Ch 17 - Acid-Base Equilibria
Annotated Answers:

1.	**(B) HCN is an acid and CN⁻ is its conjugate base.** Look for two species that differ only by H^+. The acid has the extra H^+.
2.	**(B) NH_4^+** A Brønsted-Lowry base requires a lone pair as a place for the proton, H^+ ion, to attach.
3.	**(D) HCN and HOCN.** These two differ by more than an H^+.
4.	**(E) Chloroacetic acid, $K_a = 1.4 \times 10^{-3}$** The strongest acid is the acid with the largest K_a value.
5.	**(A) H_2CO_3** All of the choices are acids. A weak acid is a weak electrolyte. Eliminate all other choices because they are strong acids (memorize the strong acid list).
6.	**(B) 3.38** HBr is a strong acid, so $[H^+] = [HBr] = 4.2 \times 10^{-4}$ M. pH = $-\log(4.2 \times 10^{-4}) = 3.38$
7.	**(E) OH⁻** If you want the conjugate base of H_2O, H_2O must be an acid (a donor). After it donates, OH^- is left.
8.	**(C) 3.87** pH = $-\log [H^+]$ = $-\log(1.34 \times 10^{-4}) = 3.87$
9.	**(E) very basic** pH >> 7 is very basic
10.	**(A) 2.0×10^{-11}** The equation shows the conjugate base of HNO_2 going through hydrolysis to form OH^-. $K_a \cdot K_b = K_w$ for a weak acid & its conjugate. So… $K_b = K_w/K_a = 1.0 \times 10^{-14}/5.0 \times 10^{-4} = 2.0 \times 10^{-11}$.
11.	**(D) an electron pair donor.** When a coordinate covalent bond is formed, the species that supplies the pair of electrons for the bond is the Lewis Base. The species with the empty orbital is the Lewis Acid.
12.	**(A) equilibrium lies to the left (reactant-favored)** Equilibrium favors the weaker acid/weaker base. HCN (the reactant) will be favored.
13.	**(C) 1.1×10^{-7}** HA is the formula for a generalized acid. HA = $H^+ + A^-$ ("A" stands for "anion".) $[H^+] = 10^{-pH} = 10^{-3.82} = 1.51 \times 10^{-4}$ M. Make an ICE box with this information and substitute the values into the K_a expression.

14.	**(C) 4.0**
	Treat this acid as though it were monoprotic with the $K_a = 1.0 \times 10^{-7}$. This is an ICE box problem.
	$$H_2S \rightleftharpoons H^+ + HS^-$$
	(ICE table: I: .10 M, 0, 0; C: $-x$, $+x$, $+x$; E: $.10-x$, x, x)
	Assume $x \ll .10$ $\therefore (.10 - x) \approx (.10)$
	$$K_a = \frac{[H^+][HS^-]}{[H_2S]} = 1.0 \times 10^{-7}$$
	$$\frac{x^2}{(.10)} = 1.0 \times 10^{-7}$$
	$$x = \sqrt{x^2} = \sqrt{1.0 \times 10^{-8}} = 1.0 \times 10^{-4}$$
	$$x = [H^+] \Rightarrow pH = -\log[H^+] = \boxed{4.00}$$

15.	**(A) $HIO_2 < HIO_3 < HClO_3 < HClO_4$**
	When comparing oxoacids, the more O's, the stronger the acid.

16.	**(A) 4.4×10^{-7}**
	This is similar to problem 13, an ICE Box problem. $[H^+] = 10^{-pH} = 10^{-3.18} = 6.607 \times 10^{-4}$ $\underline{M}$.
	$$HA \rightleftharpoons H^+ + A^-$$
	(ICE table: I: 1.00 M, 0, 0; C: -6.6×10^{-4}, $+6.6 \times 10^{-4}$, $+6.6 \times 10^{-4}$; E: $.99934$, 6.6×10^{-4}, 6.6×10^{-4})
	$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(6.6 \times 10^{-4})^2}{.99934}$$
	$$= 4.3588 \times 10^{-7}$$

17.	**(D) 12.73**
	NaOH is a strong base. $[OH^-] = [NaOH] = 0.054$. pOH = -log $[OH^-]$ = 1.27 pH = 14 - pOH

18.	**(D) 8.15**
	Recognize that NaF is the salt of the weak acid and strong base (HF + NaOH). The solution will be basic because of the conjugate base, F^-. The reaction is: $F^- + H_2O \rightleftharpoons HF + OH^-$. You will need to calculate the K_b of F^- knowing that $K_a \times K_b = K_w$. $K_b = 1.0 \times 10^{-14}/7.1 \times 10^{-4} = 1.4 \times 10^{-11}$
	This is an ICE box problem. $[F^-] = [NaF] = 0.144$ $\underline{M}$. Notice that $x = [OH^-]$.
	$$F^- + H_2O_{(l)} \rightleftharpoons HF + OH^-$$
	(ICE table: I: .144M, —, 0, 0; C: $-x$, —, $+x$, $+x$; E: $.144-x$, —, x, x)
	assume $x \ll .144$ $\therefore (.144 - x) \approx (.144)$
	$$K_b = \frac{[HF][OH^-]}{[F^-]} = 1.4 \times 10^{-11}$$
	$$\frac{x^2}{.144} = 1.4 \times 10^{-11}$$
	$$x = [OH^-] = \sqrt{x^2} = \sqrt{(1.4 \times 10^{-11})(.144)}$$
	$$= \sqrt{2.028 \times 10^{-12}} = 1.42 \times 10^{-6}$$
	$$pOH = -\log(1.42 \times 10^{-6}) = 5.846 \; ; \; pH = 14 - pOH = \boxed{8.15}$$

19.	**(A) NH_4Cl**
	"Lowest pH" means most acidic. We are looking for the conjugate acid of a weak base.
	(B) and (D) will make the solution basic; (C) & (E) will leave the solution neutral .

| 20. | **(B) 5.05**
HCN is a weak acid. This is an ICE box problem.

$HCN \rightleftharpoons H^+ + CN^-$

	HCN	H^+	CN^-
I	.200M	O	O
C	$-x$	$+x$	$+x$
E	.200$-x$	x	x

assume $x << .200$
$\therefore (.200 - x) \approx (.200)$

$K_a = \dfrac{[H^+][CN^-]}{[HCN]} = 4.0 \times 10^{-10}$

$= \dfrac{x^2}{.200} = 4.0 \times 10^{-10}$

$x = [H^+] = \sqrt{x^2} = \sqrt{(4.0 \times 10^{-10})(.200)}$

$= \sqrt{8.0 \times 10^{-11}} = 8.944 \times 10^{-6}$

$pH = -\log [H^+] = -\log (8.944 \times 10^{-6}) = 5.048$ |

| 21. | **(E) NaBr**
The pH will NOT change if the salt is the result of the neutralization of a strong acid + strong base.
HBr (strong) + NaOH (strong) → H_2O + NaBr.   Note that (A) and (B) come from diprotic acids, but (A) is the conjugate base of a weak acid while (B) is the conjugate base of a strong acid and will continue to donate H^+'s to make the solution acidic.
(B) and (D) will make the solution acidic. (A) and (C) will make the solution basic. |

| 22. | **(C) 1.3 x 10^{-13}**
The $[S^{2-}] = K_{a2} = 1.3 \times 10^{-13}$. This is a special case for diprotic acids. Essentially all of the H+'s come from the first dissociation.

$H_2S \rightleftharpoons H^+ + HS^-$ $\therefore [H^+] \approx [HS^-]$

$K_{a_2} = \dfrac{[H^+][S^{2-}]}{[HS^-]}$ $\therefore [S^{2-}] = K_{a2}$ |