Surname :________________Given name :______________________________
Student number:__________________________

Lab Section:______________________________Lab TA: ____________________________

Chemistry 1311D
Test 2 V1
November 22, 2012
80 minutes You may remove last 2 pages

1. **(3 points)** Complete the following acid/base reactions (paying particular attention to the arrow used and states of products) and, where appropriate, identify acid/base conjugate pairs:

0.5 per rxn

a) $HNO_3(aq) + H_2O\ (l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$
A_1 B_1

b) $HClO_3(aq) + H_2O\ (l) \rightleftharpoons H_3O^+(aq) + ClO_3^-(aq)$
A_1 B_1

-0.5 for no labelling

-0.1 error in arrow

c) $ClO_4^-(aq) + H_2O\ (l) \not\longrightarrow$ No reaction

d) $NO_2^-(aq) + H_2O\ (l) \rightleftharpoons HNO_2(aq) + OH^-(aq)$

e) $HClO_4(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaClO_4(aq)$
or $\longrightarrow H_2O(l) + Na^+(aq) + ClO_4^-(aq)$

f) $HNO_2(aq) + NaOH\ (aq) \longrightarrow H_2O(l) + NaNO_2(aq)$
or $\longrightarrow H_2O(l) + Na^+(aq) + NO_2^-(aq)$

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2. A particular chemical reaction was studied and the half-life, found to be independent of concentration, was 25.7 s at 227 °C. At 277 °C the reactant concentration dropped from 1.5 M to 0.45 M in 5s.

a) **(1 point)** What is the order of the reaction?

1st order because $t_{1/2}$ independent of conc.

b) **(1 point)** What is the rate constant at 227 °C?

$t_{1/2} = \frac{\ln 2}{k}$; $k = \frac{\ln 2}{25.7 s} = 2.697 \times 10^{-2} s^{-1} \approx 2.70 \times 10^{-2} s^{-1}$

c) **(2 points)** What is the rate constant at 277 °C?

$\ln A = \ln A_0 - kt$

$\ln .45 = \ln 1.5 - k(5s)$

$k = 2.41 \times 10^{-1} s^{-1}$

d) **(3 points)** What is the activation energy associated with this process?

$\ln \frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

k_1 $2.70 \times 10^{-2} s^{-1}$ @ 500 K

k_2 $2.41 \times 10^{-1} s^{-1}$ @ 550 K

$\ln \frac{2.7 \times 10^{-2}}{24.1 \times 10^{-2}} = \frac{E_a}{8.314}\left(\frac{1}{550} - \frac{1}{500}\right)$

$18.201 = E_a(-1.81818 \times 10^{-4})$

$E_a = 100.1$ kJ

e) **(3 points)** At what temperature will the rate of the reaction be twice that of 277 °C?

$k_2 = 2.41 \times 10^{-1} s^{-1}$ @ 550

$k_1 = 2(2.41 \times 10^{-1}) s^{-1} = 4.82 \times 10^{-1} s^{-1}$ @ T?

$\ln(2) = \frac{100.1 \times 10^3}{8.314}\left(\frac{1}{550} - \frac{1}{T_1}\right)$

$5.757 \times 10^{-5} = \frac{1}{550} - \frac{1}{T_1}$

$1.7606 \times 10^{-3} = \frac{1}{T_1}$

$T_1 = 568 K$

295 °C

Version 1 Jenna ·

3. a) **(4 points)** A container contains CO_2 at p = 0.464 bar. Graphite is added to the container and some CO_2 is converted to CO; at equilibrium the total pressure is 0.746 bar. Determine the equilibrium constant for the reaction:

$P_i = 0.464$ bar

$P_T = 0.746$ bar

$P_T = P_{CO_2} + P_{CO}$

	$CO_2(g)$ +	C (graphite) $\rightleftharpoons$	$2CO(g)$
i	0.464	—	—
c	−x		+2x
e	.464−x		2x

$$P_T = 0.746 = (\overset{P_{CO_2}}{0.464 - x}) + \overset{P_{CO}}{2x} = 0.464 + x$$

$x = 0.282$ bar

$P_{CO_2} = 0.464 - 0.282 = 0.182$ bar

$P_{CO} = 0.564$ bar $= 2(0.282$ bar$)$

$$K_p = P_{CO}^2 / P_{CO_2} = \frac{(0.564)^2}{0.182} = 1.748$$

b) **(1 point)** The equilibrium system above is compressed to 1/3 its initial volume. Applying Le Chatelier's principle will reaction shift to left or right? Explain briefly.

Volume decreased, pressure increased, shifts to fewer moles of gas LEFT

c) **(4 points)** Find the new equilibrium pressure for the system.

Need starting non-equilibrium pressure following compression

$P_1V_1 = P_2V_2$ $P_1 = 0.464$ $V_f = V_i$, $P_2 = ?$ $V_2 = \frac{1}{3}V_1$

$(1.464)V_i = \frac{1}{3}P_2V_1$; $P_2 = 1.392$ bar.

	CO_2 +	C $\rightleftharpoons$	2CO
i	1.392	—	—
c	−x		2x
e	1.392−x		2x

$$K_p = \frac{(2x)^2}{1.392 - x} = 1.748$$

$$= \frac{4x^2}{1.392 - x}$$

→

d) **(1 point)** A catalyst is added to the reaction mixture. What is the role of a catalyst and how will this effect the equilibrium concentrations?

Catalyst speeds up reaction.

Catalyst does NOT effect equilibrium concentrations.

$4x^2 = 2.433 - 1.748x$

$4x^2 + 1.748x - 2.433 = 0$

$$x = \frac{-1.748 \pm \sqrt{1.748^2 - 4(4)(-2.433)}}{2(4)}$$

$$= \frac{-1.748 \pm 6.479}{8}$$

$x = 0.592.$

$P_{CO_2} = 1.392 - 0.591 = 0.801$ bar

$P_{CO} = 2(.592) = 1.184$

$P_T = 1.985$ bar

All vesions

4. Sodium hypochlorite, NaOCl, is the active ingredient in laundry bleach. Typical bleach solutions contain 0.64 M of NaOCl. Hypochlorous acid, HOCl, has a Ka of 4.0 x 10^-8^.

a) **(2 points)** What happens when solid NaOCl is added to water? Write all relevant chemical reations.

$$NaOCl(s) \xrightarrow{H_2O(l)} Na^+(aq) + OCl^-(aq)$$

$$OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$$

b) **(4 points)** Determine the equilibrium concentration of ALL species in an aqueous solution of 0.64 M NaOCl (there are five species, not including H_2O).

	OCl^-	$+ H_2O \rightleftharpoons$	$HOCl$	$+ OH^-$
i	0.64	—	—	10^{-7}
c	$-x$	—	x	$+x$
e	$0.64-x$	—	x	$10^{-7}+x$

$Ka = 4.0\times10^{-8}$

$Kb = \frac{10^{-14}}{4\times10^{-8}}$

$= 2.5\times10^{-7}$

$$Kb = \frac{x(10^{-7}+x)}{0.64-x} = \frac{x^2}{0.64-x} \approx \frac{x^2}{0.64} = 2.5\times10^{-7}$$

$$x = 4.0\times10^{-4} = [OH^-]$$

$[H_3O^+] = \frac{10^{-14}}{4\times10^{-4}} = 2.5\times10^{-11}\ M$

$[OH^-] = 4.0\times10^{-4} M$

$[Na^+] = 0.64M$

$[OCl^-] = 0.64 - 4.0\times10^{-4} \approx 0.64M$

$[HOCl] = 4\times10^{-4}\ M$

c) **(1 point)** What is the pH of this solution?

$$pH = -\log(2.5\times10^{-11}) = 10.6$$

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5. a) **(4 points)** What volumes of water and 1.06 M of ammonia [NH_3] should be used to prepare 2 L of buffer solution that contains 78.312 g of dissolved NH_4Cl and has pH 9.12. K_b of NH_3 is 1.8×10^{-5}.

NH_4Cl MM = = 53.489 moles = $\frac{78.312}{53.489 g/mol}$ = 1.464

$[NH_4^+]$ = 1.464/2L = 0.732 M

pH = 9.12 kb = 1.8×10^{-5} ka = Kw/Kb = 5.556×10^{-10}

pH = pKa + log $\frac{NH_3}{NH_4^+}$ = 9.12 = 9.26 + log $\frac{NH_3}{0.732}$

9.12 = 9.26 + log $\frac{NH_3}{0.732}$

−0.14 = log NH3/.732

$[NH_3]$ = 0.53 M

V_{NH_3} [1.06] = 2[.53]

V_{NH_3} = 1L

V_{H_2O} = 1L.

b) (1 points) Determine the resulting pH when 5 g of solid NaOH is added to 2 L of water. MM NaOH = 40.0g/mol ; 5g/40g/mole = 0.125 moles

0.125 moles in 2L $[OH^-]$ = 0.125/2L = 0.0625 M

pOH = −log (0.0625) = 1.20

pH = 12.80

c) **(4 points)** If 5 g of solid NaOH were added to you 2L buffer solution, what would the be resulting change in pH of the buffer solution ?

5g = 0.125 moles = 0.0625 M base

$[NH_4^+]$ = 0.732 − 0.0625 = 0.6695

$[NH_3]$ = 0.530 + 0.0625 = 0.5925

pH = 9.26 + log $\frac{.5925}{.6695}$

= 9.21

pH increases from 9.12 to 9.21 or 0.09 pH units

d) **(1 point)** Is the buffer effective?

Yes!