

**GENERAL CHEMISTRY 2
SECOND HOUR EXAM**

Name _____ Version 4 _____

Panthersoft ID _____

Signature _____

Part 1 _____ (20 points)

Part 2 _____ (44 points)

Part 3 _____ (36 points)

TOTAL _____ (100 points)

Unless otherwise stated, you may assume $T = 25.0\text{ }^{\circ}\text{C}$ in all of the problems below.

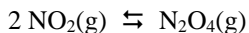
Do all of the following problems. Show your work.

Part 1. Multiple choice. Circle the letter corresponding to the correct answer. There is one and only one correct answer per problem. [4 points each]

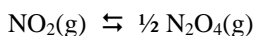
1) Which of the following will cause a change in the numerical value for K_C ?

- a) A change in temperature
- b) A change in pressure
- A** c) A change in volume
- d) Both b and c
- e) Both a and b and c

2) The numerical value for the equilibrium constant for the reaction



is $K_C = 220$, at $T = 25.^\circ\text{C}$. The numerical value for the equilibrium constant for the reaction



at $T = 25.^\circ\text{C}$ is

- a) $K_C = 220$.
- b) $K_C = 110$.
- C** c) $K_C = 14.8$
- d) $K_C = 4.5 \times 10^{-3}$
- e) Cannot tell from the information given in the problem

3) Consider the following chemical reaction



A system containing $\text{N}_2(\text{g})$, $\text{O}_2(\text{g})$, and $\text{NO}(\text{g})$ is initially at equilibrium at $T = 25.^\circ\text{C}$. If the temperature of the system is changed to $T = 80.^\circ\text{C}$ under conditions of constant volume, which of the following will occur as the system reestablishes equilibrium?

- a) Moles of $\text{N}_2(\text{g})$ will increase
- b) Moles of $\text{O}_2(\text{g})$ will increase
- C** c) Moles of $\text{NO}(\text{g})$ will increase
- d) Both a and b
- e) Both a and b and c

4) An Arrhenius base

- a) forms $\text{H}^+(\text{aq})$ when added to liquid water
- b) forms $\text{OH}^-(\text{aq})$ when added to liquid water
- B** c) donates a proton in an acid-base reaction
- d) accepts a proton in an acid-base reaction
- e) is an electron pair donor

5) A weak diprotic acid has $K_{a1} = 4.6 \times 10^{-9}$. Which of the following are possible values for K_{a2} ?

- a) 3.7×10^{-3}
- b) 7.7×10^{-7}
- C** c) 2.5×10^{-12}
- d) Both b and c
- e) Both a and b and c

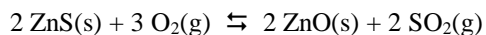
Version 1: C, B, A, C, C

Version 2: B, B, C, E, A

Version 3: B, C, A, D, A

Part 2. Short answer.

1) For the following chemical reaction write the expressions for K_C , K_p , and K . If it is not possible to write an expression, then indicate this by writing n/a (not available). [8 points]



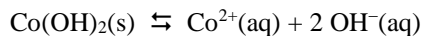
$$K_C = \frac{[\text{SO}_2]^2}{[\text{O}_2]^3}$$

$$K_p = \frac{(p_{\text{SO}_2})^2}{(p_{\text{O}_2})^3}$$

$$K = \frac{(p_{\text{SO}_2})^2}{(p_{\text{O}_2})^3}$$

Note that versions 2 and 3 of the exam have the reverse of the above reaction, and so for those versions the numerator and denominator should be switched.

2) For the chemical reaction



$$\Delta H^\circ_{\text{rxn}} = + 21.5 \text{ kJ/mol}$$

$$\Delta S^\circ_{\text{rxn}} = - 213.6 \text{ J/mol}\cdot\text{K}$$

$$\Delta G^\circ_{\text{rxn}} = + 85.5 \text{ kJ/mol}$$

Based on the above information, find the numerical value for K for the above reaction. Assume $T = 25.^\circ\text{C}$. [8 points]

$$\ln K = -\frac{\Delta G^\circ_{\text{rxn}}}{RT} = -\frac{(85.5 \text{ kJ/mol})(1000 \text{ J/kJ})}{(8.314 \text{ J/mol}\cdot\text{K})(298. \text{ K})} = - 34.51$$

$$\text{So } K = e^{-34.51} = 1.0 \times 10^{-15}$$

3) For each of the following questions circle the correct answer. There is only one correct answer per problem. [4 points each]

a) The value for Q_C for a system at equilibrium

$Q_C > K_C$

$Q_C = K_C$

$Q_C < K_C$

b) An insoluble base

Co(OH)_2

RbOH

Sr(OH)_2

c) The pH of a 0.040 M aqueous solution of analinium bromide, $\text{C}_6\text{H}_5\text{NH}_4\text{Br}$, the salt of a strong acid and a weak base.

$\text{pH} < 7.0$

$\text{pH} = 7.0$

$\text{pH} > 7.0$

4) An aqueous solution has $\text{pH} = 3.71$ $[\text{OH}^-] = 2.4 \times 10^{-11} \text{ M}$ sed on this, find the values for $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for the solution.

[4 points each]

$[\text{H}_3\text{O}^+] = \underline{1.9 \times 10^{-4} \text{ M}}$

$[\text{OH}^-] = \underline{5.1 \times 10^{-11} \text{ M}}$

$\text{pH} + \text{pOH} = 14.00$ and so $\text{pOH} = 14.00 - \text{pH} = 14.00 - 3.71 = 10.29$

$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.71} = 1.9 \times 10^{-4} \text{ M}$

$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-10.29} = 5.1 \times 10^{-11} \text{ M}$

Version 1: $[\text{H}_3\text{O}^+] = 4.2 \times 10^{-4} \text{ M}$ Version 2: $[\text{H}_3\text{O}^+] = 1.4 \times 10^{-3} \text{ M}$ Version 3: $[\text{H}_3\text{O}^+] = 3.6 \times 10^{-3} \text{ M}$
 $[\text{OH}^-] = 2.4 \times 10^{-11} \text{ M}$ $[\text{OH}^-] = 7.1 \times 10^{-12} \text{ M}$ $[\text{OH}^-] = 2.8 \times 10^{-12} \text{ M}$

5) Consider the following four weak bases

ammonia NH_3 $K_b = 1.8 \times 10^{-5}$
 aniline $\text{C}_6\text{H}_5\text{NH}_2$ $K_b = 3.8 \times 10^{-10}$

methylamine CH_3NH_2 $K_b = 4.4 \times 10^{-4}$
 pyradine $\text{C}_5\text{H}_5\text{N}$ $K_b = 1.7 \times 10^{-9}$

For each of the following questions circle the correct answer. There is only one correct answer per problem. [4 points each]

a) The strongest base from the choices below

NH_3

$\text{C}_6\text{H}_5\text{NH}_2$

CH_3NH_2

$\text{C}_5\text{H}_5\text{N}$

b) The strongest conjugate acid from the choices below

NH_4^+

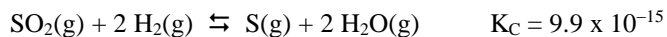
$\text{C}_6\text{H}_5\text{NH}_3^+$

CH_3NH_3^+

$\text{C}_5\text{H}_5\text{NH}^+$

Part 3. Problems.

1) In the gas phase sulfur dioxide (SO₂) and hydrogen (H₂) will exist in equilibrium with sulfur (S) and water (H₂O). The reaction that occurs is as follows



A system initially has [SO₂] = 0.074 M, [H₂] = 0.053 M, and [H₂O] = 5.6 x 10⁻⁵ M. There is initially no S(g) in the system.

a) Give an appropriate ICE table for the above system. [8 points]

$$K_C = \frac{[\text{S}][\text{H}_2\text{O}]^2}{[\text{SO}_2][\text{H}_2]^2} = 9.9 \times 10^{-15}$$

	Initial	Change	Equilibrium
S	0	x	x
H ₂ O	5.6 x 10 ⁻⁵	2x	5.6 x 10 ⁻⁵ + 2x
SO ₂	0.074	- x	0.074 - x
H ₂	0.053	- 2x	0.053 - 2x

b) What is the value for [S], the concentration of S(g), when the above system reaches equilibrium. [8 points]

Based on the ICE table and the expression for K_C

$$\frac{(x)(5.6 \times 10^{-5} + 2x)^2}{(0.074 - x)(0.053 - 2x)^2} = 9.9 \times 10^{-15} \quad \text{Assume } x \ll 5.6 \times 10^{-5}$$

$$\text{Then } \frac{(x)(5.6 \times 10^{-5})^2}{(0.074)(0.053)^2} = 9.9 \times 10^{-15}$$

$$x = \frac{(9.9 \times 10^{-15})(0.074)(0.053)^2}{(5.6 \times 10^{-5})^2} = 6.6 \times 10^{-10} \quad \text{Our assumption that } x \ll 5.6 \times 10^{-5} \text{ was good}$$

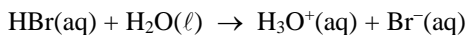
So [S] = 6.6 x 10⁻¹⁰ M

Version 1: [S] = 4.1 x 10⁻¹⁰ M Version 2: [S] = 7.8 x 10⁻¹¹ M Version 3: [S] = 3.6 x 10⁻¹⁰ M

2) Find the pH for the following:

a) A 0.044 M solution of hydrobromic acid (HBr, MW = 80.92 g/mol), a strong monoprotic acid.
[8 points]

HBr is a strong monoprotic acid, and reacts by the process



Since the acid is strong the reaction goes to completion.

$$\text{Therefore } [\text{H}_3\text{O}^+] = \frac{0.044 \text{ mol HBr}}{\text{L soln}} \cdot \frac{1 \text{ mol H}_3\text{O}^+}{1 \text{ mol HBr}} = \frac{0.044 \text{ mol H}_3\text{O}^+}{\text{L soln}}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(0.044) = 1.36$$

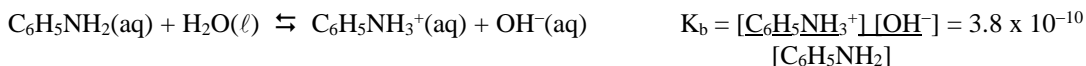
Version 1: pH = 1.09

Version 2: pH = 1.64

Version 3: pH = 1.21

b) A 7.2×10^{-3} M solution of aniline ($\text{C}_6\text{H}_5\text{NH}_2$, MW = 93.13 g/mol), a weak base, with $K_b = 3.8 \times 10^{-10}$
[12 points]

$\text{C}_6\text{H}_5\text{NH}_2$ is a weak base, so



Now we need the ICE table

	Initial	Change	Equilibrium
$\text{C}_6\text{H}_5\text{NH}_3^+$	0	x	x
OH^-	0	x	x
$\text{C}_6\text{H}_5\text{NH}_2$	7.2×10^{-3}	-x	$7.2 \times 10^{-3} - x$

And so $\frac{(x)(x)}{(7.2 \times 10^{-3} - x)} = 3.8 \times 10^{-10}$ Assume $x \ll 7.2 \times 10^{-3}$

Then $\frac{x^2}{(7.2 \times 10^{-3})} = 3.8 \times 10^{-10}$ $x^2 = (3.8 \times 10^{-10})(7.2 \times 10^{-3}) = 2.74 \times 10^{-12}$
 $x = (2.74 \times 10^{-12})^{1/2} = 1.65 \times 10^{-6}$ Our assumption that $x \ll 7.2 \times 10^{-3}$ was good

So $[\text{OH}^-] = 1.65 \times 10^{-6}$ M $\text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}(1.65 \times 10^{-6}) = 5.78$

$\text{pH} + \text{pOH} = 14.00$, so $\text{pH} = 14.00 - \text{pOH} = 14.00 - 5.78 = 8.22$

Version 1: pH = 8.18

Version 2: pH = 8.27

Version 3: 8.36