

This print-out should have 51 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

A complete neutralization reaction is performed by combining 200 mL 0.20 M LiOH and 100 mL 0.40 M nitrous acid (HNO_2). What is the pOH of the resulting solution?

1. 5.74 **correct**
2. 5.50
3. 1.90
4. 11.5
5. 8.26
6. 2.14

Explanation:

The complete neutralization of nitrous acid will end up with lithium nitrite (the salt of the conjugate base) as the dominant species in solution. You can first determine the K_b necessary for this problem:

$$K_w = K_a K_b$$

$$K_b = \frac{1 \times 10^{-14}}{4 \times 10^{-4}}$$

$$K_b = 2.5 \times 10^{-11}$$

Then we can solve for the concentration (C_2) of lithium nitrite in solution, accounting for the dilution involved in making the final solution:

$$C_1 V_1 = C_2 V_2$$

$$C_2 = \frac{(100\text{mL})(0.40\text{M})}{300\text{mL}}$$

$$C_2 = 0.133\text{M}$$

Finally we can use the approximation formula to determine the number of hydroxide ions in solution, which will use to solve for the pOH:

$$[\text{OH}^-] = \sqrt{K_b \times C_b}$$

$$0.00000182574 \text{ M} = \sqrt{K_b \times 0.133}$$

$$\text{pOH} = -\log(0.00000182574)$$

$$\text{pOH} = 5.74$$

002 10.0 points

What is the pH of a 0.07 M $\text{Ba}(\text{OH})_2$ aqueous solution?

1. 13.1461 **correct**
2. 8.7
3. 0.853872
4. 1.31461
5. 9.8596

Explanation:

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

$\text{Ba}(\text{OH})_2$ is a strong base which dissociates in aqueous solution to produce two moles of OH^- for every mole of $\text{Ba}(\text{OH})_2$, so 0.07 M $\text{Ba}(\text{OH})_2$ produces 0.14 M OH^- .

	$\text{Ba}(\text{OH})_2$	\rightarrow	Ba^{2+}	$+$	2OH^-
ini	0.07 M		0 M		0 M
Δ	-0.07 M		+0.07 M		2(0.07 M)
fin	0 M		+0.07 M		+0.14 M

$$\text{pH} = 14 - \text{pOH} = 14 - (-\log 0.14) = 13.1461$$

003 10.0 points

What is $[\text{OH}^-]$ in a 0.0050 M HCl solution?

1. $6.6 \times 10^{-5} \text{ M}$
2. $1.0 \times 10^{-7} \text{ M}$
3. $2.0 \times 10^{-12} \text{ M}$ **correct**
4. 1.0 M
5. $5.0 \times 10^{-3} \text{ M}$

Explanation:

$$[\text{OH}^-] = 0.0050 \text{ M}$$

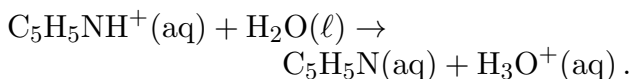
Since HCl is a strong acid, it completely dissociates and H^+ is 0.0050 M.



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

004 10.0 points

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



- 1.8×10^{-16}
- 5.6×10^8
- -1.8×10^{-9}
- 5.6×10^{-6} **correct**
- 1.8×10^{-9}

Explanation:

005 10.0 points

For the neutralization reaction involving HCl and $Mg(OH)_2$, how many liters of 0.60 M HCl are needed to react with 45 g of a 2.5 M $Mg(OH)_2$ solution? (density = 1.3 g/mL)

- 0.011 L
- 0.12 L
- 0.29 L **correct**
- 120 L
- 0.14 L
- 0.49 L

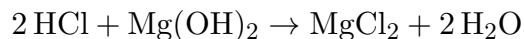
Explanation:

$$[HCl] = 0.60 \text{ M}$$

$$m_{HCl} = 45 \text{ g}$$

$$[Mg(OH)_2] = 2.5 \text{ M} \quad \text{density} = 1.3 \text{ g/mL}$$

The balanced equation for this neutralization reaction is:



We use the density to determine the volume of the $Mg(OH)_2$ solution:

$$\begin{aligned} ? \text{ L } Mg(OH)_2 \text{ soln} &= 45 \text{ g soln} \\ &\times \frac{0.001 \text{ L soln}}{1.3 \text{ g soln}} \\ &= 0.0346 \text{ L soln} \end{aligned}$$

We use the molarity to convert from L of solution to moles $Mg(OH)_2$:

$$\begin{aligned} ? \text{ mol } Mg(OH)_2 &= 0.0346 \text{ L soln} \\ &\times \frac{2.5 \text{ mol } Mg(OH)_2}{1 \text{ L soln}} \\ &= 0.0865 \text{ mol } Mg(OH)_2 \end{aligned}$$

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

$$\begin{aligned} ? \text{ mol HCl} &= 0.0865 \text{ mol } Mg(OH)_2 \\ &\times \frac{2 \text{ mol HCl}}{1 \text{ mol } Mg(OH)_2} \\ &= 0.173 \text{ mol HCl} \end{aligned}$$

We use the molarity to find the volume of HCl solution needed to provide 0.173 mol HCl:

$$\begin{aligned} ? \text{ L HCl} &= 0.173 \text{ mol HCl} \times \frac{1 \text{ L soln}}{0.60 \text{ mol HCl}} \\ &= 0.29 \text{ L HCl} \end{aligned}$$

006 10.0 points

The conjugate pair of F^- is a...

- weak acid **correct**
- strong acid
- strong base
- weak base

5. neutral salt

Explanation:

HF is a weak acid. HClO_3 is a strong acid. HSO_4^- is a weak acid.

007 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 = 9.48
3. pH = 7.12
4. pH = 10.37
5. pH = 9.26 **correct**
6. pH = 3.63
7. pH = 4.74

Explanation:

Hydroxylamine is a weak base, so use the equation to calculate weak base $[\text{OH}^-]$ 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 $[\text{OH}^-]$, 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)

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

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

008 10.0 points

What is $[\text{H}_3\text{O}^+]$ when $[\text{OH}^-] = 3.3 \times 10^{-9}$ M?

1. 6.6×10^{-5} M
2. 3.3×10^{-9} M
3. 3.0×10^{-6} M **correct**
4. 3.3×10^{-5} M
5. 1.0×10^{-7} M

Explanation:

$$[\text{OH}^-] = 3.3 \times 10^{-9} \text{ M}$$

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

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

009 10.0 points

Calculate the resulting pH if 365 mL of 2.88 M HNO_3 is mixed with 335 mL of 1.10 M $\text{Ca}(\text{OH})_2$ solution.

1. 1.46
2. 0.460
3. 2.36
4. 7.20
5. 0.067
6. 0.350 **correct**

Explanation:

$$\begin{aligned} V_{\text{HNO}_3} &= 365 \text{ mL} & [\text{HNO}_3] &= 2.88 \text{ M} \\ V_{\text{Ca}(\text{OH})_2} &= 335 \text{ mL} & [\text{Ca}(\text{OH})_2] &= 1.10 \text{ M} \end{aligned}$$

To determine the pH of the final mixture, we need to determine how much H or OH is left over after the reaction. Remember that for complete neutralization we need H and OH in equal molar amounts:



First calculate how many moles of H^+ and

OH^- we have:

$$\begin{aligned} ? \text{ mol H}^+ &= 0.365 \text{ L} \times \frac{2.88 \text{ mol HNO}_3}{1 \text{ L}} \\ &\times \frac{1 \text{ mol H}^+}{1 \text{ mol HNO}_3} \\ &= 1.05 \text{ mol H}^+ \end{aligned}$$

$$\begin{aligned} ? \text{ mol OH}^- &= 0.335 \text{ L} \times \frac{1.10 \text{ mol Ca(OH)}_2}{1 \text{ L}} \\ &\times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2} \\ &= 0.737 \text{ mol OH}^- \end{aligned}$$

Now we can see that we have more H^+ than OH^- so OH^- will be consumed and H^+ will be left over. Let's find out by how much:

$$\begin{aligned} ? \text{ mol H}^+ &= 1.05 \text{ mol} - 0.737 \text{ mol} \\ &= 0.313 \text{ mol H}^+ \end{aligned}$$

The next step is to calculate the $[\text{H}^+]$:

$$\begin{aligned} ? \text{ M H}^+ &= \frac{0.313 \text{ mol H}^+}{0.700 \text{ L}} \\ &= 0.447 \text{ M H}^+ \\ \text{pH} &= -\log[0.447] = 0.350 \end{aligned}$$

010 10.0 points

181 mL of an unknown HCl solution was neutralized in a titration with 36.2 mL of 0.250 M NaOH. What is the molarity of the unknown HCl solution?

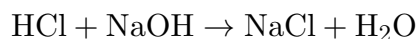
1. 5.00×10^{-2} M **correct**
2. 8.00×10^{-1} M
3. 9.05×10^{-2} M
4. 1.64×10^3 M
5. 2.50×10^{-1} M

Explanation:

$$V_{\text{HCl}} = 181 \text{ mL} \qquad V_{\text{NaOH}} = 36.2 \text{ mL}$$

$$[\text{NaOH}] = 0.250 \text{ M}$$

The balanced equation for this neutralization reaction is



We determine the moles of NaOH used:

$$\begin{aligned} ? \text{ mol NaOH} &= 0.0362 \text{ L soln} \\ &\times \frac{0.250 \text{ mol NaOH}}{1 \text{ L soln}} \\ &= 0.00905 \text{ mol NaOH} \end{aligned}$$

From the 1:1 mole ratio in the balanced chemical reaction we know we would need 0.00905 moles of NaOH to neutralize 0.00905 moles HCl. This is the amount of HCl that must have been in the 181 mL sample. Molarity is moles solute per liter of solution:

$$\begin{aligned} ? \text{ M HCl} &= \frac{0.00905 \text{ mol NaOH}}{0.181 \text{ L solution}} \\ &= 0.05 \text{ M HCl} \end{aligned}$$

011 10.0 points

Which of the following is the WEAKEST acid?

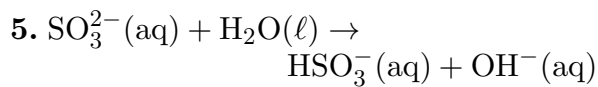
1. HClO_2
2. HClO **correct**
3. HNO_3
4. HClO_4

Explanation:

012 10.0 points

Which equation represents K_{a2} for sulfurous acid?

1. $\text{HSO}_3^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{SO}_3(\text{aq}) + \text{OH}^-(\text{aq})$
2. $\text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{HSO}_3^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
3. $\text{HSO}_3^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{SO}_3^{2-}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ **correct**
4. $\text{H}_2\text{SO}_3(\text{aq}) + 2 \text{H}_2\text{O}(\ell) \rightarrow \text{SO}_3^{2-}(\text{aq}) + 2 \text{H}_3\text{O}^+(\text{aq})$

**Explanation:****013 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. 5.30
2. 1.0×10^{-3}
3. 1.8×10^{-5}
4. 5.0×10^{-6} **correct**
5. 5.0×10^{-3}
6. 2.0×10^{-5}
7. 1.0×10^{-6}
8. 2.0×10^{-9}

Explanation:**014 10.0 points**

At 25° C, water solutions which are neutral have a pH of

1. about 7. **correct**
2. about 0.
3. about 14.
4. infinity.

Explanation:**015 (part 1 of 2) 10.0 points**

Calculate the pH of the solute in an aqueous solution of 0.45 M $\text{C}_5\text{H}_5\text{N}(\text{aq})$ (pyridine) if the K_b is 1.8×10^{-9} .

1. 9.51306
2. 9.49978

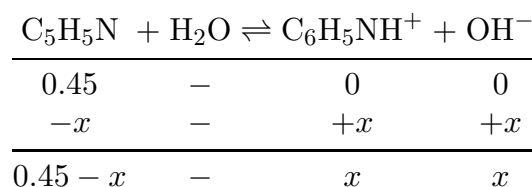
3. 9.05019
4. 9.52207
5. 9.07918
6. 9.27815
7. 9.49181
8. 9.47273
9. 9.53576
10. 9.45424

Your answer must be within ± 0.005 .

Correct answer: 9.45424.

Explanation:

$$C_{\text{pyridine}}^{\circ} = 0.45 \text{ M} \quad K_b = 1.8 \times 10^{-9}$$



$$K_b = \frac{[\text{C}_6\text{H}_5\text{NH}^+][\text{OH}^-]}{[\text{C}_5\text{H}_5\text{N}]}$$

$$1.8 \times 10^{-9} = \frac{x^2}{0.45 - x} \approx \frac{x^2}{0.45}$$

$$x = [\text{OH}^-] = \sqrt{0.45(1.8 \times 10^{-9})} = 2.84605 \times 10^{-5} \text{ mol/L.}$$

The pOH is

$$\text{pOH} = -\log(2.84605 \times 10^{-5}) = 4.54576,$$

and the pH is

$$\text{pH} = 14 - 4.54576 = 9.45424.$$

016 (part 2 of 2) 10.0 points

What is the percentage protonation of the solute?

1. 0.0080904
2. 0.00518321
3. 0.00632456
4. 0.00733017
5. 0.00781133
6. 0.00524222
7. 0.0107763
8. 0.00884652

9. 0.00914991
10. 0.00569495

Correct answer: 0.00632456%.

Explanation:

$$\begin{aligned} \% \text{ protonation} &= \frac{[\text{protonated species}]}{C_{\text{pyridine}}^{\circ}} \\ &= \frac{2.84605 \times 10^{-5}}{0.45} \times 100\% \\ &= 0.00632456\%. \end{aligned}$$

017 10.0 points

Calculate the concentration of HCO_3^- present in 0.0322 M H_2CO_3 .

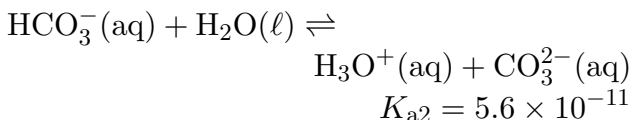
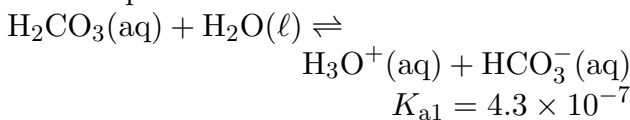
- 0.000236977
- 0.000274474
- 0.000319772
- 0.000117669
- 0.000225732
- 0.000193972
- 0.00033068
- 0.000304127
- 0.000152804
- 0.000389878

Correct answer: 0.000117669 mol/L.

Explanation:

$$K_w = 1 \times 10^{-14}$$

The equilibrium reactions of interest are



Because the second ionization constant is much smaller than the first, we can assume that the first step dominates.

H_2CO_3	+	H_2O	\rightleftharpoons	H_3O^+	+	HCO_3^-
0.0322	–			0		0
$-x$	–			$+x$		$+x$
$0.0322 - x$	–			x		x

Assuming that $x \ll 0.0322$ M,

$$\begin{aligned} K_{a1} &= \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \\ 4.3 \times 10^{-7} &= \frac{x^2}{0.0322 - x} \approx \frac{x^2}{0.0322} \\ x &= \sqrt{(0.0322)(4.3 \times 10^{-7})} \\ &= 0.000117669 \text{ M}. \end{aligned}$$

Because $x < 1\%$ of 0.0322 M, the assumption was valid, so

$$x = [\text{H}_3\text{O}^+] = [\text{HCO}_3^-] = 0.000117669 \text{ M}.$$

018 10.0 points

What is the pH of a 0.036 M HNO_3 solution?

- 2.72
- 5.36
- 12.56
- 2.88
- 1.64
- 1.44 correct
- 1.36
- 5.56

Explanation:

For a strong acid:

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

$$1.44 = -\log(0.036)$$

019 10.0 points

What is the pH of a 0.37 M solution of anilinium nitrate ($\text{C}_6\text{H}_5\text{NH}_3\text{NO}_3$)? K_b for aniline is 4.2×10^{-10} .

- 2.53959
- 2.62152
- 2.75465
- 2.70956

5. 2.73856
6. 2.55905
7. 2.60414
8. 2.67225
9. 2.52752
10. 2.58043

Your answer must be within $\pm 0.4\%$
 Correct answer: 2.52752.

Explanation:

$$M_{\text{C}_6\text{H}_5\text{NH}_3\text{NO}_3} = 0.37 \text{ M} \quad K_b = 4.2 \times 10^{-10}$$

It's a salt of a weak base (BHX). This means you need a K_a for the weak acid 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

$$\begin{aligned} [\text{H}^+] &= \sqrt{K_a \cdot C_{\text{BH}^+}} \\ &= \sqrt{(2.38095 \times 10^{-5})(0.37)} \\ &= 0.00296808 \text{ M} \end{aligned}$$

$$\text{pH} = -\log(0.00296808) = 2.52752$$

020 10.0 points

The ΔH° for the autoionization of water is 55.7 kJ/mol. Which of the following is/are true regarding the autoionization of pure water at 37°C ?

- I) The concentration of H_3O^+ will be greater at 37°C than at standard conditions
- II) $[\text{OH}^-] = [\text{H}_3\text{O}^+]$
- III) K_w is equal to 1×10^{-14}
- IV) The concentration of OH^- will be lower at 37°C than standard conditions

1. I, II, and III

2. I and II **correct**

3. II and IV

4. III only

5. II only

6. I, II, and IV

Explanation:

The autoionization of water is an endothermic process. Therefore, the reaction will shift to the products when the temperature increases. In this process, K_w will increase. At 37°C , the concentration of H_3O^+ and $[\text{OH}^-]$ will be greater. However, at any given temperature $[\text{OH}^-]$ must equal $[\text{H}_3\text{O}^+]$ for pure water.

021 10.0 points

Which of the following can exhibit amphoterism?

1. CrO_4^{2-}

2. HClO_4

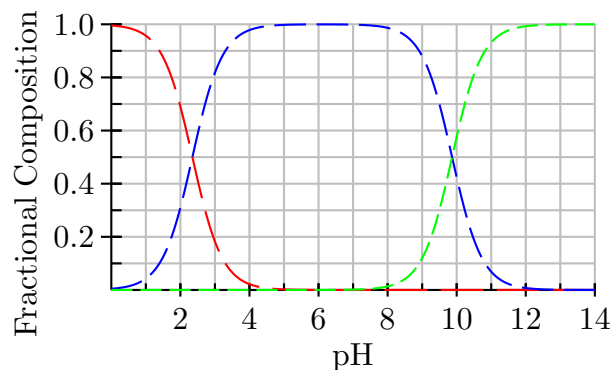
3. HPO_4^{2-} **correct**

4. Sn^{2+}

5. O^{2-}

Explanation:**022 10.0 points**

Consider the fractional composition diagram for the amino acid alanine.



What is the structure of the dominant species at pH 6?

- ${}^{-}\text{OOC}-\text{CH}(\text{CH}_3)\text{NH}_3^+$ **correct**
- ${}^{-}\text{OOC}-\text{CH}(\text{CH}_3)\text{NH}_2$
- $\text{HOOC}-\text{CH}(\text{CH}_3)\text{NH}_3^+$

Explanation:

To the left of 2.348, the red graph representing $\text{HOOC}-\text{CH}(\text{CH}_3)\text{NH}_3^+$ is dominant.

Between 2.348 and 9.867, the blue graph representing ${}^{-}\text{OOC}-\text{CH}(\text{CH}_3)\text{NH}_3^+$ is dominant.

To the right of 9.867, the green graph representing ${}^{-}\text{OOC}-\text{CH}(\text{CH}_3)\text{NH}_2$ is dominant.

023 10.0 points

Trichloroacetic acid is a harsh chemical, typically used for cosmetic treatments such as tattoo removal. What is the pH of a .0800 M trichloroacetic acid solution (CCl_3COOH)?

- 2.000
- 1.097
- 0.990
- 0.824
- 1.205 **correct**
- 2.019

Explanation:

Due to the large size of K_a , you will need to solve the full quadratic equation to get the correct answer. For .0800 M trichloroacetic acid solution, the quadratic will net $[\text{H}^+] = 0.0623369$ M. Therefore, your pH will = 1.205.

024 10.0 points

What is the hydroxide concentration $[\text{OH}^-]$ in an aqueous solution in which the hydronium ion concentration $[\text{H}_3\text{O}^+]$ is 1×10^{-5} M?

- 1×10^{-9} M **correct**
- 1×10^{-14} M
- 1×10^{-5} M

- 1×10^{-7} M

Explanation:

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-5} \text{ M}$$

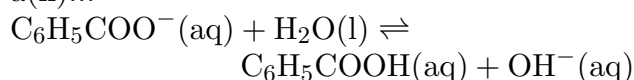
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]}$$

$$= \frac{1.0 \times 10^{-14}}{1 \times 10^{-5}} = 1 \times 10^{-9} \text{ M}$$

025 10.0 points

In the following equation, water is acting as a(n)...



- Acid **correct**
- Base
- Neither

Explanation:

$\text{C}_6\text{H}_5\text{COO}^-$ is a base and accepts a proton from H_2O . Therefore H_2O acts as an acid.

026 10.0 points

The generic weak acid HA has a percent ionization equal to 10.8% at a 0.025 M concentration. What is the K_a of this acid?

- 1.43×10^{-3}
- 4.54×10^{-4}
- 2.50×10^{-4}
- 5.97×10^{-4}
- 3.27×10^{-4} **correct**

Explanation:

For a generic weak acid, the K_a expression is:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Starting with 0.025 M HA, your values are:

$$[\text{HA}] = 0.025 \text{ M} - (0.108 \cdot 0.025 \text{ M}) = 0.022300 \text{ M}$$

$$[\text{H}^+] = [\text{A}^-] = 0.00270 \text{ M}$$

$$K_a = \frac{0.00270^2}{0.022300}$$

$$K_a = 0.000327$$

027 10.0 points

A student titrated a sample containing 107 mL of 0.28 M NaOH with 0.72 M HCl. What is the pH of the solution after 23 mL of the hydrochloric acid has been added?

1. 13.02 correct

2. 0.98

3. -0.02

4. 7

5. 14.02

Explanation:

$$V_{\text{NaOH}} = 107 \text{ mL} \quad [\text{NaOH}] = 0.28 \text{ M}$$

$$V_{\text{HCl}} = 23 \text{ mL} \quad [\text{HCl}] = 0.72 \text{ M}$$

Initially

$$n_{\text{NaOH}} : (107 \text{ mL})(0.28 \text{ M}) = 29.96 \text{ mmol}$$

$$n_{\text{HCl}} : (23 \text{ mL})(0.72 \text{ M}) = 16.56 \text{ mmol}$$



$$\text{ini} \quad 29.96 \text{ mmol} \quad 16.56 \text{ mmol}$$

$$\Delta \quad -16.56 \text{ mmol} \quad -16.56 \text{ mmol}$$

$$\text{fin} \quad 13.4 \text{ mmol} \quad 0 \text{ mmol}$$

$$V_{\text{new}} = 107 \text{ mL} + 23 \text{ mL} = 130 \text{ mL}$$

$$[\text{NaOH}] = \frac{13.4 \text{ mmol}}{130 \text{ mL}} = 0.103 \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.103) = 0.98$$

$$\text{pH} = 14 - \text{pOH} = 14 - 0.98 = 13.02$$

028 10.0 points

What is the concentration of H_3O^+ ions in a 0.20 M solution of $\text{Ca}(\text{OH})_2$ at 25°C ?

1. 0.40 M

2. 2.5×10^{-14} M correct

3. 1×10^{-13} M

4. 1.2×10^{-14} M

5. 0.20 M

6. 3.7×10^{-14} M

Explanation:

0.20 M $\text{Ca}(\text{OH})_2$ completely dissociates in aqueous solution to give a total hydroxide concentration equal to 0.40 M. You can use the following expression to solve for $[\text{OH}^-]$:

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]}$$

$$2.5 \times 10^{-14} \text{ M} = \frac{1 \times 10^{-14}}{0.40 \text{ M}}$$

029 10.0 points

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

1. 0.55

2. 5.25

3. 2.01 correct

4. 3.17

5. 1.46

Explanation:

$$M = 0.28 \text{ M}$$

$$P = 3.5\%$$

3.5% of the 0.28 M is ionized (contributes to pH), so

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

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

030 10.0 points

The hydronium ion concentration in a solution at pH 10 has what relationship to the hydronium ion concentration in a solution at pH 13?

- 3 times greater than
- 100 times less than
- 1,000 times less than
- 1,000 times greater than **correct**

Explanation:

For pH 10, $M_{\text{H}^+} = 1 \times 10^{-10}$

For pH 13, $M_{\text{H}^+} = 1 \times 10^{-13}$

031 10.0 points

List the the following solution species in order of increasing acidity (weakest acid to strongest acid). Important note: measure the acidity/basicity of the solution species as written.



- $(\text{CH}_3)_3\text{NH}^+ < \text{HIO} < \text{NH}_4^+ < \text{HCN}$
- $\text{HIO} < (\text{CH}_3)_3\text{NH}^+ < \text{NH}_4^+ < \text{HCN}$ **correct**
- $\text{HCN} < \text{NH}_4^+ < (\text{CH}_3)_3\text{NH}^+ < \text{HIO}$
- $\text{HIO} < \text{NH}_4^+ < \text{HCN} < (\text{CH}_3)_3\text{NH}^+$
- $\text{HCN} < (\text{CH}_3)_3\text{NH}^+ < \text{HIO} < \text{NH}_4^+$
- $(\text{CH}_3)_3\text{NH}^+ < \text{NH}_4^+ < \text{HCN} < \text{HIO}$

Explanation:

Look at the coversheet for the values of pK_a and pK_b . The following information is provided:

$\text{HCN}, pK_a = 9.21$

$\text{HIO}, pK_a = 10.7$

$(\text{CH}_3)_3\text{N}, pK_b = 4.19$

$\text{NH}_3, pK_b = 4.74$

The question either asks for increasing base strength or increasing acid strength. For acids, you are comparing the pK_a values:

$\text{HCN}, pK_a = 9.21$

$\text{HIO}, pK_a = 10.7$

$(\text{CH}_3)_3\text{NH}^+, pK_a = 14 - 4.19 = 9.81$

$\text{NH}_4^+, pK_a = 14 - 4.74 = 9.3$

This gives the rank: $\text{HIO} < (\text{CH}_3)_3\text{NH}^+ < \text{NH}_4^+ < \text{HCN}$

For bases, you are comparing the pK_b values:

$\text{CN}^-, pK_b = 14 - 9.21 = 4.79$

$\text{IO}^-, pK_b = 14 - 10.7 = 3.3$

$(\text{CH}_3)_3\text{N}, pK_b = 4.19$

$\text{NH}_3, pK_b = 4.74$

This gives the opposite rank: $\text{CN}^- < \text{NH}_3 < (\text{CH}_3)_3\text{N} < \text{IO}^-$

032 10.0 points

According to the Bronsted-Lowry Theory of acids and bases, a base is

- a substance which when dissolved in water yields OH^- .
- an electron acceptor.
- a proton donor.
- a proton acceptor. **correct**
- an electron donor.

Explanation:

Bases are defined by the Arrhenius theory as substances which in water produce hydroxide ions, by the Bronsted-Lowry theory as substances which accept protons, and by the Lewis theory as substances which provide electron pairs.

033 10.0 points

A 50.0 mL sample of 6 M HCl is added to 100.0 mL of 1.2 M NaOH. What is the pH of the resulting solution?

- 0.09
- 1.20
- 0.08 **correct**

4. 0.18

5. -1.20

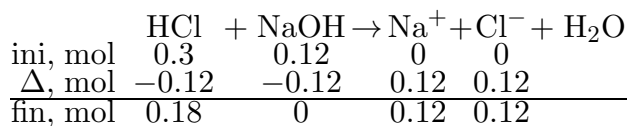
Explanation:

$$V_{\text{HCl}} = 50.0 \text{ mL} \quad [\text{HCl}] = 6 \text{ M}$$

$$V_{\text{NaOH}} = 100.0 \text{ mL} \quad [\text{NaOH}] = 1.2 \text{ M}$$

$$n_{\text{HCl}} = (0.05 \text{ L})(6 \text{ M}) = 0.3 \text{ mol}$$

$$n_{\text{NaOH}} = (0.1 \text{ L})(1.2 \text{ M}) = 0.12 \text{ mol}$$



HCl is a strong acid, Na⁺ and Cl⁻ are spectator ions.

$$[\text{H}_3\text{O}^+] = \frac{0.18 \text{ mol}}{0.150 \text{ L}} = 1.2 \text{ M}$$

$$\text{pH} = -\log(1.2 \text{ M}) = -0.0791812$$

034 10.0 points

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

1. $\text{NH}_3 + \text{H}_3\text{O}^+ \rightleftharpoons \text{NH}_4^+ + \text{H}_2\text{O}$
2. $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$ **correct**
3. The term is misleading, because the ammonium ion is not an acid.
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. $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

Explanation:**035 10.0 points**

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

What is the percentage aniline protonated?

1. 0.69%
2. 2.1%

3. 0.21%

4. 0.12%

5. 0.021% **correct****Explanation:****036 10.0 points**

What is the conjugate acid of NO_3^- ?

1. NO_2^-
2. H^+
3. NO_3^{2-}
4. NH_3
5. OH^-
6. HNO_3 **correct**

Explanation:

Since the question asks for the conjugate acid, we can assume NO_3^- is acting as a base. This means that it is a proton acceptor. To form the conjugate acid, it accepts a H making HNO_3 .

037 10.0 points

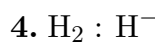
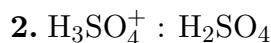
Write the charge balance equation for a dilute aqueous solution of HI.

1. $[\text{H}_3\text{O}^+] = [\text{I}^-]$
2. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$
3. $[\text{H}_3\text{O}^+] = [\text{I}^-] + [\text{OH}^-]$ **correct**
4. $[\text{I}^-] = [\text{OH}^-] + [\text{H}_3\text{O}^+]$
5. $[\text{HI}]_{\text{initial}} = [\text{I}^-]$

Explanation:**038 10.0 points**

Which is NOT a conjugate acid-base pair?

1. $\text{HCl} : \text{Cl}^-$

**Explanation:**

Except for H_2SO_4 and SO_4^{2-} , the members of all of the pairs differ by one proton.

039 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	1.0×10^{-3}
2	3.0×10^{-5}
3	2.6×10^{-7}
4	4.0×10^{-9}
5	7.3×10^{-11}

The anion of which acid is the strongest base?

1. 1

2. 5 correct

3. 3

4. 4

5. 2

Explanation:

040 10.0 points

What is the pH of a 0.24 M solution of potassium generate (KR-COO)? K_a for the generic acid (R-COOH) is 2.7×10^{-8} .

1. 3.526

2. 10.604

3. 7.000

4. 6.431

5. 10.234

6. 10.474 correct

7. 10.844

8. 3.396

9. 7.569

10. 10.324

Explanation:

$$M_{\text{KR-COO}} = 0.24 \text{ M} \quad K_a = 2.7 \times 10^{-8}$$

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

$$[\text{OH}^-] = \sqrt{K_b \cdot C_{\text{A}^-}} = 0.000298142 \text{ M}$$

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

$$= 14 + \log(0.000298142) = 10.4744$$

041 10.0 points

For a solution labeled "0.10 M $\text{H}_3\text{PO}_4(\text{aq})$,"

1. $[\text{H}_2\text{PO}_4^-]$ is greater than 0.10 M.2. $[\text{PO}_4^{3-}] = 0.10 \text{ M}$.3. $[\text{H}^+] = 0.10 \text{ M}$.4. $[\text{H}^+]$ is less than 0.10 M. correct5. $[\text{H}^+] = 0.30 \text{ M}$.**Explanation:**

042 10.0 points

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

1. $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
2. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ **correct**
3. $\text{pH} = 7.0$ or greater than 7.0
4. K_w decreases with increasing temperature.
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°C . $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$, which increases with temperature.

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

043 10.0 points

Determine the pH of a 0.000496 M solution of NaOH .

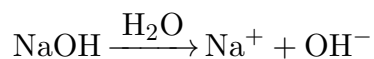
1. 10.7459
2. 10.4928
3. 10.2833
4. 10.6955
5. 10.8506
6. 10.9138
7. 10.8156
8. 10.9523
9. 10.2279
10. 10.7924

Correct answer: 10.6955.

Explanation:

$[\text{NaOH}] = 0.000496 \text{ M}$

$\text{pH} = ?$



$$[\text{OH}^-] = 4.96 \times 10^{-4} \text{ M}$$

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

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

$$= \frac{1.0 \times 10^{-14} \text{ M}^2}{4.96 \times 10^{-4} \text{ M}}$$

$$= 2.01613 \times 10^{-11} \text{ M}$$

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

$$= -\log(2.01613 \times 10^{-11})$$

$$= 10.6955$$

044 10.0 points

The pH of $0.010 \text{ M H}_3\text{PO}_4(\text{aq})$ is 2.24. Estimate the concentration of PO_4^{3-} in the solution. For H_3PO_4 , the values of K_{a1} , K_{a2} , and K_{a3} are 7.6×10^{-3} , 6.2×10^{-8} , and 2.1×10^{-13} , respectively.

1. $2.1 \times 10^{-13} \text{ M}$
2. $5.8 \times 10^{-3} \text{ M}$
3. $6.2 \times 10^{-8} \text{ M}$
4. $2.3 \times 10^{-18} \text{ M}$ **correct**
5. $7.6 \times 10^{-3} \text{ M}$

Explanation:

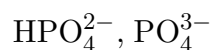
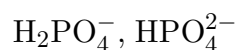
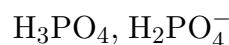
045 10.0 points

What is the conjugate base of HPO_4^{2-} ?

1. OH^-
2. HPO_4^{2-}
3. H_3PO_4
4. H_2PO_4^-
5. PO_4^{3-} **correct**

Explanation:

The conjugate base pairs are shown below:



046 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. 2.0%
2. 0.10% **correct**
3. 0.050%
4. 0.020%
5. 0.00020%

Explanation:

047 10.0 points

In the complete neutralization of butanoic acid (a weak acid) with an equal amount of sodium hydroxide (NaOH), the dominant species in the resulting solution is a...

1. weak base **correct**
2. weak acid
3. strong base
4. neutral salt
5. strong acid

Explanation:

The product of the neutralization of a weak acid is the salt of its conjugate weak base. The product of the neutralization of a weak base is the salt of its conjugate acid.

048 10.0 points

What is the pH of a 0.480 M trimethylamine ($(\text{CH}_3)_3\text{N}$) solution?

1. 8.81
2. 11.48
3. 11.87
4. 8.86
5. 2.13
6. 8.96

7. 2.25

8. 11.74 **correct**

Explanation:

For trimethylamine, K_b is small enough to use the weak base approximation formula:

$$[\text{OH}^-] = \sqrt{K_b \times C_b}$$

$$0.00558570\dots\text{M} = \sqrt{(6.5 \times 10^{-5})(0.480\text{M})}$$

$$\text{pOH} = -\log(0.00558570\dots)$$

$$\text{pH} = 14 - 2.25292\dots$$

$$\text{pH} = 11.747$$

Note: this answer ignores x, so the closest answer should be a reasonable assumption.

049 10.0 points

Which of the following aqueous solutions gives a pH greater than 7?

1. 10^{-8} M CH_3COOH
2. 10^{-8} M HCl
3. None of the solutions gives a pH greater than 7. **correct**
4. 10^{-8} M NH_4Cl
5. 10^{-8} M HCOOH

Explanation:

050 10.0 points

What is $[\text{OH}^-]$ when $[\text{H}_3\text{O}^+] = 8.2 \times 10^{-5}$ M?

1. 1.2×10^{-10} M **correct**
2. 6.3×10^{-9} M
3. 1.2×10^{-4} M
4. 1.0×10^{-7} M
5. 8.2×10^{-5} M

Explanation:

$$[\text{H}_3\text{O}^+] = 8.2 \times 10^{-5} \text{ M}$$

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

$$\begin{aligned} [\text{OH}^-] &= \frac{K_w}{[\text{H}_3\text{O}^+]} \\ &= \frac{1.0 \times 10^{-14}}{8.2 \times 10^{-5}} = 1.2 \times 10^{-10} \text{ M} \end{aligned}$$

051 10.0 points

Which pH represents a solution with 1000 times higher $[\text{OH}^-]$ than a solution with pH of 5?

1. pH = 3
2. pH = 6
3. pH = 2
4. pH = 8 **correct**
5. pH = 7
6. pH = 0.005
7. pH = 4
8. pH = 5000
9. pH = 1

Explanation:

$$\text{pH} = 5$$

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

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

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

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

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