**Chem 112** | **Fall 2019** | **Exam III** | **Whelan**

SID ________ Last Key ________ First Answer ________

**Question 1**
9 Points

The substance **hydrocyanic acid** (HCN) is a weak acid (**Ka = 4.90x10$^{-10}$**).
What is the **pH** of a **0.322 M** aqueous solution of **sodium cyanide**?

| | $CN^-$ | + | $H_2O$ | ⇌ | HCN | + | $OH^-$ |
|---|---|---|---|---|---|---|---|
| I | 0.322 | | | | 0 | | 0 |
| C | -x | | | | x | | x |
| E | 0.322-x | | | | x | | x |

$K_{CN^-} = \frac{1\times10^{-14}}{4.90\times10^{-10}} = 2.04\times10^{-5}$

$0.322 > 100(2.04\times10^{-5})$

$\therefore\ 0.322 - x \approx 0.322$

$x = \sqrt{0.322(2.04\times10^{-5})}$
$= 2.56\times10^{-3} = [OH^-]$

$pOH = -\log_{10}(2.56\times10^{-3})$
$= 2.59$

$pH = 14 - 2.59$

pH = 11.41

**Question 2**
9 Points

With respect to the following acid base reactions, indicate whether the resulting solution will be **acidic**, **basic**, or **neutral**:

1. When **35** mL of **0.40 M HClO** and **35** mL of **0.40 M sodium hydroxide** are combined: B
2. When **35** mL of **0.400** M **nitric acid** and **35** mL of **0.400** M **sodium nitrite** are combined: A
3. When **50** mL of **0.20 M ammonium iodide** and **50** mL of **0.20 M potassium hydroxide** are combined: B

**Question 3**
9 Points

The following questions pertain to a buffer solution that is **0.102** M in $NH_3$ (**ammonia**) and **0.131** M in $NH_4Br$.
Kb($NH_3$) = 1.8x10$^{-5}$ @25°C

1. Write the net ionic equation for the removal of added $H_3O^+$ to this buffer:
   $NH_3(aq)$ + $H_3O^+$ = $H_2O(l)$ + $NH_4^+$
2. What is the **buffer capacity** for **addition** of **strong base**: 0.131
3. The choice of $NH_4^+$ suggests that the **desired pH** is close to: 9.26

**Question 4**
9 Points

Identify buffer solutions from the following list.
*Choose all that apply.*

- [x] 0.30M $HNO_2(aq)$ + 0.25M $KNO_2(aq)$
- [x] 0.15M NaOH(aq) + 0.30M ammonium chloride(aq)
- [x] 0.40M Ammonium chloride + 0.30M Ammonia
- [ ] 0.30M HCl(aq) + 0.30M KF(aq)
- [ ] 0.20M $HNO_3(aq)$ + 0.15M $NaNO_2(aq)$

| Question | |
|---|---|
| **Question 5** 9 Points | **Rank** the following salts from **1-3** in order of **increasing solubility** with **1 being the most soluble** and **3 being the least soluble**. |

- **AgCN** $K_{sp} = 1.2\times10^{-16}$ 3
- **CaF2** $K_{sp} = 3.9\times10^{-11}$ 1
- **Zn3(PO4)2** $K_{sp} = 9.1\times10^{-33}$ 2

**Question 6** 9 Points

The **maximum** amount of **chromium(III) hydroxide** that will dissolve in a **0.255M chromium(III) nitrate** solution is:

Chromium(III) hydroxide: Ksp = $6.70\times10^{-31}$

| | $Cr(OH)_3(s)$ | + | $H_2O$ | ⇌ | $Cr^{3+}$ | + | $3\ OH^-$ |
|---|---|---|---|---|---|---|---|
| I | Some | | ↓ | | 0.255 | | 0 |
| C | -S | | | | S | | 3S |
| E | | | | | 0.255+S | | 3S |

$0.255 + S \approx 0.255$

$K_{sp} = [Cr^{3+}][OH^-]^3$

$6.70\times10^{-31} = 0.255\,(3S)^3$

$27S^3(0.255) = 6.70\times10^{-31}$

$S^3 = \frac{6.70\times10^{-31}}{27(0.255)} = 9.73\times10^{-32}$

$S = \sqrt[3]{9.73\times10^{-32}}$

$4.60\times10^{-11}$ M

**Question 7** 10 Points (4+6)

**Write a balanced net ionic equation** to show why the solubility of **Mn(OH)2(s) increases** in the **presence of a strong acid** and **calculate the equilibrium constant** for the reaction **of this sparingly soluble salt with acid.**

Must show work when calculating K – Ksp Mn(OH)2 = $4.6\times10^{-14}$

$Mn(OH)_2(s)$ + $2\ H_3O^+$ = $Mn^{2+}$ + $4\ H_2O(l)$

$Mn(OH)_2(s) \rightleftharpoons Mn^{2+} + 2OH^-$ $K = 4.6\times10^{-14}$

$2H_3O^+ + 2OH^- \rightleftharpoons 4H_2O(l)$ $K = (1\times10^{-14})^{-2}$

$Mn(OH_2)s + 2H_3O^+ = Mn^{2+} + 4H_2O(l)$ $K = K_{sp}(K_w)^{-2}$

$K = \frac{4.6\times10^{-14}}{(1\times10^{-14})^2}$

K = $4.6\times10^{14}$

**Question 8** 12 Points

**Rank** the following substances from **1-4** in order of **increasing entropy** with **1 being the lowest entropy** and **4 being the highest entropy**.

- $CH_3CHO(g)$ 3
- $(CH_3)_2CO(g)$ 4
- $CH_3OH(g)$ 2
- $HCHO(g)$ 1

**Question 9** 6 Points

Consider the reaction: $2\ H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g)$

Using standard absolute entropies at 298K, calculate the entropy change for the **system** when **2.38** moles of **H2O2(l)** react at standard conditions?

S° (J/K.mol): $H_2O_2(l)$: 109.6 $O_2(g)$: 205.1 $H_2O(l)$: 69.9

$\Delta S^\circ_{RXN} = 2S^\circ H_2O(l) + S^\circ O_2(g) - 2S^\circ H_2O_2(l)$

$= 2(69.9) + 205.1 - 2(109.6)$

$= 125.7$ J/K

$2.38 \times \frac{125.7}{2}$

149.6 J/K

| | |
|---|---|
| **Question 10** 6 Points | Consider the reaction: $2\ CO(g) + 2\ NO(g) \longrightarrow 2\ CO_2(g) + N_2(g)$<br>for which $\Delta H° = -746.6$ kJ and $\Delta S° = -198$ J/K at **298 K**.<br>Calculate **the entropy change** of the **UNIVERSE when 1.57 moles of NO(g)** react under standard conditions at **298 K**. $\Delta S°_{Universe}$ = 1,811 J/K<br>• Is this reaction **reactant** or **product** favored? PRODUCT |
| **Question 11** 6 Points | **Without doing any calculations**, **match** the following **thermodynamic properties** with their appropriate numerical value given on the right for the following **endothermic reaction**.<br>$2\ H_2O(g) + 2\ Cl_2(g) \longrightarrow 4\ HCl(g) + O_2(g)$<br>• $\Delta S_{rxn}$ 1<br>• $\Delta G_{rxn}$ 4<br>• $\Delta S_{universe}$ 5<br>1. > 0<br>2. < 0<br>3. = 0<br>4. > 0 at low T, < 0 at high T<br>5. < 0 at low T, > 0 at high T |
| **Question 12** 6 Points | For the reaction<br>$Fe(s) + 2HCl(aq) = FeCl_2(s) + H_2(g)$ $\Delta H° = -7.4$ kJ and $\Delta S° = 107.9$ J/K<br>a) The **standard free energy change** for the reaction of 1.63 moles of Fe(s) at 291 K, 1 atm would be kJ. -63.2 kJ<br>b) The reaction is (**reactant**, **product**) **favored** under these conditions. PRODUCT<br>*Assume that $\Delta H°$ and $\Delta S°$ are independent of temperature.* |

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| **Exam III Score** | |