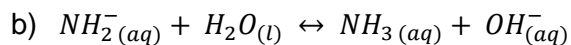
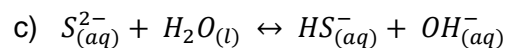
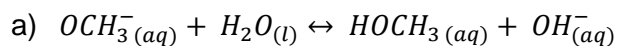


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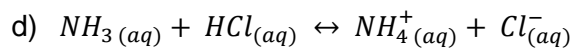
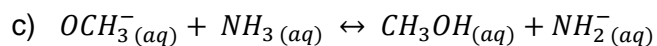
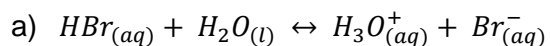
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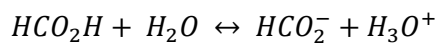
1) Write an expression for the ionization constant, K_b for each reaction.



2) Predict whether each equilibrium lies primarily to the left or to the right. Explain for each.



3) Calculate the pH of a 0.0010 M solution of formic acid, HCO_2H . $K_a = 1.8 \times 10^{-4}$



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- 4) Calculate (a) the pH and (b) the percent ionization of a 0.250 M $\text{HC}_2\text{H}_3\text{O}_2$ solution. $K_{a(\text{HC}_2\text{H}_3\text{O}_2)} = 1.8 \times 10^{-5}$. (The formula for acetic acid may also be written as CH_3COOH .)

HINT: Begin by filling out the equilibrium table below.

Balanced Equation	$\text{HC}_2\text{H}_3\text{O}_2$	\leftrightarrow	H^+	+	$\text{C}_2\text{H}_3\text{O}_2^-$
Initial Concentration (M)					
Change (M)					
Equilibrium Concentration (M)					

- 5) Calculate the pH of a 0.600 M solution of methylamine CH_3NH_2 .

$K_b = 4.4 \times 10^{-4}$ *HINT: Methylamine is a weak base. Then fill out the equilibrium table below.*

Rxn	CH_3NH_2	+	H_2O	\leftrightarrow	CH_3NH_3^+	+	OH^-
I							
C							
E							

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6) The pH of a 0.10 M solution of a weak base is 9.67. What is the K_b of the base? *Hint* - use a generic base equation when you don't know what the base is! $B^- + H_2O \leftrightarrow HB + OH^-$

7) Calculate the concentration of OH^- in a 0.25 M solution of NH_3 . Look up the K_b of NH_3 in your reference sheet.

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8) Identify the conjugate acid–base pairs in each equilibrium. Color code the pairs: acid and conjugate bases should be circled/highlighted/labeled in one color, base and conjugate acids in another.

