AP CHEMISTRY

TOPIC 7: ACIDS & BASES, REVIEW, PART I

Day 79:

- Nature of Acids and Bases
- pH and pOH scale
- Bases
- Acid Strength
- pH of strong acids
- Weak bases
- Water as an acid and a base
- pH of weak acids

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1. What is the conjugate base of the bicarbonate ion, HCO_3^{-1}?

If the question is asking for the conjugate base of something, ASSUME that something is an ACID, therefore the answer is CO_3^{-2}

2. Write the dissociation reaction for each of the following acids in water and identify the conjugate acid-base pairs:

a. $HC_2H_2ClO_2$

$\boldsymbol{HC_2H_2ClO_2}$	+	$\boldsymbol{H_2O_{(l)}}$	$\rightleftharpoons$	$\boldsymbol{C_2H_2ClO_2^{-1}}$	+	$\boldsymbol{H_3O^{+1}}$
A		B		CB		CA

b. HCN

HCN	+	$\boldsymbol{H_2O_{(l)}}$	$\rightleftharpoons$	$\boldsymbol{CN^{-1}}$	+	$\boldsymbol{H_3O^{+1}}$
A		B		CB		CA

c. HNO_2

HNO_2	+	$\boldsymbol{H_2O_{(l)}}$	$\rightleftharpoons$	$\boldsymbol{NO_2^{-1}}$	+	$\boldsymbol{H_3O^{+1}}$
A		B		CB		CA

3. Write the equilibrium expression for each of the reactions (from question 2.)

a) $$K = \frac{\left[C_2H_2ClO_2^{-1}\right]\left[H_3O^{+1}\right]}{\left[HC_2H_2ClO_2\right]} \quad \textbf{\textit{OR}} \quad K = \frac{\left[C_2H_2ClO_2^{-1}\right]\left[H^{+1}\right]}{\left[HC_2H_2ClO_2\right]}$$

b) $$K = \frac{\left[CN^{-1}\right]\left[H_3O^{+1}\right]}{\left[HCN\right]} \quad \textbf{\textit{OR}} \quad K = \frac{\left[CN^{-1}\right]\left[H^{+1}\right]}{\left[HCN\right]}$$

c) $$K = \frac{\left[NO_2^{-1}\right]\left[H_3O^{+1}\right]}{\left[HNO_2\right]} \quad \textbf{\textit{OR}} \quad K = \frac{\left[NO_2^{-1}\right]\left[H^{+1}\right]}{\left[HNO_2\right]}$$

4. Calculate the equilibrium constant, *K*, at a certain temperature for the reaction:

$$2\ NH_{3\,(g)} \rightleftharpoons 3\ H_{2\,(g)} + N_{2\,(g)}$$

if the equilibrium concentration for [H_2] is 0.850 *M*, and the initial concentration for [NH_3] was 3.22 *M*.

Answer: We <u>CANNOT</u> use the 5 % rule for this type of question !!!

	2 [NH_3]	$\rightleftharpoons$	3 [H_2]	+	[N_2]
I	3.22 *M*		0		0
C	- 2*x* = - 0.566		+ 3*x* = 0.850		+ *x* = (0.850 / 3) = 0.283
E	3.22 - 0.566 = 2.654		0.850		0.283

$$K = \frac{\left[\ H_2\ \right]^3\left[\ N_2\ \right]}{\left[\ NH_3\ \right]^2} = \frac{(0.850)^3(0.283)}{(2.654)^2} = 2.47\times10^{-2}$$

5. If the equilibrium constant at 444 ^{0}C for

$$2\ HI_{(g)} \rightleftharpoons H_{2\,(g)} + I_{2\,(g)}$$

is 1.39 x 10^{-2} calculate the equilibrium constant for the reverse reaction at 444 ^{0}C.

Answer:

$$K = \frac{\left[\ H_2\ \right]\left[\ I_2\ \right]}{\left[\ HI\ \right]^2} = 1.39\times10^{-2} \quad \textbf{\textit{THERFORE,}} \quad K' = \frac{\left[\ HI\ \right]^2}{\left[\ H_2\ \right]\left[\ I_2\ \right]} = \frac{1}{1.39\times10^{-2}} = 71.9$$

6. Fill in the missing information in the following table:

pH	pOH	[H^{+1}]	[OH^{-1}]	acid, base or neutral
11.93	***2.07***	***1.17 x 10^{-12}***	***8.51 x 10^{-3}***	***BASE***
4.46	***9.54***	3.45 x 10^{-5}	***2.90 x 10^{-10}***	***ACID***
6.65	7.35	***2.24 x 10^{-7}***	***4.47 x 10^{-8}***	***ACID***
12.76	***1.24***	***1.72 x 10^{-13}***	5.8 x 10^{-2}	***BASE***

7. The pOH of a 400 mL solution of HNO_3 is 12.44. How many grams of HNO_3 are in solution?

HNO_3 is a STRONG acid

[OH^{-1}] = antilog (- 12.44) = 3.63 x 10^{-13} M,

$$\left[\ H^{+1}\ \right] = \frac{1.00\times10^{-14}}{3.63\times10^{-13}} = 0.0275\ M = \left[\ HNO_3\ \right]$$

$$\frac{400\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.0275\ mol}{L} \times \frac{63\ g}{1\ mol\ HNO_3} = 0.69\ g$$

8. Calculate the pH of a 0.237 *M* solution of benzoic acid, HC_6H_5COO (K_a = 6.14 x 10 $^{-5}$)

	$[HC_6H_5COO]$	$\rightleftharpoons$	$[C_6H_5COO^{-1}]$	+	$[H^{+1}]$
I	0.237 *M*		0		0
C	- x		+ x		+ x
E	0.237 - x		x		x

$$K_a = 6.14 \times 10^{-5} = \frac{\left[C_6H_5COO^{-1} \right]\left[H^{+1} \right]}{\left[C_6H_5COOH \right]} = \frac{x^2}{0.237 - x} = \frac{x^2}{0.237}$$

$$6.14 \times 10^{-5} (0.237) = x^2 \; ; \; x = \sqrt{1.46 \times 10^{-5}} = 3.81 \times 10^{-3} \; M$$

$[H^{+1}] = 3.81 \times 10^{-3}$

pH = - log (3.81 x 10^{-3}) = 2.42

9. Calculate the K_a of a 0.0062 *M* weak monoprotic acid with a pH of 3.44. (Do not use the 5 % rule, and do not use the quadratic equation !!!)

monoprotic acid: an acid with one acidic proton (hydrogen ion).

$$HA \rightleftharpoons H^{+1} + A^{-1}$$

pH = 3.44

$[H^{+1}]$ = antilog (-3.44) = 3.63×10^{-4} *M*

	[HA]	$\rightleftharpoons$	$[H^{+1}]$	+	$[A^{-1}]$
I	0.0062 *M*		0		0
C	- x = - 3.63 x 10^{-4}		+ x = + 3.63 x 10^{-4}		+ x = + 3.63 x 10^{-4}
E	0.0062 – 0.000363 = 5.837 x 10^{-3}		3.63 x 10^{-4}		3.63 x 10^{-4}

$$K_a = \frac{\left[H^{+1} \right]\left[A^{-1} \right]}{\left[HA \right]} = \frac{\left[3.63 \times 10^{-4} \right]\left[3.63 \times 10^{-4} \right]}{\left[5.837 \times 10^{-3} \right]} = 2.26 \times 10^{-5}$$