

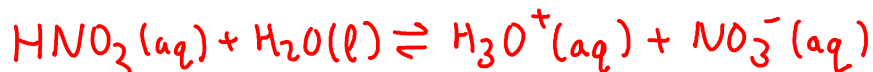
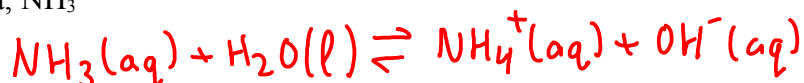
CHM 111 - Acid/Base Quiz 1

Name: SOLUTIONS

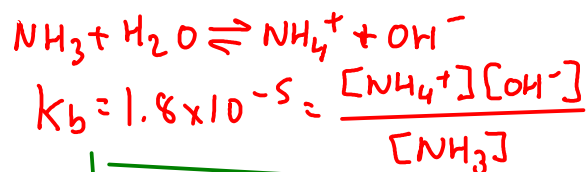
Due: 3/24/08 at 11:15 AM

Answer the following questions. Report all pH values to the nearest hundredth of a pH unit (Example: pH = 6.54). [20]

1) Write the ionization reactions for the following species when dissolved in water.

Nitric acid, HNO₃Ammonia, NH₃

Hypochlorous acid, HClO

Methylamine, CH₃NH₂Ammonium ion, NH₄⁺2) What is the pH of an 0.060 M solution of ammonia, NH₃? The K_b of ammonia is 1.8 × 10⁻⁵.• pH = 11.01

	initial	Δ	Equil
NH ₄ ⁺	0	+x	x
OH ⁻	0	+x	x
NH ₃	0.060	-x	0.060 - x

$$\frac{x^2}{0.060 - x} = 1.8 \times 10^{-5}$$

solved via quadratic eqn...

$$x = 1.0303 \times 10^{-3} = [\text{OH}^-]$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2.99$$

$$\text{pH} = 11.01$$

3) If the hydroxide ion concentration of a solution is $3.75 \times 10^{-6} \text{ M}$, what is the pH of the solution? Is the solution acidic or basic?

- pH = 8.57, so the solution is basic (pH > 7).

$$pOH = -\log_{10}(3.75 \times 10^{-6}) = 5.43$$

$$pH = 14 - pOH = 14 - 5.43$$

$$= \boxed{8.57}$$

4) Calculate the pH of a solution made by dissolving 0.0702 grams of the strong base potassium hydroxide, KOH, in enough water to make 250. mL of solution?

- pH = 11.70 KOH: FW = 56.1056 *strong base!*

$$0.0702 \text{ g KOH} \times \frac{\text{mol KOH}}{56.1056 \text{ g KOH}} = 1.251212 \times 10^{-3} \text{ mol KOH}$$

$$\frac{1.251212 \times 10^{-3} \text{ mol KOH}}{0.250 \text{ L}} = 5.004848 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

$$pOH = -\log_{10}(5.004848 \times 10^{-3}) = 2.30$$

$$pH = 14 - pOH = \boxed{11.70}$$

5) What is the pH of an 0.15 M solution of formic acid, HCHO₂? The K_a of formic acid is 1.7×10^{-4} .

- pH = 2.30



$$K_a = 1.7 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]}$$

	Initial	Δ	Equil
H ₃ O ⁺	0	+x	x
CHO ₂ ⁻	0	+x	x
HCHO ₂	0.15	-x	0.15-x

$$\frac{x^2}{0.15-x} = 1.7 \times 10^{-4}$$

solve via quadratic eqn...

$$x = 4.966 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

$$pH = \boxed{2.30}$$