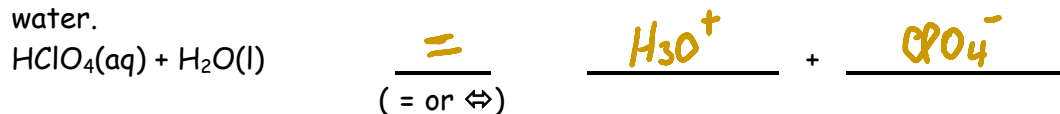
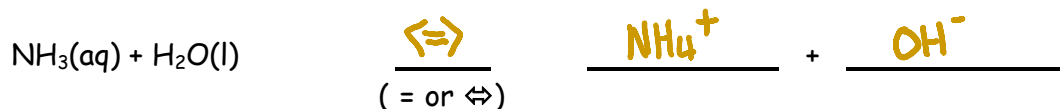


SID Last KeyFirst AnswerQuestion 1
6 Points

a. Write a net ionic equation to show that perchloric acid, behaves as an acid in water.



b. Write a net ionic equation to show how ammonia behaves as a base in water.

Question 2
8 Points

- | | | |
|------------------------------------|----------|----------------|
| a. HNO ₂ | <u>2</u> | 1. Strong Acid |
| b. C ₉ H ₇ N | <u>4</u> | 2. Weak Acid |
| c. CH ₃ COOH | <u>2</u> | 3. Strong Base |
| d. Ba(OH) ₂ | <u>3</u> | 4. Weak Base |

Question 3
6 Points
Circle the appropriate answers

	Acid	K _a
A	Acetic	1.8×10 ⁻⁵
B	Histidine	7.9×10 ⁻⁷
C	Carbonic	4.2×10 ⁻⁷

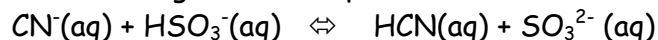
- a. The acid with the **smallest** [H₃O⁺] in a 0.10 M aqueous solution is: A B C
- b. The acid with the **smallest** pK_a: A B C
- c. The acid with the **smallest** pOH in a 0.10 M aqueous solution is: A B C

Question 4
4 Points
A student determines that the value of pK_a for HCN = 9.29.What is the value of K_a? 5.13 × 10⁻¹⁰Question 5
9 Points
The hydroxide concentration in an aqueous solution is 3.5×10⁻² M.

- a. The hydronium ion concentration is: 2.86 × 10⁻¹³ M
- b. The pH of this solution is: 12.54
- c. The pOH is: 1.46

Question 6
6 Points

1. For following net ionic equation:



- Circle the appropriate answer - B-L = Bronsted Lowry

SO ₃ ²⁻	B-L Acid	<u>B-L Base</u>
HSO ₃ ⁻	<u>B-L Acid</u>	B-L Base

2. The formula for the conjugate ACID of CN⁻ is: HCN
3. The formula for the conjugate BASE of HSO₃⁻ is: SO₃²⁻

Question 7 A buffer solution that is **0.436M** in HCN and **0.436M** in KCN has a pH of **9.40**.
6 Points

Addition of **which of the following** would **increase the capacity of the buffer** for added OH^- ?

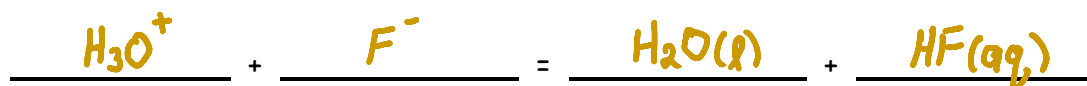
- KCN HCN
 both HCN and KCN pure water
 none of these choices

Question 8 Which of the following aqueous solutions are buffer solutions ?
5 Points

- 0.24 M HI + 0.18 M NaI 0.10 M CH_3COOH + 0.18 M CH_3COOK
 0.27 M NH_4Br + 0.31 M NH_3 0.34 M NH_4NO_3 + 0.39 M NaNO_3
 0.10 M HCl + 0.21 M NaF

Question 9 A buffer solution is made that is **0.434M** in HF and **0.434M** in KF
6 Points

1. If K_a for HF is 7.2×10^{-4} , what is the pH of the buffer solution? 3.14
2. Write the net ionic equation for the reaction that occurs when 0.129 mol HCl is added to 1.00 L of the buffer solution.



Question 10 A buffer solution is **0.414M** in H_2CO_3 and **0.324M** in KHCO_3 . If K_a for H_2CO_3 is 4.2×10^{-7} , what is the pH of this buffer solution?
5 Points
Must show work

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log_{10} \frac{[\text{Base}]}{[\text{Acid}]} \\ &= -\log_{10}(4.2 \times 10^{-7}) + \log_{10} \left(\frac{0.324}{0.414} \right) \\ &= 6.38 + \log_{10}(0.783) \\ &= 6.27 \end{aligned}$$

pH = 6.27

Question 11 A small amount of **strong base** is added to a **buffer** made from **HCN** and **NaCN**. What changes if any will occur to the following.
8 Points

Choose from the following choices:

	Increase significantly Decrease significantly	Increase Decrease	Increase slightly Decrease slightly
1. pH		<u>Increase slightly</u>	
2. [OH ⁻]		<u>Increase slightly</u>	
3. [HCN]		<u>Decrease</u>	
4. [CN ⁻]		<u>Increase</u>	

Question 12 When the nuclide ²¹⁸Po decays to ²¹⁴Pb, what kind of decay does ²¹⁸Po undergo?
6 Points
⁴/₂He. The instability of ²¹⁸Po is probably due to the fact that it has too many Protons.

Question 13 Write a balanced nuclear equation for the following:
6 Points

- ⁵⁹₂₆Fe undergoing beta decay: $\frac{^{59}_{26}\text{Fe}}{26} = \frac{0}{-1}e + \frac{59}{27}\text{Co}$
- ²⁵₁₃Al undergoing positron emission: $\frac{^{25}_{13}\text{Al}}{13} = \frac{0}{+1}e + \frac{25}{12}\text{Mg}$
- ⁴¹₂₀Ca undergoing electron capture: $\frac{^{41}_{20}\text{Ca}}{20} + \frac{0}{-1}e = \frac{41}{19}\text{K}$

Question 14 What **volume** of **hydrogen gas** is produced when **1.33 mol** of **iron** reacts completely according to the following reaction at **25°C** and **1 atm**?
6 Points

iron (s) + hydrochloric acid (aq) = iron(II) chloride (aq) + hydrogen (g)
For full credit you must show work and include a balanced chemical equation.



$$1.33 \text{ mol Fe} \left| \frac{1 \text{ H}_2}{1 \text{ Fe}} \right. = 1.33 \text{ mol H}_2$$

$$PV = nRT$$

$$(1)V = 1.33 \times 0.08205 \times 298$$

$$V = 32.5 \text{ L}$$

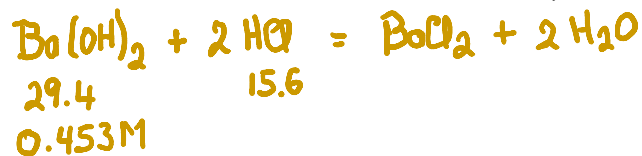
32.5 L

Question 15
8 Points

An aqueous solution of **hydrochloric acid** is standardized by titration with a **0.453 M** solution of **barium hydroxide**.

If **29.4 mL** of base are required to neutralize **15.6 mL** of the acid, what is the **molarity** of the **hydrochloric acid** solution?

For full credit you must show work and include a balanced chemical equation.



$$\# \text{ mol Ba(OH)}_2 = 0.453 \times 0.0294 = 1.33 \times 10^{-2} \text{ mol Ba(OH)}_2$$

$$1.33 \times 10^{-2} \text{ mol Ba(OH)}_2 \left| \begin{array}{l} 2 \text{ HCl} \\ 1 \text{ Ba(OH)}_2 \end{array} \right. = 2.66 \times 10^{-2} \text{ mol HCl}$$

$$M = \frac{2.66 \times 10^{-2}}{0.0156} = 1.71$$

1.71 M

Question 16
5 Points

How many **grams** of **iron(II) bromide** are there in **43.5 mL** of an aqueous solution that has a concentration of **0.166 M**?

Must show work

$$\# \text{ mol FeBr}_2 = 0.166 \times 0.0435 = 7.22 \times 10^{-3}$$

$$\text{FeBr}_2 : 55.85 + 2(79.90) = 216.65 \text{ g} \cdot \text{mol}^{-1}$$

$$7.22 \times 10^{-3} \text{ mol FeBr}_2 \left| \begin{array}{l} 216.65 \text{ g} \\ 1 \text{ mol} \end{array} \right. = 1.56 \text{ g}$$

1.56 grams

Exam III Score