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):

a. $HCO_3^- + H^+ \leftrightarrow H_2CO_3$
Base conj acid: Bronsted

b. $HCO_3^- \leftrightarrow CO_3^{2-} + H^+$
Acid conj base : Arrhenius

c. $CH_3NH_2 + H_2O \leftrightarrow CH_3NH_3^+ + OH^-$
Base acid conj acid conj base : Lewis

d. $C_6H_5OH + H_2O \leftrightarrow C_6H_5O^- + H_3O^+$
Acid base conj base conj acid : Lewis, Arrhenius, Bronsted

e. $H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$
Acid base conj acid conj base

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2. Assuming $K_w = 1x10^{-14}$, calculate the molarity of OH^- in solutions at 25ºC when the H^+ concentration is:

a. 0.2M
At 25ºC, $Kw = [OH^-][H^+] = 1x10^{-14}$
$[OH^-] = 1x10^{-14}/ .2 = 5x10^{-14}$

b. 5×10^{-10} M
$[OH^-] = 1x10^{-14}/ 5 \times 10^{-10} = 2 \times 10^{-5}$

c. 100 M
$[OH^-] = 1x10^{-14}/ 100 = 1x10^{-16}$

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

a. 0.01M NaOH

$[OH^-] = 0.01$
pOH = 2
$[H^+] = 1x10^{-14}/.01 = 1x10^{-12}$
pH = 12

b. 5M HNO_3

$[H^+] = 5$ M
$[OH^-] = e\text{-}14 / 5 = 2x10^{-15}$
pOH = 14.7
pH = -.7

c. 0.007M $Ba(OH)_2$
$[OH^-] = .014$ M
$[H^+] = 1x10^{-14}/ .014 = 7.14x10^{-13}$
pOH = 1.85
pH = 12.15

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

$HA \leftrightarrow H^+ + A^-$

Acid	[HA]	$[H^+]$	$[A^-]$	K_a
Chlorous acid $HClO_2$	0.6	.077	.077	1.0×10^{-2}
Nitrous acid	18.6	.20	4	4.3×10^{-4}
Hydrocyanic acid HCN	28	7×10^{-5}	2×10^{-4}	4.9×10^{-10}
Phosphoric acid H_3PO_4	0.076	0.3	.00192	7.6×10^{-3}

5. The pH of a 0.115M solution of chloroacetic acid, $ClCH_2COOH$, is measured to be 1.92. Calculate Ka for this monoprotic acid.

pH = -log[H3O+]
[H3O+] = 10-pH = 10-1.92 = 0.012 M

	HA +	H2O → H3O+ +	A
Initial	0.115 M	0 M	0 M
Change	-0.012 M	+0.012 M	+0.012M
Equilibrium	0.103 M	0.012 M	0.012 M

Ka = [H3O+][A-]/[HA] = (0.012)(0.012)/0.103 = 1.4×10^{-3}

6. Calculate the concentrations of all the species and the pH in 0.10 M hypochlorous acid, HOCl. For HOCl, Ka = 3.5×10^{-8}.

	HOCl + H2O ↔	H3O+ +	OCl-
Initial	0.10 M	0 M	0 M
Change	-x M	+xM	+xM
Equil	(0.10 – x)M	xM	xM

Ka = [H3O+][OCl-]/[HOCl] = (x)(x)/(0.10-x) = 3.5 x 10-8
x2/0.10 ≈ 3.5 x 10-8 x2 ≈ 3.5 x 10-9 x2 ≈ 5.9 x 10-5
[H3O+] = xM = 5.9 x 10-5 M
[OCl-] = xM = 5.9 x 10-5 M
[HOCl] = (0.10 – x) M = (0.10 – 0.000059) M = 0.10 M
[OH-] = 1x1014/[H3O+] = 1x1014/5.9 x 10-5 1.7 x 10-10 M
pH = -log(5.9 x 10-5) = 4.23

7. One liter of saturated barium sulfate solution contains 0.0025 grams of dissolved $BaSO_4$.

(a) What is the molar solubility of $BaSO_4$?

BaSO4(s) ↔ Ba2+(aq) + SO42-(aq)
Ksp = [Ba2+][SO42-]

0.0025 g BaSO4/1.0 L x 1 mol BaSO4/233 g BaSO4 = 1.1 x 10-5 mol BaSO4/L (dissolved)

(b) Calculate the solubility product constant for BaSO4.

BaSO4(s) ↔ Ba2+(aq) + SO42-(aq)

1.1 x 10-5 mol/L 1.1 x 10-5 M 1.1 x 10-5 M
Ksp = [Ba2+][SO42-] = (1.1 x 10-5)(1.1 x 10-5) = 1.2 x 10-10

8. Calculate the molar solubility, silver and chloride ion concentrations, and solubility in grams per liter for saturated AgCl (Ksp = 1.8 x 10-10).

AgCl(s) ↔ Ag+ + Cl
xM xM xM

Ksp = [Ag+][Cl-] = (x)(x) = x2 = 1.8 x 10-10
x = molar solubility = 1.3 x 10-5 mol/L
[Ag+] = [Cl-] = 1.3 x 10-5 mol/L
1.3 x 10-5 mol AgCl/L x 143 g AgCl/1mol AgCl = 1.9 x 10-3 g AgCl/L