

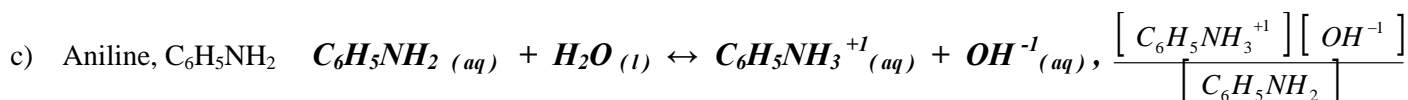
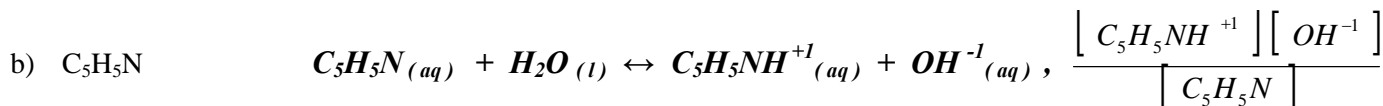
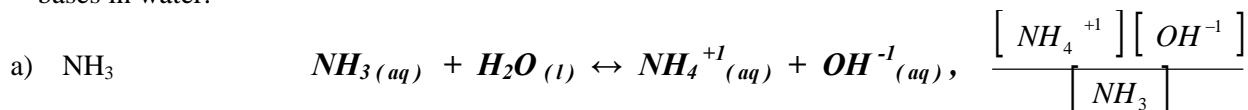
AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART C

Day 77:

- Bases
- Weak bases

1. Write the reaction and the corresponding K_b equilibrium expression for each of the following substances acting as bases in water.



2. Calculate the pH of the following substance.

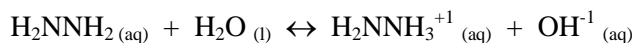


3. Calculate the concentration of an aqueous NaOH solution that has a pH = 10.50.

$$\text{pH} = 10.50, p\text{OH} = 14 - 10.50 = 3.50, [\text{OH}^-] = \text{antilog}(-3.50) = 3.16 \times 10^{-4} \text{ M} = [\text{NaOH}]$$

NaOH is a strong base = 100% dissociation.

4. For the reaction of hydrazine (N_2H_4) in water.



$K_b = 3.0 \times 10^{-6}$. Calculate the concentration of all the species and the pH of a 2.0 M solution of hydrazine in water.

Hydrazine is a WEAK BASE ...

	$[\text{H}_2\text{NNH}_2]$	+	$[\text{H}_2\text{O}]$	\rightleftharpoons	$[\text{H}_2\text{NNH}_3^+]$	+	$[\text{OH}^-]$
I	2.0 M		-		0		0
C	-x		-		+x		+x
E	2.0 - x		-		x		x

$$K_b = 3.0 \times 10^{-6} = \frac{[\text{H}_2\text{NNH}_3^+][\text{OH}^-]}{[\text{H}_2\text{NNH}_2]} = \frac{x^2}{2.0 - x} = \frac{x^2}{2.0}$$

$$3.0 \times 10^{-6} (2.0) = x^2 ; x = \sqrt{6.0 \times 10^{-6}} = 2.45 \times 10^{-3} \text{ M}$$

$$[\text{OH}^-] = 2.45 \times 10^{-3}; p\text{OH} = -\log(2.45 \times 10^{-3}) = 2.61; \text{pH} = 14 - 2.61 = 11.39$$

5. Calculate the pOH for a 2.75 M solution of an aniline, $C_6H_5NH_2$, solution. $K_b = 3.8 \times 10^{-10}$.

Answers: WEAK BASE ...

	[$C_6H_5NH_2$]	+	[H_2O]	\rightleftharpoons	[$C_6H_5NH_3^{+1}$]	+	[OH^{-1}]
I	2.75 M		-		0		0
C	- x		-		+ x		+ x
E	2.75 - x		-		x		x

$$K_b = 3.8 \times 10^{-10} = \frac{[C_6H_5NH_3^{+1}][OH^{-1}]}{[C_6H_5NH_2]} = \frac{(x)(x)}{2.75-x} = \frac{x^2}{2.75}$$

$$K_b = 3.8 \times 10^{-10} = \frac{x^2}{2.75}; \quad (3.8 \times 10^{-10})(2.75) = x^2$$

$$x = \sqrt{1.045 \times 10^{-9}} = 3.23 \times 10^{-5} M = [OH^{-1}]$$

$$pOH = -\log(3.23 \times 10^{-5}) = 4.49$$

6. Calculate the pH of a 0.0045 M solution of $C_2H_5NH_2$ solution. $K_b = 5.6 \times 10^{-4}$. (you may **NOT** use the 5 % rule for this problem.)

Answers: WEAK BASE ...

	[$C_2H_5NH_2$]	+	[H_2O]	\rightleftharpoons	[$C_2H_5NH_3^{+1}$]	+	[OH^{-1}]
I	0.0045 M		-		0		0
C	- x		-		+ x		+ x
E	0.0045 - x		-		x		x

$$K_b = 5.6 \times 10^{-4} = \frac{[C_2H_5NH_3^{+1}][OH^{-1}]}{[C_2H_5NH_2]} = \frac{x^2}{0.0045-x}$$

We cannot “drop” the x in this problem, the percent of error is greater than 5 %

You need to know how to do equilibrium problems that use the quadratic equation – I will place a question like this on the quiz !!! Everyone SHOULD know how to use the quadratic equation (WITHOUT a graphing calculator !!!)

$$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$K_b = 5.6 \times 10^{-4} = \frac{x^2}{0.0045 - x}$$

$$(5.6 \times 10^{-4})(0.0045 - x) = x^2$$

$$2.52 \times 10^{-6} - 5.6 \times 10^{-4}x = x^2$$

$$0 = x^2 + 5.6 \times 10^{-4}x - 2.52 \times 10^{-6}$$

$$\frac{-(5.6 \times 10^{-4}) + \sqrt{(5.6 \times 10^{-4})^2 - 4(1)(-2.52 \times 10^{-6})}}{2(1)}$$

$$x = 1.33 \times 10^{-3}$$

$$\text{pOH} = -\log(0.00133) = 2.88$$

$$\text{pH} = 14 - 2.88 = 11.12$$

*If you did not use the quadratic equation, $x = 1.58 \times 10^{-3}$ (15.8 % error),
and your WRONG pH would have been equal to 11.20*