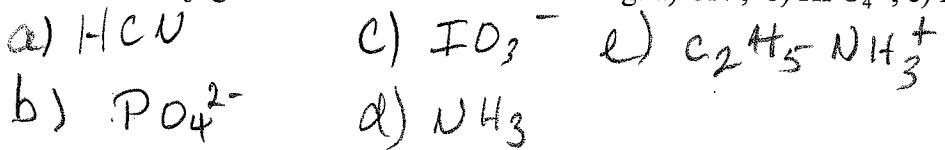


CH 223 – Worksheet 1

1. Give the conjugate acid or base of the following: a) CN^- , b) HPO_4^{2-} , c) HIO_3 , d) NH_4^+ , e) $\text{C}_2\text{H}_5\text{NH}_2$



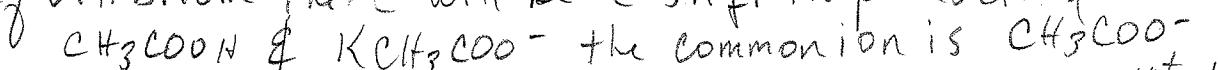
2. Predict whether aqueous solutions of the following substances are acidic, basic, or neutral: a) CrBr_3 ,
b) $\underline{\text{LiI}}$, c) K_3PO_4 , d) $\underline{\text{KCN}}$ e) $\underline{\text{NH}_4^+}$

basic basic basic acidic

acidic
single charged metal

3. Briefly define the following terms:

Common-Ion Effect: When a common ion is added to a system at equilibrium there will be a shift in products & reactants



Brunsted-Lowry acid and base: acid donates H^+ $\Rightarrow \text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$
base accepts H^+ $\Rightarrow \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{O}^{2-}$

A Strong acid: dissociates completely, example HNO_3

4. Calculate the pH of the following acid solutions: a) 0.00835 M HNO_3 , b) 0.525 g of HClO_4 in

a) HNO_3 strong acid

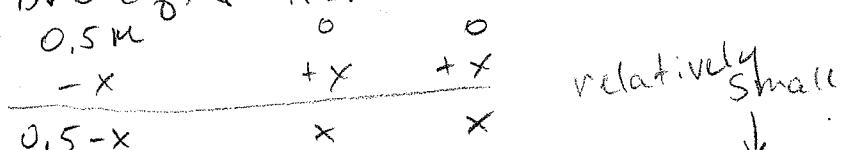
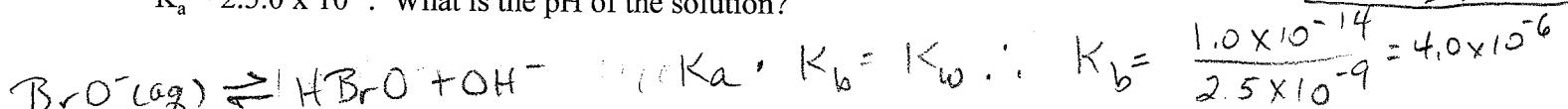
$$\text{pH} = -\log[\text{H}^+] = -\log(0.00835) = 2.07$$

c) $\text{Ca}(\text{OH})_2$ strong base

$$\begin{aligned} \text{b)} \quad & \text{HClO}_4 \text{ strong Acid } M/M = 100.459 \text{ mol} \\ & (0.5259)(\frac{1 \text{ mol}}{100.459}) = 0.00523 \text{ mol} \\ & [\text{HClO}_4] = \frac{0.00523 \text{ mol}}{0.575 \text{ L}} = 9.1 \times 10^{-3} \text{ mol/L} \\ & \therefore \text{pH} = -\log(9.1 \times 10^{-3}) = 2.04 \end{aligned}$$

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. Assume it all dissolves then $\text{pOH} = -\log[\text{OH}^-] = -\log(2 \cdot 0.0842) = 0.774$

5. Calculate the molar concentration of OH^- ions in 0.5 M solution of hypobromite ion BrO^- , $\text{pH} = 14 - 0.774 = 13.2$
 $K_a = 2.50 \times 10^{-9}$. What is the pH of the solution?



$$\% \text{ ionization} = \frac{2.0 \times 10^{-4}}{0.5} \times 100\% = 5.2 \times 10^{-4} \%$$

$$\begin{aligned} \text{K}_b &= \frac{[\text{HBrO}][\text{OH}^-]}{[\text{BrO}^-]} = \frac{x^2}{0.5-x} = 4.0 \times 10^{-6} \\ &\text{relatively large} \end{aligned}$$

$$\therefore \text{pOH} = -\log[\text{OH}^-] = -\log(2.0 \times 10^{-4}) = 5.70$$

$$\therefore \text{pH} = 14 - 5.70 = 8.30$$

$$x = \sqrt{(0.5-x)(4.0 \times 10^{-6})} = 2.0 \times 10^{-4}$$