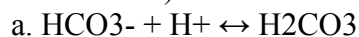
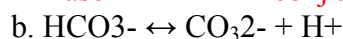


Worksheet 5. Aqueous Equilibrium Problems; Simple Equilibria

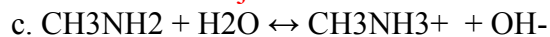
1. Identify the acid/base and their conjugate base/acid, and which definition you use to determine (Bronsted, Arrhenius or Lewis):



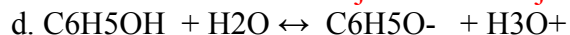
Base conj acid: Bronsted



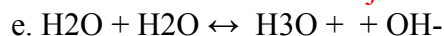
Acid conj base : Arrhenius



Base acid conj acid conj base : Lewis



Acid base conj base conj acid : Lewis, Arrhenius, Bronsted



Acid base conj acid conj base

2. Assuming $K_w = 1 \times 10^{-14}$, calculate the molarity of OH^- in solutions at 25°C when the H^+ concentration is:

a. 0.2M

At 25°C , $K_w = [\text{OH}^-] [\text{H}^+] = 1 \times 10^{-14}$

$[\text{OH}^-] = 1 \times 10^{-14} / .2 = 5 \times 10^{-14}$

b. 5×10^{-10} M

$[\text{OH}^-] = 1 \times 10^{-14} / 5 \times 10^{-10} = 2 \times 10^{-5}$

c. 100 M

$[\text{OH}^-] = 1 \times 10^{-14} / 100 = 1 \times 10^{-16}$

3. For each of these **strong** acid/base solutions, calculate the molarity of OH^- , H^+ , pH and pOH

a. 0.01M NaOH

$[\text{OH}^-] = 0.01$

pOH = 2

$[\text{H}^+] = 1 \times 10^{-14} / .01 = 1 \times 10^{-12}$

pH = 12

b. 5M HNO_3

$[\text{H}^+] = 5 \text{ M}$

$[\text{OH}^-] = 1 \times 10^{-14} / 5 = 2 \times 10^{-15}$

pOH = 14.7

pH = -.7

c. 0.007M $\text{Ba}(\text{OH})_2$

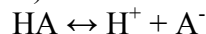
$[\text{OH}^-] = .014 \text{ M}$

$[\text{H}^+] = 1 \times 10^{-14} / .014 = 7.14 \times 10^{-13}$

pOH = 1.85

pH = 12.15

4. Fill in the blank (all concentrations are at equilibrium):



Acid	[HA]	[H ⁺]	[A ⁻]	K _a
Chlorous acid HClO ₂	0.6	.077	.077	1.0 x 10 ⁻²
Nitrous acid	18.6	.20	4	4.3 x 10 ⁻⁴
Hydrocyanic acid HCN	28	7 x 10 ⁻⁵	2 x 10 ⁻⁴	4.9 x 10 ⁻¹⁰
Phosphoric acid H ₃ PO ₄	0.076	0.3	.00192	7.6 x 10 ⁻³

5. The pH of a 0.115M solution of chloroacetic acid, ClCH₂COOH, is measured to be 1.92. Calculate K_a for this monoprotic acid.

$$pH = -\log[H_3O^+]$$

$$[H_3O^+] = 10^{-pH} = 10^{-1.92} = 0.012 \text{ M}$$

	HA	+	H ₂ O	→	H ₃ O ⁺	+	A ⁻
Initial	0.115 M				0 M		0 M
Change	-0.012 M				+0.012 M		+0.012 M
Equilibrium	0.103 M				0.012 M		0.012 M

$$K_a = [H_3O^+][A^-]/[HA] = (0.012)(0.012)/0.103 = 1.4 \times 10^{-3}$$

6. Calculate the concentrations of all the species and the pH in 0.10 M hypochlorous acid, HOCl. For HOCl, K_a = 3.5 x 10⁻⁸.

	HOCl	+	H ₂ O	↔	H ₃ O ⁺	+	OCl ⁻
Initial	0.10 M				0 M		0 M
Change	-x M				+xM		+xM
Equil	(0.10 - x)M				xM		xM

$$K_a = [H_3O^+][OCl^-]/[HOCl] = (x)(x)/(0.10-x) = 3.5 \times 10^{-8}$$

$$x^2/0.10 \approx 3.5 \times 10^{-8} \quad x^2 \approx 3.5 \times 10^{-9} \quad x \approx 5.9 \times 10^{-5}$$

$$[H_3O^+] = xM = 5.9 \times 10^{-5} \text{ M}$$

$$[OCl^-] = xM = 5.9 \times 10^{-5} \text{ M}$$

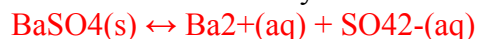
$$[HOCl] = (0.10 - x) \text{ M} = (0.10 - 0.000059) \text{ M} = 0.10 \text{ M}$$

$$[OH^-] = 1 \times 10^{-14}/[H_3O^+] = 1 \times 10^{-14}/5.9 \times 10^{-5} = 1.7 \times 10^{-10} \text{ M}$$

$$pH = -\log(5.9 \times 10^{-5}) = 4.23$$

7. One liter of saturated barium sulfate solution contains 0.0025 grams of dissolved BaSO₄.

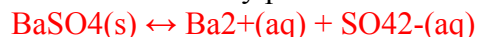
(a) What is the molar solubility of BaSO₄?



$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$

$$0.0025 \text{ g BaSO}_4 / 1.0 \text{ L} \times 1 \text{ mol BaSO}_4 / 233 \text{ g BaSO}_4 = 1.1 \times 10^{-5} \text{ mol BaSO}_4 / \text{L (dissolved)}$$

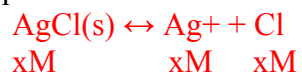
(b) Calculate the solubility product constant for BaSO₄.



$1.1 \times 10^{-5} \text{ mol/L}$ $1.1 \times 10^{-5} \text{ M}$ $1.1 \times 10^{-5} \text{ M}$

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (1.1 \times 10^{-5})(1.1 \times 10^{-5}) = 1.2 \times 10^{-10}$$

8. Calculate the molar solubility, silver and chloride ion concentrations, and solubility in grams per liter for saturated AgCl ($K_{sp} = 1.8 \times 10^{-10}$).



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (x)(x) = x^2 = 1.8 \times 10^{-10}$$

$$x = \text{molar solubility} = 1.3 \times 10^{-5} \text{ mol/L}$$

$$[\text{Ag}^+] = [\text{Cl}^-] = 1.3 \times 10^{-5} \text{ mol/L}$$

$$1.3 \times 10^{-5} \text{ mol AgCl/L} \times 143 \text{ g AgCl/mol AgCl} = 1.9 \times 10^{-3} \text{ g AgCl/L}$$