

This print-out should have 45 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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**001 10.0 points**

Assume that five weak acids, identified only by numbers (1, 2, 3, 4, and 5), have the following ionization constants.

Acid	Ionization Constant $K_a$ value
1	$1.0 \times 10^{-3}$
2	$3.0 \times 10^{-5}$
3	$2.6 \times 10^{-7}$
4	$4.0 \times 10^{-9}$
5	$7.3 \times 10^{-11}$

The anion of which acid is the strongest base?

1. 4

2. 5 **correct**

3. 2

4. 3

5. 1

**Explanation:**

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**002 10.0 points**

The term “ $K_a$  for the ammonium ion” describes the equilibrium constant for which of the following reactions?

1.  $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$  **correct**

2.  $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

3.  $\text{NH}_3 + \text{H}_3\text{O}^+ \rightleftharpoons \text{NH}_4^+ + \text{H}_2\text{O}$

4.  $\text{NH}_4^+ + \text{OH}^- \rightleftharpoons \text{NH}_3 + \text{H}_2\text{O}$

5.  $\text{NH}_4\text{Cl}(\text{solid}) + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{Cl}^-$

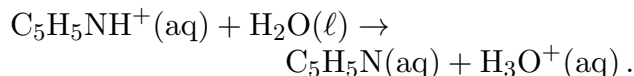
6. The term is misleading, because the ammonium ion is not an acid.

**Explanation:**

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**003 10.0 points**

If the value of  $K_b$  for pyridine is  $1.8 \times 10^{-9}$ , calculate the equilibrium constant for



1.  $5.6 \times 10^{-6}$  **correct**

2.  $1.8 \times 10^{-9}$

3.  $1.8 \times 10^{-16}$

4.  $5.6 \times 10^8$

5.  $-1.8 \times 10^{-9}$

**Explanation:**

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**004 10.0 points**

Which of the following is true in pure water at any temperature?

1.  $K_w$  decreases with increasing temperature.

2.  $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

3.  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$  **correct**

4.  $\text{pH} = 7.0$  or greater than 7.0

5.  $\text{pH} = 7.0$

**Explanation:**

$K_w$  is shown to INCREASE with increasing temperature.  $\text{pH} = 7$  is only true when water is at  $24^\circ\text{C}$ .  $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$ , which increases with temperature.

At high temperatures  $\text{pH}$  can be less than 7. Thus  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$  is the only case that is true.

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**005 10.0 points**

Which is NOT a conjugate acid-base pair?

1.  $\text{H}_2\text{O} : \text{OH}^-$

2. HCl : Cl<sup>-</sup>3. H<sub>3</sub>SO<sub>4</sub><sup>+</sup> : H<sub>2</sub>SO<sub>4</sub>4. H<sub>2</sub> : H<sup>-</sup>5. H<sub>2</sub>SO<sub>4</sub> : SO<sub>4</sub><sup>2-</sup> **correct****Explanation:**

Except for H<sub>2</sub>SO<sub>4</sub> and SO<sub>4</sub><sup>2-</sup>, the members of all of the pairs differ by one proton.

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**006 10.0 points**What is the conjugate acid of NO<sub>3</sub><sup>-</sup>?1. NO<sub>2</sub><sup>-</sup>2. NH<sub>3</sub>3. H<sup>+</sup>4. HNO<sub>3</sub> **correct**5. NO<sub>3</sub><sup>2-</sup>6. OH<sup>-</sup>**Explanation:**

Since the question asks for the conjugate acid, we can assume NO<sub>3</sub><sup>-</sup> is acting as a base. This means that it is a proton acceptor. To form the conjugate acid, it accepts a H making HNO<sub>3</sub>.

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**007 10.0 points**What is [H<sub>3</sub>O<sup>+</sup>] when [OH<sup>-</sup>] = 3.3 × 10<sup>-9</sup> M?1. 1.0 × 10<sup>-7</sup> M2. 3.3 × 10<sup>-9</sup> M3. 3.3 × 10<sup>-5</sup> M4. 3.0 × 10<sup>-6</sup> M **correct**5. 6.6 × 10<sup>-5</sup> M**Explanation:**[OH<sup>-</sup>] = 3.3 × 10<sup>-9</sup> M

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{14}$$

$$\begin{aligned} [\text{H}_3\text{O}^+] &= \frac{K_w}{[\text{OH}^-]} \\ &= \frac{1.0 \times 10^{14}}{3.3 \times 10^{-9}} = 3.0 \times 10^{-6} \text{ M} \end{aligned}$$

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**008 10.0 points**What is [OH<sup>-</sup>] in a 0.0050 M HCl solution?1. 6.6 × 10<sup>-5</sup> M2. 5.0 × 10<sup>-3</sup> M3. 1.0 × 10<sup>-7</sup> M4. 2.0 × 10<sup>-12</sup> M **correct**

5. 1.0 M

**Explanation:**[OH<sup>-</sup>] = 0.0050 M

Since HCl is a strong acid, it completely dissociates and H<sup>+</sup> is 0.0050 M.



$$\begin{aligned} K_w &= [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14} \\ [\text{OH}^-] &= \frac{K_w}{[\text{H}^+]} \\ &= \frac{1 \times 10^{-14}}{0.0050} = 2 \times 10^{-12} \text{ M} \end{aligned}$$

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**009 10.0 points**Which pH represents a solution with 1000 times higher [OH<sup>-</sup>] than a solution with pH of 5?

1. pH = 2

2. pH = 0.005

3. pH = 8 **correct**

4. pH = 1

5. pH = 3

6. pH = 5000

7. pH = 7

8. pH = 4

9. pH = 6

**Explanation:**

pH = 5

$$\text{pOH} = 14 - \text{pH} = 14 - 5 = 9$$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-9} \text{ M}$$

$$\begin{aligned} [\text{OH}^-]_x &= 1000 [\text{OH}^-] = (10^3)(10^{-9} \text{ M}) \\ &= 10^{-6} \text{ M} \end{aligned}$$

$$\text{pOH}_x = -\log(\text{OH}_x) = 6$$

$$\text{pH}_x = 14 - \text{pOH}_x = 14 - 6 = 8$$

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**010 10.0 points**What is the pH of a 0.12 M Ba(OH)<sub>2</sub> aqueous solution?

1. 1.33802

2. 8.7

3. 0.619789

4. 13.3802 **correct**

5. 10.0352

**Explanation:**

$$[\text{Ba}(\text{OH})_2] = 0.15 \text{ M}$$

Ba(OH)<sub>2</sub> is a strong base which dissociates in aqueous solution to produce two moles of OH<sup>-</sup> for every mole of Ba(OH)<sub>2</sub>, so 0.12 M Ba(OH)<sub>2</sub> produces 0.24 M OH<sup>-</sup>.

	Ba(OH) <sub>2</sub>	→	Ba <sup>2+</sup>	+	2 OH <sup>-</sup>
ini	0.12 M		0 M		0 M
Δ	-0.12 M		+0.12 M		2(0.12 M)
fin	0 M		+0.12 M		+0.24 M

$$\text{pH} = 14 - \text{pOH} = 14 - (-\log 0.24) = 13.3802$$

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**011 10.0 points**Hydroxylamine is a weak molecular base with  $K_b = 6.6 \times 10^{-9}$ . What is the pH of a 0.0500 M solution of hydroxylamine?

1. pH = 8.93

2. pH = 7.12

3. pH = 3.63

4. pH = 4.74

5. pH = 9.26 **correct**

6. pH = 9.48

7. pH = 10.37

**Explanation:**Hydroxylamine is a weak base, so use the equation to calculate weak base [OH<sup>-</sup>] concentration (note that this is the approximate equation. Why? Because  $K_b$  is very small and the concentration is reasonable) :

$$\begin{aligned} [\text{OH}^-] &= \sqrt{K_b C_b} \\ &= \sqrt{(6.6 \times 10^{-9})(0.0500)} \\ &= 1.82 \times 10^{-5} \end{aligned}$$

After finding [OH<sup>-</sup>], you can find pH using either method below:

A)

$$\text{pOH} = -\log(1.82 \times 10^{-5}) = 4.74$$

$$\text{pH} = 14 - 4.74 = 9.26$$

or B)

$$\begin{aligned} [\text{H}^+] &= \frac{K_w}{[\text{OH}^-]} \\ &= \frac{1.0 \times 10^{-14}}{1.82 \times 10^{-5}} = 5.52 \times 10^{-10} \end{aligned}$$

$$\text{pH} = -\log(5.52 \times 10^{-10}) = 9.26$$

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**012 10.0 points**What is the pH of a 0.2 M solution of potassium generate (KR-COO)?  $K_a$  for the generic acid (R-COOH) is  $2.7 \times 10^{-8}$ .

1. 10.285

2. 7.000

3. 10.565

4. 10.195

5. 3.565

6. 7.569

7. 6.431

8. 3.435

9. 10.435 **correct**

10. 10.805

**Explanation:**

$$M_{\text{KR-COO}} = 0.2 \text{ M} \quad K_a = 2.7 \times 10^{-8}$$

It's a salt of a weak generic acid (KA). Get it? Generic acid makes generic ions. Ha! This means you need a  $K_b$  for the weak base  $A^-$ . Use  $K_b = \frac{K_w}{K_a}$  and you'll get the  $K_b = 3.7037 \times 10^{-7}$ . You CAN use the approximation for the equilibrium which means that

$$[\text{OH}^-] = \sqrt{K_b \cdot C_{A^-}} = 0.000272166 \text{ M}$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ &= 14 + \log(0.000272166) = 10.4348 \end{aligned}$$

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**013 10.0 points**

At 25° C, the pH of a water solution of a salt of a WEAK acid and a STRONG base is

- less than 7.
- greater than 7. **correct**
- about 7.
- equal to the hydrogen ion concentration.

**Explanation:**

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**014 10.0 points**

What is the pH of a 0.16 M solution of anilinium nitrate ( $\text{C}_6\text{H}_5\text{NH}_3\text{NO}_3$ )?  $K_b$  for aniline is  $4.2 \times 10^{-10}$ .

Your answer must be within  $\pm 0.4\%$   
Correct answer: 2.70956.

**Explanation:**

$$M_{\text{C}_6\text{H}_5\text{NH}_3\text{NO}_3} = 0.16 \text{ M} \quad K_b = 4.2 \times 10^{-10}$$

It's a salt of a weak base (BHX). This means you need a  $K_a$  for the weak acid  $\text{BH}^+$ :

$$\begin{aligned} K_a &= \frac{K_w}{K_b} \\ &= \frac{1.0 \times 10^{-14}}{4.2 \times 10^{-10}} \\ &= 2.38095 \times 10^{-5} \end{aligned}$$

You CAN use the approximation for the equilibrium which means that

$$\begin{aligned} [\text{H}^+] &= \sqrt{K_a \cdot C_{\text{BH}^+}} \\ &= \sqrt{(2.38095 \times 10^{-5})(0.16)} \\ &= 0.0019518 \text{ M} \end{aligned}$$

$$\text{pH} = -\log(0.0019518) = 2.70956$$

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**015 10.0 points**

The pH of lemon juice is approximately 2.4. At this pH, the hydronium ion concentration is closest to which value?

- $2.50 \times 10^{-12} \text{ M}$
- $5.62 \times 10^{-4} \text{ M}$
- $4.00 \times 10^{-3} \text{ M}$  **correct**
- 250 M

**Explanation:**

$$\text{pH} = 2.4, \text{ so}$$

$$M_{\text{H}^+} = 10^{-2.4} = 0.00398107 \text{ M}$$

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**016 10.0 points**

Which solution has the highest pH?

- 0.1 M of  $\text{KHCOO}$ ,  
 $K_a \text{HCOOH} = 1.8 \times 10^{-4}$
- 0.1 M of  $\text{KCl}$ ,  $K_a \text{HCl} = \text{very large}$

3. 0.1 M of  $\text{KCH}_3\text{COO}$ ,  
 $K_{\text{a HC}_2\text{H}_3\text{O}_2} = 1.8 \times 10^{-5}$

4. 0.1 M of  $\text{KNO}_2$ ,  $K_{\text{a HNO}_2} = 4.5 \times 10^{-4}$

5. 0.1 M of  $\text{KClO}$ ,  $K_{\text{a HClO}} = 3.5 \times 10^{-8}$   
**correct**

**Explanation:**

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**017 10.0 points**

What is the pH of a solution that contains 11.7 g of  $\text{NaCl}$  for every 200 mL of solution?

1. 1.0

2.  $10^{-1}$

3. 7.0 **correct**

4.  $1.0 \times 10^{-7}$

**Explanation:**

$m_{\text{NaCl}} = 11.7 \text{ g}$   $V_{\text{soln}} = 200 \text{ mL}$   
 $\text{NaCl}$  completely dissociates in water to give  $\text{Na}^+$  and  $\text{Cl}^-$ , neither of which hydrolyzes and so in aqueous  $\text{NaCl}$  the  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  ions result from autoionization of water.

$$K_{\text{w}} = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1 \times 10^{-7}$$

and  $\text{pH} = -\log[\text{H}_3\text{O}^+] = 7.0$

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**018 10.0 points**

A 0.010 M solution of a weak acid  $\text{HA}$  has a pH of 4.20. What is the pOH of the solution?

1. 14.0

2. None of these

3. 4.20

4. 7.0

5. 9.80 **correct**

**Explanation:**

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**019 10.0 points**

A solution has a pH of 4.35. Find the pOH.

1. 4.35

2. 9.65 **correct**

3. None of these

4. 18.35

**Explanation:**

$\text{pH} = 4.35$

$$\text{pOH} = 14 - \text{pH} = 9.65$$

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**020 (part 1 of 2) 10.0 points**

The pH of an aqueous solution is measured as 1.21. Calculate the  $[\text{H}_3\text{O}^+]$ .

Correct answer: 0.0616595 M.

**Explanation:**

$\text{pH} = 1.21$

$[\text{H}_3\text{O}^+] = ?$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\log[\text{H}_3\text{O}^+] = -\text{pH}$$

$$[\text{H}_3\text{O}^+] = \text{antilog}(-\text{pH})$$

$$= 1 \times 10^{-\text{pH}}$$

$$= 1 \times 10^{-1.21}$$

$$= 0.0616595 \text{ M}$$

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**021 (part 2 of 2) 10.0 points**

Calculate the  $[\text{OH}^-]$ .

Correct answer:  $1.62181 \times 10^{-13} \text{ M}$ .

**Explanation:**

$\text{pH} = 1.21$

$[\text{OH}^-] = ?$

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14} \text{ M}^2$$

$$[\text{OH}^-] = \frac{1 \times 10^{-14} \text{ M}^2}{[\text{H}_3\text{O}^+]}$$

$$= \frac{1 \times 10^{-14} \text{ M}^2}{0.0616595 \text{ M}}$$

$$= 1.62181 \times 10^{-13} \text{ M}$$

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**022 10.0 points**

What is the pH of a solution made by mixing 0.05 mol of NaCN with enough water to make a liter of solution?

$K_a$  for HCN is  $4.9 \times 10^{-10}$  and  $K_w = 1 \times 10^{-14}$ .

Correct answer: 11.0044.

**Explanation:**

$$\begin{aligned} [\text{OH}^-] &= \sqrt{K_b C} \\ &= \sqrt{\frac{K_w}{K_a} C} \\ &= \sqrt{\frac{1 \times 10^{-14}}{4.9 \times 10^{-10}} (0.05)} = 0.00101015 \end{aligned}$$

$$\begin{aligned} [\text{H}^+] &= \frac{K_w}{[\text{OH}^-]} \\ &= \frac{1 \times 10^{-14}}{0.00101015} = 9.89949 \times 10^{-12} \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log[\text{H}^+] \\ &= -\log(9.89949 \times 10^{-12}) = 11.0044 \end{aligned}$$

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**023 10.0 points**

Identify the list in which all salts produce a basic aqueous solution.

1. AgNO<sub>3</sub>, NaCHO<sub>2</sub>, CrI<sub>3</sub>
2. NH<sub>4</sub>Cl, C<sub>6</sub>H<sub>4</sub>NH<sub>3</sub>NO<sub>3</sub>, FeI<sub>3</sub>
3. AlCl<sub>3</sub>, Zn(NO<sub>3</sub>)<sub>2</sub>, KClO<sub>4</sub>
4. CH<sub>3</sub>NH<sub>3</sub>Cl, KNO<sub>3</sub>, NaBz (sodium benzoate)
5. KCH<sub>3</sub>COO, NaCN, KF **correct**

**Explanation:**

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**024 10.0 points**

What is the pH in a solution made by

dissolving 0.100 mole of sodium acetate (NaCH<sub>3</sub>COO) in enough water to make one liter of solution?  $K_a$  for CH<sub>3</sub>COOH is  $1.80 \times 10^{-5}$ .

1. 8.87 **correct**

2. 9.25

3. 5.13

4.  $5.56 \times 10^{-11}$

5. 10.25

6. 5.74

7.  $5.56 \times 10^{-10}$

8.  $1.80 \times 10^{-6}$

9.  $7.46 \times 10^{-6}$

10.  $1.34 \times 10^{-9}$

**Explanation:**

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**025 10.0 points**

A 0.200 M solution of a weak monoprotic acid HA is found to have a pH of 3.00 at room temperature. What is the ionization constant of this acid?

1.  $5.0 \times 10^{-3}$

2.  $2.0 \times 10^{-5}$

3.  $1.0 \times 10^{-6}$

4. 5.30

5.  $5.0 \times 10^{-6}$  **correct**

6.  $1.8 \times 10^{-5}$

7.  $2.0 \times 10^{-9}$

8.  $1.0 \times 10^{-3}$

**Explanation:**

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**026 10.0 points**

What is the percent ionization for a weak acid HX that is 0.40 M?  $K_a = 4.0 \times 10^{-7}$ .

1. 0.00020%
2. 0.050%
3. 0.020%
4. 0.10% **correct**
5. 2.0%

**Explanation:**

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**027 10.0 points**

A 0.28 M solution of a weak acid is 3.5% ionized. What is the pH of the solution?

1. 2.01 **correct**
2. 1.46
3. 5.25
4. 0.55
5. 3.17

**Explanation:**

$M = 0.28 \text{ M}$   $P = 3.5\%$   
 3.5% of the 0.28 M is ionized (contributes to pH), so

$$[\text{H}^+] = (0.28 \text{ M}) \times \frac{3.5}{100} = 0.0098 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -\log(0.0098) = 2.00877$$

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**028 10.0 points**

The pH of 0.010 M aniline(aq) is 8.32.

What is the percentage aniline protonated?

1. 2.1%
2. 0.021% **correct**
3. 0.12%

4. 0.21%

5. 0.69%

**Explanation:**

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**029 10.0 points**

A 20 mL sample of 0.20 M nitric acid solution is required to neutralize 40 mL of barium hydroxide solution. What is the molarity of the barium hydroxide solution?

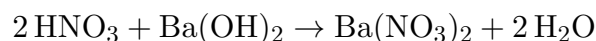
1. 0.050 M **correct**
2. 0.025 M
3. 0.100 M
4. 0.0025 M
5. 0.200 M

**Explanation:**

$$V_{\text{HNO}_3} = 20 \text{ mL} \qquad [\text{HNO}_3] = 0.20 \text{ M}$$

$$V_{\text{Ba(OH)}_2} = 40 \text{ mL}$$

The balanced equation for this neutralization reaction is



We determine the moles of  $\text{HNO}_3$  used:

$$\begin{aligned} ? \text{ mol HNO}_3 &= 0.020 \text{ L soln} \\ &\times \frac{0.20 \text{ mol HNO}_3}{1 \text{ L soln}} \\ &= 0.0040 \text{ mol HNO}_3 \end{aligned}$$

Using the mole ratio from the chemical equation we calculate the moles  $\text{Ba(OH)}_2$  needed to react with 0.0040 mol of  $\text{HNO}_3$ :

$$\begin{aligned} ? \text{ mol Ba(OH)}_2 &= 0.0040 \text{ mol HNO}_3 \\ &\times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol HNO}_3} \\ &= 0.0020 \text{ mol Ba(OH)}_2 \end{aligned}$$

There are 0.0020 moles  $\text{Ba(OH)}_2$  in the 40 mL sample. Molarity is moles solute per liter of solution:

$$\begin{aligned} ? \text{ M Ba(OH)}_2 &= \frac{0.0020 \text{ moles Ba(OH)}_2}{0.040 \text{ L solution}} \\ &= 0.050 \text{ M Ba(OH)}_2 \end{aligned}$$

**030 10.0 points**

When an acid and base neutralize each other, the products are generally water

1. a salt. **correct**
2. a gel.
3. a colloid.
4. an ion.

**Explanation:**

The general format for neutralization reactions is acid + base  $\rightarrow$  salt + water.

**031 10.0 points**

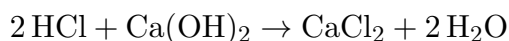
How many moles of  $\text{Ca}(\text{OH})_2$  are needed to neutralize three moles of  $\text{HCl}$ ?

1. three
2. 1.5 **correct**
3. four
4. eight
5. 0.5
6. two
7. six
8. one

**Explanation:**

$$n_{\text{HCl}} = 3 \text{ mol}$$

For acid base neutralization we need one mole of  $\text{H}^+$  for every mole of  $\text{OH}^-$ . Therefore the balanced equation is



$$\begin{aligned} ? \text{ mol Ca}(\text{OH})_2 &= 3 \text{ mol HCl} \\ &\times \frac{1 \text{ mol Ca}(\text{OH})_2}{2 \text{ mol HCl}} \\ &= 1.5 \text{ mol Ca}(\text{OH})_2 \end{aligned}$$

**032 10.0 points**

A 29.1 mL sample of a solution of  $\text{RbOH}$  is neutralized by 22.51 mL of a 2.735 M solution of  $\text{HBr}$ . What is the molarity of the  $\text{RbOH}$  solution?

Correct answer: 2.11563 M.

**Explanation:**

$$\begin{aligned} V_{\text{RbOH}} &= 29.1 \text{ mL} & V_{\text{HBr}} &= 22.51 \text{ mL} \\ [\text{HBr}] &= 2.735 \text{ M} & [\text{RbOH}] &= ? \end{aligned}$$



$$\begin{aligned} \left( \frac{2.735 \text{ mol HBr}}{\text{L}} \right) (22.51 \text{ mL}) \left( \frac{\text{L}}{1000 \text{ mL}} \right) \\ = 0.0615649 \text{ mol HBr} \end{aligned}$$

$$\begin{aligned} (0.0615649 \text{ mol HBr}) \left( \frac{1 \text{ mol RbOH}}{1 \text{ mol HBr}} \right) \\ \times \left( \frac{1}{29.1 \text{ mL}} \right) \left( \frac{1000 \text{ mL}}{\text{L}} \right) \\ = 2.11563 \frac{\text{mol}}{\text{L}} \text{ RbOH} \\ = 2.11563 \text{ M RbOH} \end{aligned}$$

**033 10.0 points**

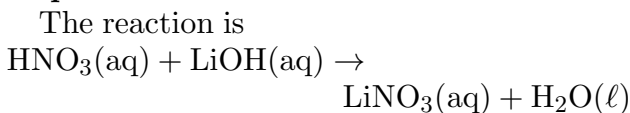
For the neutralization reaction involving  $\text{HNO}_3$  and  $\text{LiOH}$ , how much of 2.10 M  $\text{HNO}_3$  is needed to neutralize 22.2 L of a 4.66 M  $\text{LiOH}$  solution? The molar mass of  $\text{LiOH}$  is 23.95 g/mol. The molar mass of  $\text{HNO}_3$  is 63.1 g/mol. The density of the  $\text{HNO}_3$  solution is 1.06 g/mL. The density of the  $\text{LiOH}$  solution is 1.15 g/mL.

1. 0.567 g
2. 109.7 g
3. 56,600 g
4. 56.6 g
5. 52,200 g **correct**
6. 103.5 g



7. 49.3 g

8. 1,620,000 g

**Explanation:**

Find the number of moles of LiOH used:

$$(22.2 \text{ L LiOH}) \times \frac{4.66 \text{ mol LiOH}}{1 \text{ L LiOH}} = 103.452 \text{ mol LiOH}$$

Find the moles of HNO<sub>3</sub> needed:

$$(103.452 \text{ mol LiOH}) \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol LiOH}} = 103.452 \text{ mol HNO}_3$$

Finally, find the mass of HNO<sub>3</sub>:

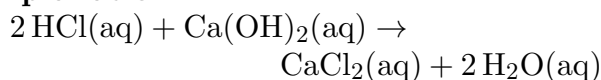
$$(103.452 \text{ mol HNO}_3) \times \frac{1 \text{ L HNO}_3}{2.1 \text{ mol HNO}_3} \times \frac{1.06 \text{ g HNO}_3}{1 \text{ mL HNO}_3} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 52218.6 \text{ g HNO}_3$$

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**034 10.0 points**

An aqueous solution is prepared with 2 moles of HCl and 1 mole of Ca(OH)<sub>2</sub>. The resulting solution contains mainly of

1. water and Cl<sup>-</sup>, H<sup>+</sup>, and Ca<sup>2+</sup> ions.
2. water and Cl<sup>-</sup> and Ca<sup>2+</sup> ions. **correct**
3. water and Cl<sup>-</sup>, H<sup>+</sup>, OH<sup>-</sup>, and Ca<sup>2+</sup> ions.
4. water and Cl<sup>-</sup>, OH<sup>-</sup>, and Ca<sup>2+</sup> ions.

**Explanation:**

1 mole of Ca(OH)<sub>2</sub> reacts with 2 moles of HCl, so there will be no Ca(OH)<sub>2</sub> nor HCl left. The CaCl<sub>2</sub>(aq) will exist as Ca<sup>2+</sup>(aq) and Cl<sup>-</sup>(aq). The H<sup>+</sup> from the HCl and the OH<sup>-</sup> from the Ca(OH)<sub>2</sub> have all reacted. Only a miniscule amount of H<sup>+</sup> and OH<sup>-</sup> remain from the autoionization of the water.

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**035 10.0 points**

Assume you have a 0.4 M solution of acetic

acid that is 1.3 percent ionized or dissociated. What is the pH?

1. 2.3 **correct**

2. 0.3

3. 0.4

4. 1.5

5. 4.3

**Explanation:**

[CH<sub>3</sub>COOH] = 0.4 M      percent = 1.3%  
First calculate the concentration of acetic acid that is ionized:



$$0.4 \text{ M} \times \frac{1.3}{100} = 0.0052 \text{ M H}^+.$$

Thus

$$\text{pH} = -\log [\text{H}^+] = -\log (0.0052) = 2.284.$$

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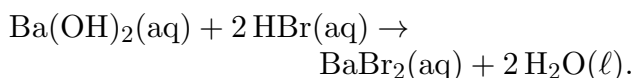
**036 10.0 points**

Determine the total ionic equation for the reaction between HBr(aq) and Ba(OH)<sub>2</sub>(aq).

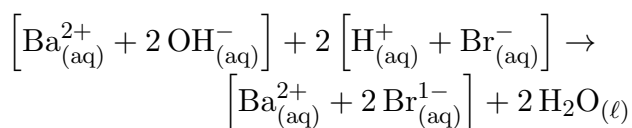
1.  $2 \text{H}^+ + 2 \text{OH}^- \rightarrow 2 \text{H}_2\text{O}$
2.  $2 \text{Br}^- + \text{Ba}^{2+} \rightarrow \text{BaBr}_2$
3.  $2 \text{HBr} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaBr}_2 + 2 \text{H}_2\text{O}$
4.  $2 \text{H}^+ + 2 \text{Br}^- + \text{Ba}^{2+} + 2 \text{OH}^- \rightarrow \text{Ba}^{2+} + 2 \text{Br}^- + 2 \text{H}_2\text{O}$  **correct**

**Explanation:**

An acid and base react to produce a salt and water. In this case the salt, formed from the available cations and anions other than H<sup>+</sup> and OH<sup>-</sup>, is barium bromide (BaBr<sub>2</sub>), which is soluble. The formula unit equation is



In the total ionic equation soluble compounds are written as their ions:

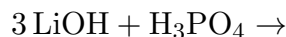
**037 10.0 points**

If aqueous acetic acid is reacted with sodium hydroxide, which of the following substances are in the net ionic equation?

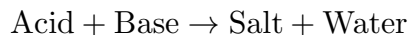
1. acetate ion, hydroxide ion, hydronium ion, and water
2. acetate ion, hydronium ion, and water
3. acetic acid, hydroxide ion, acetate ion, and water **correct**
4. acetic acid, hydroxide ion, hydronium ion, acetate ion, and water
5. acetic acid, sodium ion, hydroxide ion, and acetate ion

**Explanation:****038 10.0 points**

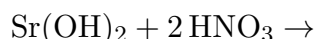
Identify the products of the chemical equation



1.  $3 \text{LiH} + (\text{OH})_3\text{PO}_4$
2.  $\text{Li}_3\text{PO}_4 + 3 \text{H}_2\text{O}$  **correct**
3.  $3 \text{H} + 3 \text{O}_2 + \text{H}_3\text{Li}_3$
4.  $\text{Li}_3\text{P} + 2 \text{H}_2\text{O} + \text{H}_3\text{O}_5$

**Explanation:****039 10.0 points**

What are the products of the following reaction?



1.  $\text{Sr}(\text{NO}_2)_2 + 2 \text{H}_2\text{O}_2$

2.  $\text{Sr}(\text{NO}_3)_2 + 2 \text{H}_2\text{O}$  **correct**

3.  $\text{SrNO}_3 + \text{H}_2\text{O}$

4.  $\text{SrH}_2 + \text{HNO}_5$

**Explanation:**

$\text{Sr}(\text{OH})_2$  is a base and  $\text{HNO}_3$  an acid; they create a salt and water.

**040 10.0 points**

Aqueous ammonia can be used to neutralize sulfuric acid ( $\text{H}_2\text{SO}_4$ ) and nitric acid ( $\text{HNO}_3$ ) to produce two salts extensively used as fertilizers. They are

1.  $(\text{NH}_4)_2\text{SO}_4$  and  $\text{NH}_4\text{NO}_3$ , respectively. **correct**
2.  $\text{NH}_4\text{SO}_4$  and  $\text{NH}_4\text{NO}_3$ , respectively.
3.  $\text{NH}_4\text{SO}_3$  and  $\text{NH}_4\text{OH}$ , respectively.
4. cyanamide and cellulose nitrate, respectively.

**Explanation:**

Aqueous ammonia is a weak base which reacts with acids to form salts. With sulfuric acid and nitric acid, it forms ammonium sulfate and ammonium nitrate, respectively, both of which are used as fertilizers.

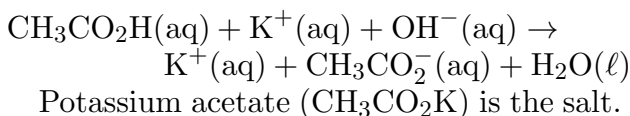
**041 10.0 points**

Identify the salt that is produced from the acid-base neutralization reaction between potassium hydroxide and acetic acid ( $\text{CH}_3\text{COOH}$ ).

1. potassium cyanide
2. potassium acetate **correct**
3. potassium formate
4. potassium amide

**Explanation:**

The balanced equation is

**042 10.0 points**

What volume of 0.585 M  $\text{Ca}(\text{OH})_2$  would be needed to neutralize 15.8 L of 1.51 M HCl?

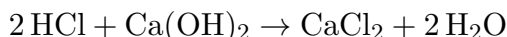
1. 40.8 L
2. 12.2 L
3. 6.12 L
4. 3.06 L
5. 20.4 L **correct**

**Explanation:**

$$[\text{Ca}(\text{OH})_2] = 0.585 \text{ M} \quad V_{\text{HCl}} = 15.8 \text{ L}$$

$$[\text{HCl}] = 1.51 \text{ M}$$

The balanced equation for this neutralization reaction is



We determine the moles of HCl present:

$$? \text{ mol HCl} = 15.8 \text{ L soln} \times \frac{1.51 \text{ mol HCl}}{1 \text{ L soln}}$$

$$= 23.86 \text{ mol HCl}$$

Using the mole ratio from the chemical equation we calculate the moles of  $\text{Ca}(\text{OH})_2$  needed to react with this amount of HCl:

$$? \text{ mol Ca}(\text{OH})_2 = 23.86 \text{ mol HCl}$$

$$\times \frac{1 \text{ mol Ca}(\text{OH})_2}{2 \text{ mol HCl}}$$

$$= 11.93 \text{ mol Ca}(\text{OH})_2$$

We use the molarity of the  $\text{Ca}(\text{OH})_2$  solution to convert from moles to volume of  $\text{Ca}(\text{OH})_2$ :

$$? \text{ L Ca}(\text{OH})_2 = 11.93 \text{ mol Ca}(\text{OH})_2$$

$$\times \frac{1 \text{ L soln}}{0.585 \text{ mol Ca}(\text{OH})_2}$$

$$= 20.4 \text{ L Ca}(\text{OH})_2$$

**043 10.0 points**

It was found that 25 mL of 0.012 M HCl neutralized 40 mL of NaOH solution. What was the molarity of the base solution?

1. 0.006 M
2. 0.012 M
3. 0.050 M
4. 0.0075 M **correct**

**Explanation:**

$$V_{\text{HCl}} = 25 \text{ mL} \quad M_{\text{HCl}} = 0.012 \text{ M}$$

$$V_{\text{NaOH}} = 40 \text{ mL} = 0.04 \text{ L}$$

The base is NaOH. To neutralize,  $\text{mol H}^+$  =  $\text{mol OH}^-$ .

$$n_{\text{H}^+} = \frac{0.012 \text{ mol}}{\text{L}} (25 \text{ mL HCl})$$

$$\times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HCl}}$$

$$= 0.0003 \text{ mol H}^+ = n_{\text{OH}^-} = n_{\text{NaOH}}$$

$$M_{\text{NaOH}} = \frac{\text{mol}}{\text{L}} = \frac{0.0003 \text{ mol NaOH}}{0.04 \text{ L}}$$

$$= 0.0075 \text{ M NaOH}$$

**044 10.0 points**

The pH of a solution of hydrochloric acid is 1.57. What is the molarity of the acid?

Correct answer: 0.0269 mol/L.

**Explanation:****045 10.0 points**

How many moles of NaOH are needed to neutralize three moles of HCl?

1. 0.5
2. one
3. six
4. 1.5

5. three **correct**

6. two

7. eight

8. four

**Explanation:**

For acid base neutralization we need one mole of  $\text{H}^+$  for every mole of  $\text{OH}^-$ . Therefore the balanced equation is

