

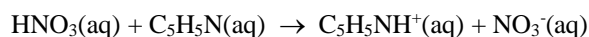
WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NOTE: EXAM 2 is **Wednesday, October 10th**. It will cover Chapter 15, and Sections 16.1, 16.3, 16.4, and 16.5 of Chapter 16. This includes the Chapter 16 ppt slides 1-31, and the Chapter 16 problems 1-10.

NAME _____ Panther ID _____

For problems involving calculations you must show your work for credit.

1) For the acid-base reaction below, identify the acid, the base, the conjugate base of the acid, and the conjugate acid of the base.



Acid HNO₃ Conjugate base NO₃⁻

Base C₅H₅N Conjugate acid C₅H₅NH⁺

2) For each of the following questions circle the correct answer. There may be more than one (or no) correct answer per problem.

A strong acid

HF HNO₂ HNO₃ HClO₃

A strong soluble base

AgOH KOH Cu(OH)₂ Fe(OH)₃

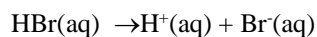
3) An aqueous solution has pH = 3.91 at T = 25. °C. Find the pOH, [H₃O⁺] (concentration of hydronium ion), and [OH⁻] (concentration of hydroxide ion) for the solution.

pOH 10.09 [H₃O⁺] 1.2 x 10⁻⁴ M [OH⁻] 8.1 x 10⁻¹¹ M

pOH = 14.00 - pH [H₃O⁺] = 10^{-pH} = 10^{-3.91} = 1.2 x 10⁻⁴ M
 = 14.00 - 3.91 = 10.09 [OH⁻] = 10^{-pOH} = 10^{-10.09} = 8.1 x 10⁻¹¹ M

4) Find the pH of each of the following solutions (at T = 25. °C):

a) A 0.0300 M aqueous solution of hydrobromic acid (HBr), a strong acid.

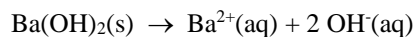


$$[\text{H}^{\text{+}}] = \frac{0.0300 \text{ mol HBr}}{\text{L soln}} \frac{1 \text{ mol H}^{\text{+}}}{1 \text{ mol HBr}} = 0.0300 \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}^{\text{+}}] = -\log_{10}(0.0300) = 1.52$$

Note I have written the reaction for HBr in the Arrhenius acid/base picture, but could have also done this using the Bronsted acid/base definitions.

b) 400.0 mL of an aqueous solution containing 13.8 g of barium hydroxide (Ba(OH)₂), MW = 171.3 g/mol), a strong soluble base.



$$\text{moles Ba(OH)}_2 = 13.8 \text{ g} \frac{1 \text{ mol}}{171.3 \text{ g}} = 0.0806 \text{ mol}$$

$$[\text{Ba(OH)}_2] = \frac{0.0806 \text{ mol}}{0.4000 \text{ L}} = 0.2014 \text{ M}$$

$$[\text{OH}^{-}] = \frac{0.2014 \text{ mol Ba(OH)}_2}{\text{L soln}} \frac{2 \text{ mol OH}^{-}}{1 \text{ mol Ba(OH)}_2} = 0.403 \text{ M}$$

$$\text{pOH} = -\log_{10}[\text{OH}^{-}] = -\log_{10}(0.403) = 0.39$$

$$\text{At } T = 25.0 \text{ }^{\circ}\text{C}, \text{pH} + \text{pOH} = 14.00, \text{ so } \text{pH} = 14.00 - \text{pOH} = 14.00 - 0.39 = 13.61$$