

# AP CHEMISTRY

## TOPIC 7: ACIDS & BASES, REVIEW, PART I

Day 79:

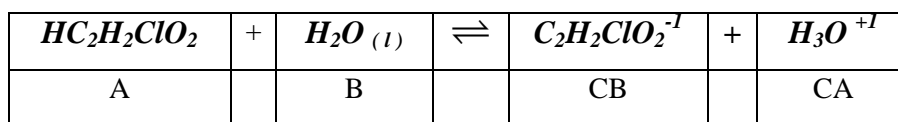
- Nature of Acids and Bases
- Acid Strength
- Water as an acid and a base
- pH and pOH scale
- pH of strong acids
- pH of weak acids
- Bases
- Weak bases

1. What is the conjugate base of the bicarbonate ion,  $\text{HCO}_3^{-1}$ ?

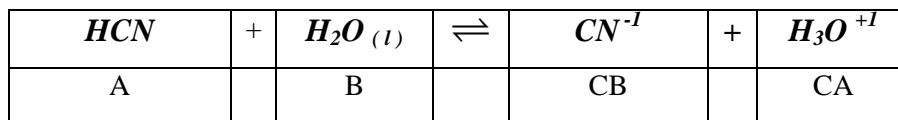
*If the question is asking for the conjugate base of something, ASSUME that something is an ACID, therefore the answer is  $\text{CO}_3^{-2}$*

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

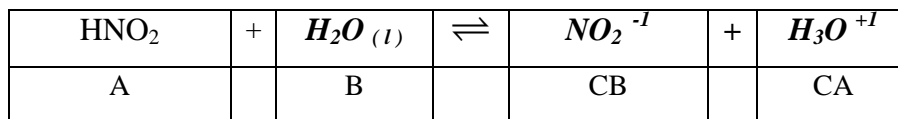
a.  $\text{HC}_2\text{H}_2\text{ClO}_2$



b. HCN



c.  $\text{HNO}_2$



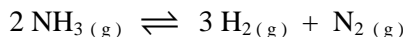
3. Write the equilibrium expression for each of the reactions ( from question 2.)

a) 
$$K = \frac{[\text{C}_2\text{H}_2\text{ClO}_2^{-1}][\text{H}_3\text{O}^{+1}]}{[\text{HC}_2\text{H}_2\text{ClO}_2]} \quad \text{OR} \quad K = \frac{[\text{C}_2\text{H}_2\text{ClO}_2^{-1}][\text{H}^{+1}]}{[\text{HC}_2\text{H}_2\text{ClO}_2]}$$

b) 
$$K = \frac{[\text{CN}^{-1}][\text{H}_3\text{O}^{+1}]}{[\text{HCN}]} \quad \text{OR} \quad K = \frac{[\text{CN}^{-1}][\text{H}^{+1}]}{[\text{HCN}]}$$

c) 
$$K = \frac{[\text{NO}_2^{-1}][\text{H}_3\text{O}^{+1}]}{[\text{HNO}_2]} \quad \text{OR} \quad K = \frac{[\text{NO}_2^{-1}][\text{H}^{+1}]}{[\text{HNO}_2]}$$

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



if the equilibrium concentration for  $[\text{H}_2]$  is  $0.850 \text{ M}$ , and the initial concentration for  $[\text{NH}_3]$  was  $3.22 \text{ M}$ .

**Answer: We CANNOT use the 5 % rule for this type of question !!!**

	$2 [\text{NH}_3]$	$\rightleftharpoons$	$3 [\text{H}_2]$	+	$[\text{N}_2]$
<b>I</b>	$3.22 \text{ M}$		$0$		$0$
<b>C</b>	$- 2x = - 0.566$		$+ 3x = 0.850$		$+ x = (0.850 / 3) = 0.283$
<b>E</b>	$3.22 - 0.566 = 2.654$		$0.850$		$0.283$

$$K = \frac{[\text{H}_2]^3 [\text{N}_2]}{[\text{NH}_3]^2} = \frac{(0.850)^3 (0.283)}{(2.654)^2} = 2.47 \times 10^{-2}$$

5. If the equilibrium constant at  $444^\circ\text{C}$  for



is  $1.39 \times 10^{-2}$  calculate the equilibrium constant for the reverse reaction at  $444^\circ\text{C}$ .

**Answer:**

$$K = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = 1.39 \times 10^{-2} \quad \text{THEREFORE,} \quad K' = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{1}{1.39 \times 10^{-2}} = 71.9$$

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

pH	pOH	$[\text{H}^+]$	$[\text{OH}^-]$	acid, base or neutral
11.93	<b>2.07</b>	<b><math>1.17 \times 10^{-12}</math></b>	<b><math>8.51 \times 10^{-3}</math></b>	<b>BASE</b>
<b>4.46</b>	<b>9.54</b>	$3.45 \times 10^{-5}$	<b><math>2.90 \times 10^{-10}</math></b>	<b>ACID</b>
<b>6.65</b>	7.35	<b><math>2.24 \times 10^{-7}</math></b>	<b><math>4.47 \times 10^{-8}</math></b>	<b>ACID</b>
<b>12.76</b>	<b>1.24</b>	<b><math>1.72 \times 10^{-13}</math></b>	$5.8 \times 10^{-2}$	<b>BASE</b>

7. The pOH of a 400 mL solution of  $\text{HNO}_3$  is 12.44. How many grams of  $\text{HNO}_3$  are in solution?

**$\text{HNO}_3$  is a STRONG acid**

$$[\text{OH}^-] = \text{antilog}(-12.44) = 3.63 \times 10^{-13} \text{ M,}$$

$$[\text{H}^+] = \frac{1.00 \times 10^{-14}}{3.63 \times 10^{-13}} = 0.0275 \text{ M} = [\text{HNO}_3]$$

$$\frac{400 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1 \text{ L}} \times \frac{0.0275 \text{ mol}}{1 \text{ L}} \times \frac{63 \text{ g}}{1 \text{ mol HNO}_3} = 0.69 \text{ g}$$

8. Calculate the pH of a 0.237 M solution of benzoic acid, HC<sub>6</sub>H<sub>5</sub>COO (  $K_a = 6.14 \times 10^{-5}$  )

	[HC <sub>6</sub> H <sub>5</sub> COO ]	⇌	[C <sub>6</sub> H <sub>5</sub> COO <sup>-1</sup> ]	+	[ H <sup>+1</sup> ]
<b>I</b>	0.237 M		0		0
<b>C</b>	- x		+ x		+ x
<b>E</b>	0.237 - x		x		x

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

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

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

$$pH = -\log ( 3.81 \times 10^{-3} ) = 2.42$$

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 ).**



$$pH = 3.44$$

$$[H^{+1}] = \text{antilog} ( -3.44 ) = 3.63 \times 10^{-4} M$$

	[ HA ]	⇌	[ H <sup>+1</sup> ]	+	[ A <sup>-1</sup> ]
<b>I</b>	0.0062 M		0		0
<b>C</b>	- x = - 3.63 x 10 <sup>-4</sup>		+ x = + 3.63 x 10 <sup>-4</sup>		+ x = + 3.63 x 10 <sup>-4</sup>
<b>E</b>	0.0062 - 0.000363 = 5.837 x 10 <sup>-3</sup>		3.63 x 10 <sup>-4</sup>		3.63 x 10 <sup>-4</sup>

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