



Question 1 Are the following salts expected to be soluble or insoluble in water?

12 Points

1. $\text{FeF}_3$	Soluble	Insoluble	4. $\text{Ba}(\text{NO}_3)_2$	Soluble	Insoluble
2. $\text{Na}_3\text{PO}_4$	Soluble	Insoluble	5. $\text{FeCO}_3$	Soluble	Insoluble
3. $\text{NH}_4\text{OH}$	Soluble	Insoluble	6. $\text{BaS}$	Soluble	Insoluble

Question 2 Classify each of the following substances:

12 Points

1. HF	B	A) Strong Acid
2. NaI	E	B) Weak Acid
3. $\text{NH}_3$	D	C) Strong Base
4. HCl	A	D) Weak Base
5. NaOH	C	E) Soluble Salt
6. $\text{Cr}_3(\text{PO}_4)_2$	F	F) Insoluble Salt

Question 3 1. The formula for the conjugate base of HCN:

$\text{CN}^-$

6 Points

2. The formula for the conjugate acid of  $\text{CO}_3^{2-}$

$\text{HCO}_3^-$

Question 4 The  $[\text{H}^+]$  in an aqueous solution is found to be  $5.43 \times 10^{-2} \text{M}$ .

8 Points

- The pH of this solution is: 1.265
- The  $[\text{OH}^-]$  of this solution is:  $1.84 \times 10^{-13}$
- The pOH of this solution is: 12.735
- The solution is (circle one) Basic Acidic Neutral

Question 5 What is the expected pH of an aqueous solution of 0.302M hydrocyanic acid (HCN) at  $25^\circ\text{C}$ ?

6 Points

$$[\text{H}^+] = \{\text{Ka}[\text{Acid}]\}^{\frac{1}{2}}$$
$$[\text{H}^+] = \{4 \times 10^{-10}(0.302)\}^{\frac{1}{2}}$$

$$[\text{H}^+] = 1.10 \times 10^{-5}$$

$$\text{pH} = 4.96$$

Question 6 How would the pH of a 0.302M aqueous hydrochloric acid solution compare to the 0.302M aqueous HCN solution at the same temperature.  
4 Points

1.  $\text{pH HCl(aq)} > \text{pH HCN(aq)}$
2.  $\text{pH HCl(aq)} < \text{pH HCN(aq)}$
3.  $\text{pH HCl(aq)} = \text{pH HCN(aq)}$

Question 7 Give the net ionic equation for the following reactions:  
6 Points

1.  $\text{NaOH(aq)} + \text{HNO}_2(\text{aq}) \quad \text{OH}^- + \text{HNO}_2(\text{aq}) = \text{NO}_2^- + \text{H}_2\text{O(l)}$
2.  $\text{NH}_3(\text{aq}) + \text{HCl(aq)} \quad \text{NH}_3(\text{aq}) + \text{H}^+ = \text{NH}_4^+$
3.  $\text{HI(aq)} + \text{LiOH(aq)} \quad \text{H}^+ + \text{OH}^- = \text{H}_2\text{O(l)}$

Question 8 A 1L buffer solution contains 0.112M KF and 0.396M HF. Calculate the expected pH of this solution?  
6 Points

$$[\text{H}^+] = K_a \{ [\text{Weak Acid}] / [\text{Conjugate Base}] \}$$

$$[\text{H}^+] = 7.2 \times 10^{-4} \{ 0.396 / 0.112 \}$$

$$[\text{H}^+] = 2.54 \times 10^{-3}$$

$$\text{pH} = 2.59$$

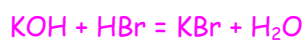
Question 9 The addition of 0.012 moles of HBr to the buffer described in question 8 would result in:  
14 Points

- |  |          |          |              |
|--|----------|----------|--------------|
| 1. pH  | Increase | Decrease | No Change    |
| 2. $[\text{H}_3\text{O}^+]$                                    | Increase | Decrease | No Change    |
| 3. $[\text{OH}^-]$   | Increase | Decrease | No Change    |
| 4. $[\text{F}^-]$  | Increase | Decrease | No Change    |
| 5. $[\text{HF}]$   | Increase | Decrease | No Change    |
| 6. $[\text{HF}] / [\text{F}^-]$                                | Increase | Decrease | No Change    |
| 7. The maximum amount of HCl that this buffer could withstand? |          |          | ~0.112 moles |

Question 10 10 Points The reaction  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$  has a  $K_c = 55.6$  and a  $\Delta H^\circ = -10 \text{ kJ/mol}$  at 696K. The production of  $\text{HI}(\text{g})$  is favor by:

- |  |      |       |
|--|------|-------|
| 1. Decreasing the temperature.                     | True | False |
| 2. Increasing the pressure by changing the volume. | True | False |
| 3. Decreasing the volume.                          | True | False |
| 4. Adding $\text{I}_2$                             | True | False |
| 5. Removing $\text{HI}$                            | True | False |

Question 11 8 Points An aqueous solution of **potassium hydroxide** is standardized by titration with a **0.369 M** solution of **hydrobromic acid**. If **30.5 mL** of base are required to neutralize **13.4 mL** of the acid, what is the molarity of the **potassium hydroxide** solution?

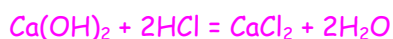


$$0.369 \times 0.0305 = 1.12 \times 10^{-2} \text{ mol KOH}$$
$$1.12 \times 10^{-2} \text{ mol KOH} \times (1 \text{ HBr} / 1 \text{ KOH}) = 1.12 \times 10^{-2} \text{ mol HBr}$$

$$M = 1.12 \times 10^{-2} / 0.0134 = 0.839\text{M}$$

0.839M

Question 12 8 Points How many grams of solid **calcium hydroxide** are needed to exactly neutralize **27.6 mL** of a **1.68 M** **hydrochloric acid** solution? Assume that the volume remains constant.



$$1.68 \times 0.0276 = 4.63 \times 10^{-2} \text{ mol HCl}$$

$$4.63 \times 10^{-2} \text{ mol HCl} \times (1 \text{ Ca}(\text{OH})_2 / 2 \text{ HCl}) = 2.31 \times 10^{-2} \text{ mol Ca}(\text{OH})_2$$

$$2.31 \times 10^{-2} \text{ mol Ca}(\text{OH})_2 \times (74.1\text{g} / 1 \text{ mol}) = 1.72\text{g Ca}(\text{OH})_2$$

1.72g