Exam II – CHM 152/154 – Fall 2005 – Dr. Smith – Form A

CRIB

1. Which of the following statements is FALSE regarding equilibrium?
   a. Equilibrium can be reached starting with products only.
   b. Equilibrium is when the concentration of products equals the concentration of reactants.
   c. Equilibrium does not depend on the initial concentrations.
   d. Equilibrium is when the forward rate equals the reverse rate.
   e. Equilibrium can be reached starting with reactants only.

2. What is the correct equilibrium constant expression in concentrations for the following reaction?
   $2\ NO_2(g) \leftrightarrows 2\ NO(g) + O_2(g)$
   a. $K_c = [O_2][2\ NO] / [2\ NO_2]$
   b. $K_c = [O_2][NO]^2 / [NO_2]^2$
   c. $K_c = [NO_2]^2 / [NO]^2\ [O_2]$
   d. $K_c = [O_2][2\ NO]^2 / [2\ NO_2]^2$
   e. $K_c = [NO]^2 / [NO_2]^2$

3. If K_c for the following reaction is 823, calculate K_p if T = 25.0°C. $2\ Cu_2S(s) + 3\ O_2(g) \leftrightarrows 2\ Cu_2O(s) + 2\ SO_2(g)$

$K_p = 823\ [(0.0821\ Latm/molK)(298K)]^{-1} = 33.6$

4. Consider this reaction $H_2CO_3 + H_2O \leftrightarrows H_3O^+ + HCO_3^-$ with $K_c = 2.5 \times 10^{-4}$. Now consider this reaction $HCO_3^- + H_2O \leftrightarrows H_3O^+ + CO_3^{2-}$ with $K_c = 5.6 \times 10^{-11}$. What is the equilibrium constant for this reaction $H_2CO_3 + 2\ H_2O \leftrightarrows 2\ H_3O^+ + CO_3^{2-}$?

Note the reaction in question is the sum of the other two reactions, so the desired K will be the product of the other two constants. $KK = 1.4 \times 10^{-14}$

5. For the reaction $2\ HI(g) \leftrightarrows H_2(g) + I_2(g)$ $K_c = 12.3$. If $[H_2] = [I_2] = [HI] = 3.21 \times 10^{-3}$ M, which one of the following statements is true?
   a. The concentrations of H_2 and HI will decrease as the system approaches equilibrium
   b. The concentrations of H_2 and I_2 will increase as the system approaches equilibrium
   c. The system is at equilibrium, so the concentrations will not change
   d. The concentration of HI will rise as the system approaches equilibrium
   e. We need pressures to solve this problem

   $Q = (3.21 \times 10^{-3})\ (3.21 \times 10^{-3}) / (3.21 \times 10^{-3})^2 = 1$
   Q < K so reaction goes forward so there will be more products

6. What is the initial concentration of PCl_5 if at equilibrium $[Cl_2] = 0.50$. Given: $PCl_5 \leftrightarrows PCl_3 + Cl_2$ and K = 1.00? *(Hint: ICE table)*

The initial $[PCl_5]$ is our unknown. The change is 0.50. So the equilibrium $[PCl_5] = x-0.50$
From the ICE table: $1.00 = (0.50)^2 / (x-0.50)$

Solving for x = 0.75M = $[PCl_5]$ initial

7. For this reaction: $3\ H_2(g) + N_2(g) \leftrightarrows 2\ NH_3(g)$ there is more ammonia gas present at higher temperatures than at lower temperatures. Is the forward reaction exothermic or **endothermic? Heat is a reactant – when added the rxn goes to products.**

8. How will the following affect the equilibrium for this reaction: $3\ H_2(g) + N_2(g) \leftrightarrows 2\ NH_3(g)$ 1) removing nitrogen gas 2) increasing the volume 3) decreasing pressure
   a. **1) shift left, 2) shift left, 3) shift left**
   b. 1) shift right, 2) shift right, 3) shift right
   c. 1) shift left, 2) shift left, 3) shift right
   d. 1) shift left, 2) shift right, 3) shift right
   e. 1) shift right, 2) shift left, 3) shift left

9. Which of the following factors can change the value of the equilibrium constant?
   a. pressure b. concentration c. catalyst **d. temperature** e. volume

10. What is the conjugate acid of HCO_3^- ? **H_2CO_3**

11. What is the hydronium ion concentration in a 0.00255M sodium hydroxide solution?

$[OH^-] = 0.00255M$
$K_w / 0.00255M = [H^+] = 3.92 \times 10^{-12}M$

12. What is the hydroxide ion concentration if the pH is 8.88?

$8.88 = -\log[H^+]$
$[H^+] = 1.32 \times 10^{-9}M$
$K_w / [H^+] = [OH^-] = 7.59 \times 10^{-6}M$

13. What is the pH of 24.5 mL of 3.33×10^{-5}M nitric acid solution?

$[H^+] = 3.33 \times 10^{-5}M$
$pH = -\log 3.33 \times 10^{-5} = 4.48$

14. Which of the following is NOT a strong acid?
   a. HCl b. H_2SO_4 c. $HClO_4$ **d. H_3PO_4** e. HNO_3

15. The pH of a 0.115M weak monoprotic acid solution is 2.25. Calculate K_a for this acid.

$2.25 = -\log[H^+]$
$[H^+] = 0.00562M$
Set up an ICE table with the generic acid dissociation equation $HA \leftrightarrows H^+ + A^-$
Initial [HA] = 0.115M, equilibrium[H^+] = 0.00562M which also equals $[A^-]$
[HA] at equilibrium is 0.115 – 0.00562 = 0.109M
$K_a = (0.00562)^2 / (0.109) = 2.89 \times 10^{-4}$

16. Which of the following is NOT true as an acid gets stronger?
   a. The hydronium ion concentration increases
   b. The pH decreases

c. The K_a value decreases
d. Percent ionization increases
e. The hydroxide ion concentration decreases

17. At a certain temperature, K_a for acetic acid is 2.4×10^{-6}. What is the pH of a 0.155M solution of acetic acid at this temperature?

Set up ICE table with $CH_3COOH \leftrightarrows CH_3COO^- + H^+$
Initial acid is 0.155 and x is the change.
Subbing into K_a and using the approximation you get $2.4 \times 10^{-6} = x^2 / 0.155$
Solving for $x = 6.099 \times 10^{-4}$ and check approximation – is good!
pH = -log x = 3.2

18. Which one of the following salts will form an acidic solution when dissolved in water?
a. $NaNO_3$
b. KF
c. $LiCH_3COO$
d. $NaHCO_3$
e. NH_4Cl (a is neutral, b, c and d are basic)

19. Which of the following definitions are correct?
a. A Bronsted acid is a proton acceptor. A Lewis acid is an electron acceptor.
b. A Bronsted acid is a proton donor. A Lewis acid is an electron donor.
c. A Bronsted acid is a proton acceptor. A Lewis acid is an electron donor.
d. A Bronsted acid is a proton donor. A Lewis acid is an electron acceptor.

20. Phosphoric acid can donate three H^+ ions. Which H^+ is the easiest to pull off?
a. **the first** b. the second c. the third d. they are all the same

Bonus 1: What is the pH for a 0.42M solution of LiCN? Given: K_a for HCN is 4.9×10^{-8}

LiCN is a soluble salt and ionized completely into Li^+ and CN^- ions. CN^- is the conjugate base for the weak acid HCN. Bases gain H from water. The reaction is $CN^- + H_2O \leftrightarrows HCN + OH^-$
Set up ICE table with initial $[CN^-] = 0.42$ and the change is x.
This is a basic reaction so we need $K_b = K_w / K_a = 2.04 \times 10^{-7}$
Solving for x: $2.04 \times 10^{-7} = x^2 / 0.42$
$x = 2.928 \times 10^{-4}$
pOH = -log x = 3.53
pH = 14 – 3.53 = 10.47 rounding to 2 sig dig = 1.0×10^1

Bonus 2: What is the pH for a 6.25×10^{-3}M solution of barium hydroxide?

$Ba(OH)_2$ gives off 2 hydroxides for every mole
$[OH^-] = 2 (6.25 \times 10^{-3}M) = 0.0125M$
pOH = -log 0.0125 = 1.90
pH = 14.00 – 1.90 = 12.10