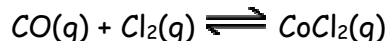


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Last KeyFirst Answer**Question 1**

8 Points

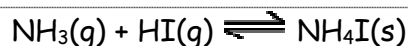
Consider the following reaction where $K_c = 77.52$ at 600 K:A reaction mixture was found to contain 0.128 moles of $\text{COCl}_2(\text{g})$, 5.22×10^{-2} moles of CO(g) , and 4.35×10^{-2} moles of $\text{Cl}_2(\text{g})$, in a 1.00 Liter container.Indicate **True (T)** or **False (F)** for each of the following:

- a) In order to reach equilibrium $\text{COCl}_2(\text{g})$ 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

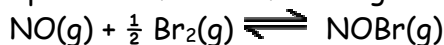
Consider the following equilibrium:

Circle the statement that is correct with respect to K_c and K_p 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.Calculate K_c and K_p at this temperature for the following reaction at 262K:

$$R = 0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$$

$$K_c = \left(\frac{1}{3.05 \times 10^{-3}} \right)^{1/2}$$

$$= \sqrt{327.8}$$

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

$$= 18.1 (0.0821 \times 262)^{-1/2}$$

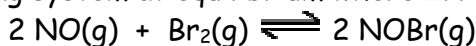
$$= \frac{18.1}{\sqrt{0.0821 \times 262}}$$

$$K_c = \underline{18.1}$$

$$K_p = \underline{3.90}$$

Question 4

9 Points

Consider the following system at equilibrium where $\Delta H^\circ = -16.1 \text{ kJ}$, and $K_c = 1.54 \times 10^2$, at 298 K.If the **TEMPERATURE** on the equilibrium system is suddenly increased:The value of K_c

- Increases
- Decreases
- Remains the same

The value of Q

- Is greater than K_c
- Is less than K_c
- Is equal to K_c

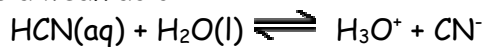
 $[\text{Br}_2]$

- Increases
- Decreases
- Remains the same

Question 5

4 Points

HCN is a weak acid -



$$K_c = 4.0 \times 10^{-10} \text{ @ } 298\text{K}$$

Addition of OH^- to this equilibrium will cause the $[\text{HCN}]$ to

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

Question 6
8 Points

Consider the following system at equilibrium where $K_c = 6.50 \times 10^{-3}$ and $\Delta H^\circ = 16.1 \text{ kJ/mol}$ at 298 K.



The production of $\text{NO}(g)$ is favored by:

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

- a) Decreasing the temperature. F c) Adding Br_2 . F
b) Decreasing the volume. 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.



If ΔH° for this reaction is 92.7 kJ, what is the value of K_p at 408 K?

Must Show Work for Full Credit: $R = 8.314 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$

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

$$\begin{aligned} K_1 &= 0.110 & K_2 &=? \\ T_1 &= 298 & T_2 &= 408 \\ \Delta H^\circ &= 92,700 \text{ J} \end{aligned}$$

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

$$\begin{aligned} \ln K_2 - \ln 0.110 &= -11,149.8 (-9.047 \times 10^{-4}) \\ \ln K_2 - \ln 0.110 &= 10.09 \\ \ln K_2 + 2.21 &= 10.09 \\ \ln K_2 &= 10.09 - 2.21 \\ &= 7.88 \end{aligned}$$

$$K_p = \underline{2.65 \times 10^3}$$

Question 8
9 Points

a) What is the conjugate acid of CO_3^{2-}



b) What is the conjugate base of HCO_3^-



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



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

lithium nitrite:

B

sodium nitrate:

N

potassium cyanide:

B

Question 10
9 Points

Calculate the pH of a 0.369 M aqueous solution of CH_3COOH , $K_a = 1.8 \times 10^{-5}$ @25°C.

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

	HA	+ H ₂ O	⇌	H ₃ O ⁺	+ A ⁻
I	0.369			0	0
C	-x			x	x
E	0.369-x			x	x

$$0.369 > 100(1.8 \times 10^{-5})$$

$$\therefore 0.369 - x \approx 0.369$$



$$x = \sqrt{0.369(1.8 \times 10^{-5})}$$

$$= 2.57 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

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

$$\text{pH} = \underline{2.59}$$

Question 11
9 Points

Calculate the pH of a 0.401 M aqueous solution of $\text{C}_9\text{H}_7\text{N}$, $K_b = 6.30 \times 10^{-10}$ @25°C.

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

	B(aq)	+ H ₂ O	⇌	BH ⁺	+ OH ⁻
I	0.401			0	0
C	-x			x	x
E	0.401-x			x	x

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

$$\therefore 0.401 - x \approx 0.401$$



$$x = \sqrt{0.401(6.30 \times 10^{-10})}$$

$$= 1.59 \times 10^{-5} = [\text{OH}^-]$$

$$\text{pOH} = -\log_{10}(1.59 \times 10^{-5})$$

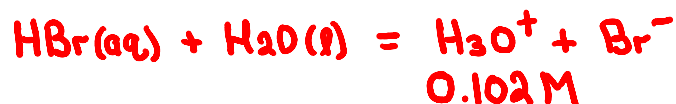
$$= 4.80$$

$$\text{pH} = 14 - 4.80$$

$$\text{pH} = \underline{9.20}$$

Question 12
6 Points

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



Strong acid

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

$$\text{pOH} = 14 - 0.991$$

$$\text{pOH} = \underline{13.01}$$

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 K_a 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

$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] \\ -\log_{10} [\text{H}_3\text{O}^+] &= 1.854 \\ \log_{10} [\text{H}_3\text{O}^+] &= -1.854 \\ [\text{H}_3\text{O}^+] &= 0.0139 = x \end{aligned}$$

$$\begin{aligned} K_a &= \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \\ &= \frac{x \cdot x}{0.312 - x} \\ &= \frac{(0.0139)^2}{0.312 - 0.0139} \end{aligned}$$

$$K_a = \underline{6.48 \times 10^{-4}}$$

Do Not Write Below This Line

Exam II Score