CHEM 1114 Exam 2, March 28, 2008

Name Key

By submitting this exam, you affirm that you have not received or given unauthorized aid.

Useful Information: , $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$, pH =-log[H_3O^+], $K_a \cdot K_b = K_w$, pH + pOH = 14.00, PV=nRT, R=0.0821 L•atm/mol•K, 8.314 J/mol•K K_W= 1.00 x 10^{-14}, (n/V)•RT = P,

$\ln\frac{k_2}{k_1} = \frac{-E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$, $\ln\frac{[A]_t}{[A]_0} = -kt$

(1)(4 points) Define the following

(a) A Lewis Base An electron-pair donor

(b) A Lewis Acid An electron-pair acceptor

(c) A Brønsted-Lowry Base an H^+ acceptor

(d) A Brønsted-Lowry Acid a H^+ donor

(2)(6 points) What is the pH of the following solutions?

(a) 0.500 M $HClO_4$

$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$ K = very large

	$HClO_4$		H_3O^+	ClO_4^-
init	0.500		0	0
change	−0.500		0.500	0.500
final	0		0.500	0.500

$pH = -\log(0.500) = 0.301$

(b) 0.100 M HCN

$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$ $K_a = 4.9 \times 10^{-10}$

	HCN		H_3O^+	CN^-	
init	0.100		0	0	
change	−x		x	x	$\frac{x^2}{0.100} = 4.9 \times 10^{-10}$
final	≈0.100		x	x	$x = 7.0 \times 10^{-6} = [H_3O^+]$

$pH = -\log(7.0 \times 10^{-6}) = 5.15$

(c) 0.300 M $NaNO_2$ NO_2^- $K_b = \frac{1.00\times10^{-14}}{4.5\times10^{-4}} = 2.2\times10^{-11}$

$NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$ $K_b = 2.2\times10^{-11}$

	NO_2^-		HNO_2	OH^-
init	0.300	→	0	0
change	$-x$		$+x$	$+x$
final	≈ 0.300		x	x

$\frac{x^2}{0.300} = 2.2\times10^{-11}$

$x = 2.6\times10^{-6}$

$pOH = -\log 2.6\times10^{-6} = +5.59$

$pH = 14 - 5.59 = \boxed{8.41}$

(3)(4 points) Will the following salts be acidic, basic, or neutral

(a) $LiNO_2$ basic

(b) NH_4CN $\frac{1.00\times10^{-14}}{4.5\times10^{-10}} = 2.0\times10^{-5} = K_b$ basic

5.5×10^{-10}

(c) NaOH basic

(d) $CaCO_3$ basic

(4)(6 points) Given the following information

$$H_2O(g) + C(s) \rightleftharpoons H_{2(g)} + CO_{(g)} \quad K = 55$$

(a) Write the formula for Kc.

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

(b) If 1.00 mole of H_2O is with excess C contained in a 1.00 L vessel and allowed to react until equilibrium is established, what is [CO] at equilibrium?

$H_2O + C_{(s)} \rightleftharpoons H_2 + CO$

	H_2O		H_2	CO
init	1.00 M	→	0	0
change	$-x$		$+x$	$+x$
final	$1.00-x$		x	x

$\frac{x^2}{1.00-x} = 55$

$x^2 = 55 - 55x$

$x^2 + 55x - 55 = 0$

$x = \frac{-55 \pm \sqrt{(55)^2 - 4(1)(-55)}}{2}$

$x = \frac{-55 \pm 57.0}{2}$

$x = 0.98$

$\boxed{[CO] = 0.98\ M}$

$\frac{(0.98)^2}{0.02} = 50 \approx 55$ right

(5)(2 points) Label the following as acid, base, conjugate acid, and conjugate base

acid base conjugate acid conjugate base

$NH_3 + CH_3^- \rightleftharpoons CH_4 + NH_2^-$

(6)(4 points)

(a) what is the K_a of ammonium?

NH_4 $K_a = \frac{1.00\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$

(b) What is the K_b of sulfate?

K_a of $HSO_4^- = 1.2\times10^{-2}$

$K_b = \frac{1.00\times10^{-14}}{1.2\times10^{-2}} = 8.3\times10^{-13}$

(7)(4 points) How will Q compare to K after the following procedures are performed on the equilibrium mixture (larger, smaller, equal to K)?

$NaHCO_{3(s)} \rightleftharpoons Na^+_{(aq)} + CO_{2(g)} + H_2O_{(l)}$ $Q = \frac{[Na^+][CO_2]}{1}$

(a) CO_2 is added to the system larger

(b) Na^+ is removed from the system smaller

(c) $NaHCO_3$ is added equal to K

(d) N_2 is added equal to K

(8)(4 points) For the following reaction, the initial and final concentrations are shown

	$2SO_3(g) \rightleftharpoons$	$2SO_{2(g)} + O_{2(g)}$	
initial	0.0500 M	0 M	0 M
	$-2x$	$+2x$	$+x$
final	0.0332 M	x M	y M

$-2x = 0.0332 - 0.0500 = -0.0168$

$x = 0.0084$

(a) What are the ~~values of~~ x and y?

$x = 0.0168\ M$ 0.0168 M

$y = 0.0084\ M$

(b) What is the numerical value of K?

$$K = \frac{(x)(2x)^2}{(0.0332)} = \frac{(0.0084)(0.0168)^2}{(0.0332)^2}$$

4.25×10^{-3}

0.215 $\times 10^{-3}$

(9)(4 points) For the reaction below, $K_c = 120$, If $[Cl_2] = 0.55$ M and $[ClF_3] = 5.22$ M at equilibrium, what is $[F_2]$?

$$Cl_{2(g)} + 3F_{2(g)} \rightleftharpoons 2ClF_{3(g)} \quad K_c = \frac{[ClF_3]^2}{[Cl_2][F_2]^3} = 120$$

$$120 = \frac{(5.22)^2}{(0.55)[F_2]^3}$$

$[F_2]^3 = 0.413$

$[F_2] = 0.74\ M$

(10)(4 points) Place the following acids in order of increasing acidity.

PH_3, HBr, H_2S, AlH_3

AlH_3, PH_3, H_2S HBr

(11)(4 points) A solution is made by mixing NH_4Cl and NH_3 in a 250 mL solution has a pH of 5.00. If the concentration of NH_4^+ is 0.500 M, what is $[NH_3]$? $K_a = 5.6 \times 10^{-10}$ for NH_4^+

$NH_4^+ (aq) + H_2O(aq) \rightleftharpoons NH_3 (aq) + H_3O^+ (aq)$

$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.6 \times 10^{-10}$ $[NH_4] = 0.500M$

$-\log[H_3O^+] = 5.00$

$\therefore [H_3O^+] = 1.0 \times 10^{-5}$

$K_a = \frac{[NH_3](1.0 \times 10^{-5})}{0.500} = 5.6 \times 10^{-10}$

$[NH_3] = 2.8 \times 10^{-5}$

(12)(4 points) When a saturated solution of sucrose is cooled, solid sugar forms. Is the following reaction endothermic or exothermic and why?

$\Delta H + C_{12}H_{22}O_{11}(s) \rightleftharpoons C_{12}H_{22}O_{11}(aq)$

½ for answer?

Endothermic. Heat is consumed in the forward rxn. When the temp is lowered, the K shifts to favor the product reagents including heat.

Extra Credit:(6 points) If the rate law of a reaction is found to be rate = k[A][B], which of the following mechanisms are valid and why?

Mechanism 1

not supported
rate determining step rate = $k[B]^2$

A ---> B (fast)
+2B ---> C (slow)

A + B ---> C

Mechanism 2

Mechanism 2 is supported
The rate determing step is
rate = $k[A][B]$, same as the experimental rate

A + B ----> D ~~(fast)~~ slow
+D ---> C (slow) fast

A + B ---> C