MC, Chem1B, Fall15, Test2

KEY

Read questions carefully before answering. No outside paper is allowed. Write **set up equation** for a mathematical problem, then put the raw data with **units**, before showing the calculation. Use the reverse side of your answer paper as scratch. Use the periodic table and important constants charts provided. (Total points = 66 + (16*3=)48 = 114).

Show your calculation with set up and units (when appropriate)

1) The initial rate of the reaction A + B -----> C was measured at several different concentrations of the reactants. Following **formal** methods, (a) calculate the rate law for the reaction (6 pts.) and (b) The magnitude of the rate constant (2 pts.).

1) $rate = 4\times10^{-3}[A]^2$

Experiment	Initial Concentrations [A] (M)	[B] (M)	Initial Rate ($M\ s^{-1}$)
1	0.10	0.10	4.0×10^{-5}
2	0.10	0.20	4.0×10^{-5}
3	0.20	0.10	16.0×10^{-5}

$$\frac{rate_1}{rate_2} = \frac{k[0.1]^n[0.1]^m}{k[0.1]^n[0.2]^m} = \frac{4.0\times10^{-5}}{4.0\times10^{-5}}$$

$$\left(\frac{0.1}{0.2}\right)^m = 1 \qquad m = 0$$

$$\frac{rate_1}{rate_3} = \frac{k[0.1]^n[0.1]^m}{k[0.2]^n[0.1]^m} = \frac{4\times10^{-5}}{16\times10^{-5}}$$

$$\left(\frac{0.1}{0.2}\right)^n = \frac{1}{4} \qquad n = 2$$

$$rate = k[A]^2[B]^0 = k[A]^2$$

$$4.0\times10^{-5} = k[0.1]^2$$

$$k = 4\times10^{-3}$$

$$rate = 4\times10^{-3}[A]^2$$

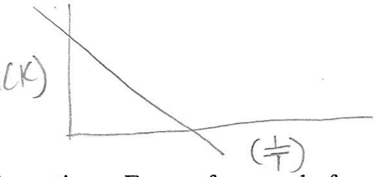
2) In the reaction N_2 (g) + 3 H_2 (g) ↔ 2 NH_3 (g), if the K_c = 9.60 at 573 K, then calculate the Kp at this temperature (4 pts.).

2) $K_p = 4.3\times10^{-3}$

$$K_p = K_c(RT)^{\Delta n} \qquad R = 0.0821\ atm\ L/mol\ K$$

$$K_p = 9.6\,(0.0821\ atm\,L/mol\,K \cdot 573\ K)^{2-4}$$

$$K_p = 4.3\times10^{-3}$$



3) Activation energies of reactions, E_a, are frequently found graphically. The Arrhenius equation: $\ln(k) = (-E_a)/RT + \ln(A)$ is used. Values of k, the rate constant, are measured at various temperatures, then ln k and 1/T are calculated and plotted. In one particular experiment the **slope** of the st. line (obtained by plotting ln(k) and 1/T) was found to be -30000 K. Calculate the energy of activation (in calories) of the reaction (6 pts.):

3) $E_a = 59616.4 \frac{calories}{mol}$

$\ln(K) = \frac{(-E_a)}{R} \cdot \frac{1}{T} + \ln(A)$

$y = m \cdot x + b$

$m = slope = \frac{-E_a}{R}$

$-30000 K = \frac{-E_a}{R}$ $R = 8.3145 J/mol K$

$-30000 K = \frac{-E_a}{8.3145 J/mol K}$

$E_a = 249435 Joules/mol$

$\frac{249435 joules}{mol} \cdot \frac{1 calorie}{4.184 joules}$

$E_a = 59616.4 Calories/mol$

4) Equilibrium was established when a mixture of 0.20 mol of NO(g), 0.10 mol of $H_2(g)$, and 0.20 mol of $H_2O(g)$ is placed in a 2.0-L vessel at 400 K. The equilibrium reaction is : $2 NO(g) + 2 H_2(g)$ <—-> $N_2(g) + 2 H_2O(g)$. If at equilibrium [NO] = 0.062 *M*, then calculate K_P. (10 pts.)

4) $K_p = 19.9$

$2 NO(g) + 2 H_2(g) \rightleftharpoons N_2(g) + 2 H_2O(g)$

I	$\frac{0.2 mol}{2 L}$	$\frac{0.1 mol}{2 L}$	0 M	$\frac{0.2 mol}{2 L}$
C	$-2x$	$-2x$	$+x$	$+2x$
E	$\frac{0.2}{2} M - 2x$	$\frac{0.1}{2} M - 2x$	x	$\frac{0.2}{2} M + 2x$

$[NO] = \frac{0.2}{2} M - 2x = 0.062 M$

$x = 0.019$

$K_c = \frac{[H_2O]^2 [N_2]}{[H_2]^2 [NO]^2} = \frac{(\frac{0.2}{2} + 2 \cdot 0.019)^2 (0.019)}{(\frac{0.1}{2} - 2 \cdot 0.019)^2 (\frac{0.2}{2} - 2 \cdot 0.019)^2} = 653.7$

$K_p = K_c (RT)^{\Delta n}$

$K_p = 653.7 (0.0821 \frac{atm L}{mol K} \cdot 400 K)^{(3-4)}$

$K_p = 19.9$

5) Calculate the pOH of a solution if 1.35 moles of HI is in 530.00 mL of water. (6 pts.)

5) $pOH = 14.4$

$[HI] = \frac{1.35 mol}{0.53 L}$

$[HI] = 2.55 M$

$HI (aq) \rightleftharpoons H^+ (aq) + I^- (aq)$

I	2.55 M	0 M	0 M
C	−2.55 M	+2.55 M	+2.55 M
E	2.55 M − 2.55 M	+2.55 M	+2.55 M

b/c HI is a strong acid, it dissociates completely.

so $[H^+] = 2.55 M$

$pH = -\log(2.55) = -0.4$

$pOH = 14 - pH = 14.4 = pOH$

$NH_3 \leftrightarrow NH_4^+$

6) Ammonia is a weak base with pKb = 4.74 at 25°C. Calculate the pH of a 0.2 M ammonia solution in water at that temperature (8 pts.). 6) pH=11.3

$NH_3 (aq) + H_2O (l) \rightleftharpoons NH_4^+ (aq) + OH^- (aq)$

I	0.2M	0	0
C	−x	+x	+x
E	0.2M − x	x	x

$pK_b = 4.74 = -\log(K_b)$

$K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{0.2}$ b/c x is negligably small compared to 0.2.

$x = 1.9 \times 10^{-3} = [OH^-]$

$pOH = -\log(1.9 \times 10^{-3}) = 2.7 \rightarrow 14 - pOH = 11.3 = pH$

7) Calculate the pH of a buffer solution that contains 0.820 grams of sodium acetate and 0.01 moles of acetic acid in 100 ml of water. The Ka of acetic acid is 1.77×10^{-5} (8 pts). 7) pH = 4.75

$pH = pKa + \log\left(\frac{[A^-]}{[HA]}\right)$

$Na^+C_2H_3O_2^-(s) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq)$

E −s +s +s

$Na^+C_2H_3O_2^- = 0.82g \cdot \frac{1 mol}{82g} = 0.01 mol$

$C_2H_4O_2 = 0.01 mol$

$pH = -\log(1.77 \times 10^{-5}) + \log\left(\frac{0.01}{0.01}\right) = 4.75 = pH$

8) K_{sp} of Ag_2CrO_4 in water at 20 °C is 1.9×10^{-12} M. Calculate it's solubility in gram per 0.1 L of solution (8 pts.). 8) $\frac{0.0026g}{0.1L}$

$Ag_2CrO_4(s) \rightleftharpoons 2Ag^{2+}(aq) + CrO_4^{2-}(aq)$

−s +2s +s

$K_{sp} = 1.9 \times 10^{-12} M = [CrO_4^{2-}][Ag^{2+}]^2 = (s)(2s)^2$

$s = \frac{7.8 \times 10^{-5} mol}{L} \cdot \frac{2 \cdot 108 + 52 + 4 \cdot 16}{1 mol\ Ag_2CrO_4} = \frac{0.026g}{1L}$

$= \frac{0.0026g}{0.1L}$

9) Calculate the molar solubilty of CaF_2 at 25°C in a solution that is 0.010 M in $Ca(NO_3)_2$. Ksp for CaF_2 = 3.9×10^{-11}. Show your calculation with ICE chart. (8 pts.)

9) 3.12×10^{-5} M

$Ca(NO_3)_{2(s)} \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq)$

$-0.01M \quad +0.01M \quad +2 \cdot 0.01M$

$CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2F^-(aq)$

	$CaF_2(s)$	$Ca^{2+}(aq)$	$2F^-(aq)$
I		0.01M	
C	−s	+s	+2s
E	−s	0.01+s	2s

$K_{sp} = [Ca^{2+}][F^-]^2 = (0.01+s)(2s)^2 = 3.9 \times 10^{-11}$

$s = 3.12 \times 10^{-5}$ M

MULTIPLE CHOICE. On your scantron, start answering from number as the question number. Select the one alternative that best completes the statement or answers the question (3 pts each).

10) The rate of a reaction depends on __________.

10) D

A) collision energy
B) collision orientation
C) collision frequency
D) all of the above
E) none of the above

11) A catalyst can increase the rate of a reaction __________.

11) A

A) by providing an alternative pathway with a lower activation energy
B) by changing the value of the frequency factor (A)
C) by increasing the overall activation energy (E_a) of the reaction
D) by lowering the activation energy of the reverse reaction
E) All of these are ways that a catalyst might act to increase the rate of reaction.

12) How does the reaction quotient of a reaction (Q) differ from the equilibrium constant (K_{eq}) of the same reaction?

12) B

A) K does not depend on the concentrations or partial pressures of reaction components.
B) Q is the same as K_{eq} when a reaction is at equilibrium.
C) Q does not depend on the concentrations or partial pressures of reaction components.
D) K_{eq} does not change with temperature, whereas Q is temperature dependent.
E) Q does not change with temperature.

13) The equilibrium constant for reaction 1 is K. The equilibrium constant for reaction 2 is __________. 13) E

(1) SO_2 (g) + (1/2) O_2 (g) $\rightleftharpoons$ SO_3 (g)

(2) $2SO_3$ (g) $\rightleftharpoons$ $2SO_2$ (g) + O_2 (g)

A) 1/2K B) K^2 C) $-K^2$ D) 2K E) $1/K^2$

14) The reaction below is exothermic: 14) B

$2SO_2$ (g) + O_2 (g) $\rightleftharpoons$ $2SO_3$ (g)

Le Châtelier's Principle predicts that __________ will result in an increase in the number of moles of SO_3 (g) in the reaction container.

A) removing some oxygen
B) increasing the pressure
C) increasing the volume of the container
D) decreasing the pressure
E) increasing the temperature

15) The equlibrium reaction $Co(H_2O)_6^{2+}$ (aq) (Pink) + 4 Cl^- (aq) <-> $CoCl_4{}^{2-}$ (aq) (Blue) + 6 H_2O(l) turns pink when placed in ice water mixture but turns blue in hot water. The reaction, as shown, is: 15) B

A) Nonthermic
B) Endothermic
C) Exothermic
D) Insufficient data

16) In which of the following aqueous solutions does the weak acid exhibit the highest percentage ionization? 16) D

A) 0.01 M HClO ($K_a = 3.0 \times 10^{-8}$)
B) 0.01 M HNO_2 ($K_a = 4.5 \times 10^{-4}$)
C) 0.01 M $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$)
D) 0.01 M HF ($K_a = 6.8 \times 10^{-4}$)
E) These will all exhibit the same percentage ionization.

17) Which of the following aqueous solutions has the highest [OH^-]? 17) E

A) a solution with a pOH of 12.0
B) a 1×10^{-3} M solution of NH_4Cl
C) a 1×10^{-4} M solution of HNO_3
D) a solution with a pH of 3.0
E) pure water

18) A 0.1 M aqueous solution of ________ will have a pH of 7.0 at 25.0 °C. 18) B

NaOCl KCl NH_4Cl $Ca(OAc)_2$

A) NaOCl
B) KCl
C) NH_4Cl
D) $Ca(OAc)_2$
E) KCl and NH_4Cl

19) Which one of the following pairs cannot be mixed together to form a buffer solution? 19) E

A) $NaC_2H_3O_2$, HCl ($C_2H_3O_2^-$ = acetate)
B) NH_3, NH_4Cl
C) KOH, HF
D) H_3PO_4, KH_2PO_4
E) RbOH, HBr

20) In which of the following aqueous solutions would you expect AgCl to have the lowest solubility? 20) A

A) 0.020 M $BaCl_2$ → 0.04 Cl
B) pure water
C) 0.020 $AgNO_3$ → 0.02 Ag
D) 0.020 KCl → 0.02 Cl
E) 0.015 NaCl → 0.015 Cl

TRUE/FALSE. On the scantron, select answer 'A' if the statement is true and 'B' if the statement is false (3 pts each).

21) Units of the rate constant of a reaction are independent of the overall reaction order. T or F

22) In an exothermic equilibrium reaction, increasing the reaction temperature favors the formation of reactants. T or F

23) H_2SO_3 and H_2SO_4 are considered an acid-base conjugate pair. T or F

24) For any buffer system, the buffer capacity depends on the amount of acid and base from which the buffer is made. T or F

25) The effect of a catalyst on a chemical reaction is to react with product, effectively removing it and shifting the equilibrium to the right. T or F