Chem 112 **Fall 2019** **Exam II** **Whelan**

SID ☐☐☐☐☐☐☐☐ Last Key First Answer

Question 1
8 Points

Consider the following reaction where K_c = **77.52** at 600 K:

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$$

A reaction mixture was found to contain **0.128** moles of **$COCl_2$(g)**, **5.22×10^{-2}** moles of **CO(g)**, and **4.35×10^{-2}** moles of **Cl_2(g)**, in a **1.00 Liter container**.

Indicate **True** (T) or **False** (F) for each of the following:

a) In order to reach equilibrium **$COCl_2$(g)** must be **produced**. T
b) In order to reach equilibrium K_c must **decrease**. F
c) In order to reach equilibrium **CO(g)** must be **produced**. F
d) **Q** is **greater** than **K**. F

Question 2
5 Points

Consider the following equilibrium: $NH_3(g) + HI(g) \rightleftharpoons NH_4I(s)$

Circle the statement that is correct with respect to Kc and Kp for this equilibrium.

☐ $K_c = K_p$ ☐ $K_c > K_p$ ☐ $K_c < K_p$

Question 3
8 Points

The equilibrium constant, K_c, for the following reaction is **3.05×10^{-3} at 262K**.

$$2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$$

Calculate K_c and K_p at this temperature for the following reaction **at 262K**:

$$NO(g) + \tfrac{1}{2}Br_2(g) \rightleftharpoons NOBr(g)$$

R = 0.0821 L.atm.mol^{-1}.K^{-1}

$$K_c = \left(\frac{1}{3.05\times10^{-3}}\right)^{1/2} = \sqrt{327.8}$$

K_c = 18.1

$$K_p = K_c(RT)^{\Delta n} = 18.1\,(0.0821\times262)^{-1/2} = \frac{18.1}{\sqrt{0.0821\times262}}$$

K_p = 3.90

Question 4
9 Points

Consider the following system at equilibrium where $\Delta H°$ = **-16.1 kJ**, and **Kc = 1.54×10^{2}**, at 298 K.

$$2\,NO(g) + Br_2(g) \rightleftharpoons 2\,NOBr(g)$$

If the **TEMPERATURE** on the equilibrium system is suddenly **increased**:

The value of **Kc**
- ☐ Increases
- ☐ Decreases
- ☐ Remains the same

$[Br_2]$
- ☐ Increases
- ☐ Decreases
- ☐ Remains the same

The value of Q
- ☐ Is **greater** than Kc
- ☐ **Is less than** Kc
- ☐ **Is equal to** Kc

Question 5
4 Points

HCN is a weak acid –

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+ + CN^- \qquad K_c = 4.0\times10^{-10} \text{ @ 298K}$$

Addition of OH^- to this equilibrium will cause the **[HCN]** to

a) **Increase**
b) **Decrease**
c) Remain **unchanged**
d) **Impossible** to determine

Question 6
8 Points

Consider the following system at equilibrium where K_c = **6.50x10^{-3}** and $\Delta H°$ = **16.1 kJ/mol** at **298 K**.

$$2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$$

The **production** of **NO(g)** is **favored** by:
Indicate **True (T)** or **False (F)** for each of the following:

a) **Decreasing** the **temperature**. F

b) **Decreasing** the **volume**. F

c) **Adding** Br_2. F

d) **Decreasing** the **pressure** (by changing the volume). T

Question 7
8 Points

The equilibrium constant, K_p, for the following reaction is **0.110** at **298 K**.

$$NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$$

If $\Delta H°$ for this reaction is **92.7 kJ**, what is **the value of K_p at 408 K**?

Must Show Work for Full Credit: R = 8.314 J.mol^{-1}.K^{-1}

$$\ln\frac{K_2}{K_1} = -\frac{\Delta H°}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$K_1 = 0.110$ $\quad K_2 = ?$
$T_1 = 298$ $\quad T_2 = 408$
$\Delta H° = 92{,}700\ J$

$$\ln\frac{K_2}{0.110} = -\frac{92{,}700}{8.314}\left(\frac{1}{408} - \frac{1}{298}\right)$$

$\ln K_2 - \ln 0.110 = -11{,}149.8(-9.047\times10^{-4})$
$\ln K_2 - \ln 0.110 = 10.09$
$\ln K_2 + 2.21 = 10.09$
$\ln K_2 = 10.09 - 2.21$
$= 7.88$

K_p = 2.65×10^3

Question 8
9 Points

a) What is the **conjugate acid** of CO_3^{2-} $\quad HCO_3^-$

b) What is the **conjugate base** of HCO_3^- $\quad CO_3^{2-}$

c) Write a **net ionic equation** to show that **methylamine** behaves as a **Brønsted-Lowry base** in water.

$CH_3NH_2(aq) + H_2O(l)$ $\rightleftharpoons$ $CH_3NH_3^+$ + OH^-

= or ⇔

Question 9
8 Points

Indicate whether each of the following compounds will give an **acidic(A)**, **basic(B)** or **neutral(N)** solution when dissolved in water.

ammonium sulfate: A

sodium nitrate: N

lithium nitrite: B

potassium cyanide: B

Question 10
9 Points

Calculate the **pH** of a **0.369 M** aqueous solution of CH_3COOH, K_a = 1.8x10^{-5} @25°C.

For Full Credit must fill in the ICE Table and Show Work.

	HA	+	H_2O	⇌	H_3O^+	+	A^-
I	0.369		↓		0		0
C	$-x$				x		x
E	$0.369-x$		↓		x		x

$0.369 > 100(1.8\times10^{-5})$

$\therefore\ 0.369 - x \approx 0.369$

HA = CH_3COOH
A^- = CH_3COO^-

$x = \sqrt{0.369(1.8\times10^{-5})}$
$= 2.57\times10^{-3} = [H_3O^+]$

$pH = -\log_{10}(2.57\times10^{-3})$

pH = 2.59

Question 11
9 Points

Calculate the pH of a 0.401 M aqueous solution of C_9H_7N, K_b = 6.30x10^{-10} @25°C.

For Full Credit must fill in the ICE Table and Show Work.

	B(aq)	+	H_2O	⇌	BH^+	+	OH^-
I	0.401		↓		0		0
C	$-x$				x		x
E	$0.401-x$		↓		x		x

$0.401 > 100(6.30\times10^{-10})$

$\therefore\ 0.401 - x \approx 0.401$

B = C_9H_7N
BH^+ = $C_9H_7NH^+$

$x = \sqrt{0.401(6.30\times10^{-10})}$
$= 1.59\times10^{-5} = [OH^-]$

$pOH = -\log_{10}(1.59\times10^{-5})$
$= 4.80$

$pH = 14 - 4.80$

pH = 9.20

Question 12
6 Points

What is the **pOH** of an aqueous solution of **0.102 M** hydrobromic acid? → Strong acid

$HBr(aq) + H_2O(l) = H_3O^+ + Br^-$
0.102 M

$pH = -\log_{10}(0.102) = 0.991$

$pOH = 14 - 0.991$

pOH = 13.01

Question 13 9 Points	In the laboratory, a general chemistry student measured the **pH** of a **0.312 M** aqueous solution of **nitrous acid** to be **1.854**. What is the **Ka** for HNO_2? **For Full Credit must fill in the ICE Table and Show Work.**

	HNO_2	+	H_2O	$\rightleftharpoons$	H_3O^+	+	NO_2^-
I	0.312				0		0
C	$-x$				x		x
E	$0.312-x$				x		x

$$pH = -\log_{10}[H_3O^+]$$
$$-\log_{10}[H_3O^+] = 1.854$$
$$\log_{10}[H_3O^+] = -1.854$$
$$[H_3O^+] = 0.0139 = x$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{x \cdot x}{0.312 - x} = \frac{(0.0139)^2}{0.312 - 0.0139}$$

K_a = 6.48×10^{-4}

Do Not Write Below This Line

Exam II Score	