with a $K_a = 6.2 \times 10^{-10}$.

a. Write the equation for the equilibrium that exists in a solution of KCN.

$CN^- + H_2O_{(l)} \rightleftharpoons HCN + OH^-$

CN^- ... K^+ is a spectator (conjugate base of HCN)

b. Calculate the K_b for CN^-.

$K_b = \dfrac{K_w}{K_a} = \dfrac{1 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}$

c. Calculate the pH of a 0.0200 M solution of KCN.

$CN^- + H_2O \rightleftharpoons HCN + OH^-$ ICE BOX

	CN^-	H_2O	HCN	OH^-
I	.0200		0	0
C	$-x$		$+x$	$+x$
E	$.0200 - x$		x	x

assume $x \ll .02$ etc.

$x = [OH^-] = \sqrt{(.0200)(1.6 \times 10^{-5})}$

$= 5.68 \times 10^{-4}$

pOH = 3.245

pH = 14 − pOH = 10.75

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STATION 9—WEAK BASES

NH_3 is a weak base with a $K_b = 1.8 \times 10^{-5}$.

a. Write the dissociation equation for NH_3 in water.

$NH_3 + H_2O(l) \rightleftharpoons NH_4^+ + OH^-$

b. Calculate the pH of a 0.100 M solution of NH_3. ICE BOX ... SHORT CUT

$x = [OH^-] = \sqrt{(.100)(1.8 \times 10^{-5})} = 1.34 \times 10^{-3}$

$pOH = -\log[OH^-] = 2.87$ $\quad pH = 14 - pOH = \boxed{11.13}$

c. Calculate the K_a for the conjugate acid of NH_3.

$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \boxed{5.56 \times 10^{-10}}$

d. Write the equation for the equilibrium that exists in a solution of NH_4Cl.

$NH_4^+ + H_2O(l) \rightleftharpoons NH_3 + H_3O^+$