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

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	1.0×10^{-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. 4

2. 5 **correct**

3. 2

4. 3

5. 1

Explanation:

002 10.0 points

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

1. $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ **correct**

2. $NH_3 + H_2O \rightleftharpoons NH_4{}^+ + OH^-$

3. $NH_3 + H_3O^+ \rightleftharpoons NH_4^+ + H_2O$

4. $NH_4^+ + OH^- \rightleftharpoons NH_3 + H_2O$

5. $NH_4Cl(solid) + H_2O \rightleftharpoons NH_4^+ + Cl^-$

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

Explanation:

003 10.0 points

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

$$C_5H_5NH^+(aq) + H_2O(\ell) \rightarrow C_5H_5N(aq) + H_3O^+(aq)\,.$$

1. 5.6×10^{-6} **correct**

2. 1.8×10^{-9}

3. 1.8×10^{-16}

4. 5.6×10^{8}

5. -1.8×10^{-9}

Explanation:

004 10.0 points

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

1. K_w decreases with increasing temperature.

2. $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

3. $[H_3O^+] = [OH^-]$ **correct**

4. pH = 7.0 or greater than 7.0

5. pH = 7.0

Explanation:

K_w is shown to INCREASE with increasing temperature. pH = 7 is only true when water is at 24°C. $[H_3O^+][OH^-] = K_w$, which increases with temperature.

At high temperatures pH can be less than 7. Thus $[H_3O^+] = [OH^-]$ is the only case that is true.

005 10.0 points

Which is NOT a conjugate acid-base pair?

1. $H_2O : OH^-$

2. $HCl : Cl^-$

3. $H_3SO_4^+ : H_2SO_4$

4. $H_2 : H^-$

5. $H_2SO_4 : SO_4^{2-}$ **correct**

Explanation:
Except for H_2SO_4 and SO_4^{2-}, 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_3^-?

1. ${NO_2}^-$

2. NH_3

3. H^+

4. HNO_3 **correct**

5. ${NO_3}^{2-}$

6. OH^-

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.

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007 10.0 points
What is $[H_3O^+]$ when $[OH^-] = 3.3 \times 10^{-9}$ M?

1. 1.0×10^{-7} M

2. 3.3×10^{-9} M

3. 3.3×10^{-5} M

4. 3.0×10^{-6} M **correct**

5. 6.6×10^{-5} M

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

$$K_w = [H_3O^+][OH^-] = 1 \times 10^{14}$$

$$\begin{aligned}[H_3O^+] &= \frac{K_w}{[OH^-]} \\ &= \frac{1.0 \times 10^{14}}{3.3 \times 10^{-9}} = 3.0 \times 10^{-6} \text{ M}\end{aligned}$$

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008 10.0 points
What is $[OH^-]$ in a 0.0050 M HCl solution?

1. 6.6×10^{-5} M

2. 5.0×10^{-3} M

3. 1.0×10^{-7} M

4. 2.0×10^{-12} M **correct**

5. 1.0 M

Explanation:
$[OH^-] = 0.0050$ M
Since HCl is a strong acid, it completely dissociates and H^+ is 0.0050 M.

$$HCl \rightleftharpoons H^+ + Cl^-$$

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

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009 10.0 points
Which pH represents a solution with 1000 times higher $[OH^-]$ 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

$$\text{pOH} = 14 - \text{pH} = 14 - 5 = 9$$
$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-9}\text{ M}$$

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

$$\text{pOH}_x = -\log(\text{OH}_x) = 6$$
$$\text{pH}_x = 14 - \text{pOH}_x = 14 - 6 = 8$$

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010 10.0 points

What is the pH of a 0.12 M $Ba(OH)_2$ aqueous solution?

1. 1.33802

2. 8.7

3. 0.619789

4. 13.3802 **correct**

5. 10.0352

Explanation:
$[Ba(OH)_2] = 0.15$ M

$Ba(OH)_2$ is a strong base which dissociates in aqueous solution to produce two moles of OH^- for every mole of $Ba(OH)_2$, so 0.12 M $Ba(OH)_2$ produces 0.24 M OH^-.

	$Ba(OH)_2$ →	Ba^{2+} +	$2\,OH^-$
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 $[OH^-]$ concentration (note that this is the approximate equation. Why? Because K_b is very small and the concentration is reasonable) :

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

After finding $[OH^-]$, you can find pH using either method below:

A)
$$\text{pOH} = -\log\left(1.82 \times 10^{-5}\right) = 4.74$$
$$\text{pH} = 14 - 4.74 = 9.26$$

or B)
$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.82 \times 10^{-5}} = 5.52 \times 10^{-10}$$
$$\text{pH} = -\log\left(5.52 \times 10^{-10}\right) = 9.26$$

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

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 $K_b = \dfrac{K_w}{K_a}$ 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

1. less than 7.

2. greater than 7. **correct**

3. about 7.

4. 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 ($C_6H_5NH_3NO_3$)? K_b for aniline is 4.2×10^{-10}.

Your answer must be within ± 0.4%
Correct answer: 2.70956.

Explanation:
$M_{C_6H_5NH_3NO_3} = 0.16$ M $\qquad K_b = 4.2 \times 10^{-10}$

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

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

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

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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?

1. 2.50×10^{-12} M

2. 5.62×10^{-4} M

3. 4.00×10^{-3} M **correct**

4. 250 M

Explanation:

pH = 2.4, so

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

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016 10.0 points

Which solution has the highest pH?

1. 0.1 M of KHCOO,
$K_{a\,\text{HCOOH}} = 1.8 \times 10^{-4}$

2. 0.1 M of KCl, $K_{a\,\text{HCl}}$ = very large

3. 0.1 M of KCH_3COO,
$K_{a\ HC_2H_3O_2} = 1.8 \times 10^{-5}$

4. 0.1 M of KNO_2, $K_{a\ HNO_2} = 4.5 \times 10^{-4}$

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

Explanation:

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017 10.0 points

What is the pH of a solution that contains 11.7 g of NaCl for every 200 mL of solution?

1. 1.0

2. 10^{-1}

3. 7.0 **correct**

4. 1.0×10^{-7}

Explanation:

m_{NaCl} =11.7 g $V_{soln} = 200$ mL

NaCl completely dissociates in water to give Na^+ and Cl^-, neither of which hydrolyzes and so in aqueous NaCl the H_3O^+ and OH^- ions result from autoionization of water.

$$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$$
$$[H_3O^+] = [OH^-] = 1 \times 10^{-7}$$

and $pH = -\log[H_3O^+] = 7.0$

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018 10.0 points

A 0.010 M solution of a weak acid 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:

pH = 4.35

$$pOH = 14 - pH = 9.65$$

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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 $[H_3O^+]$.

Correct answer: 0.0616595 M.

Explanation:

pH = 1.21 $[H_3O^+] = ?$

$$\begin{aligned} pH &= -\log[H_3O^+] \\ \log[H_3O^+] &= -pH \\ [H_3O^+] &= \text{antilog}\,(-pH) \\ &= 1 \times 10^{-pH} \\ &= 1 \times 10^{-1.21} \\ &= 0.0616595 \text{ M} \end{aligned}$$

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021 (part 2 of 2) 10.0 points

Calculate the $[OH^-]$.

Correct answer: 1.62181×10^{-13} M.

Explanation:

pH = 1.21 $[OH^-] = ?$

$$\begin{aligned} [H_3O^+][OH^-] &= 1 \times 10^{-14}\,M^2 \\ [OH^-] &= \frac{1 \times 10^{-14}\,M^2}{[H_3O^+]} \\ &= \frac{1 \times 10^{-14}\,M^2}{0.0616595 \text{ M}} \\ &= 1.62181 \times 10^{-13} \text{ M} \end{aligned}$$

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 4.9×10^{-10} and $K_w = 1 \times 10^{-14}$.

Correct answer: 11.0044.

Explanation:

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

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

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

023 10.0 points

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

1. $AgNO_3$, $NaCHO_2$, CrI_3

2. NH_4Cl, $C_6H_4NH_3NO_3$, FeI_3

3. $AlCl_3$, $Zn(NO_3)_2$, $KClO_4$

4. CH_3NH_3Cl, KNO_3, NaBz (sodium benzoate)

5. KCH_3COO, NaCN, KF **correct**

Explanation:

024 10.0 points

What is the pH in a solution made by dissolving 0.100 mole of sodium acetate ($NaCH_3COO$) in enough water to make one liter of solution? K_a for CH_3COOH is 1.80×10^{-5}.

1. 8.87 **correct**

2. 9.25

3. 5.13

4. 5.56×10^{-11}

5. 10.25

6. 5.74

7. 5.56×10^{-10}

8. 1.80×10^{-6}

9. 7.46×10^{-6}

10. 1.34×10^{-9}

Explanation:

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. 5.0×10^{-3}

2. 2.0×10^{-5}

3. 1.0×10^{-6}

4. 5.30

5. 5.0×10^{-6} **correct**

6. 1.8×10^{-5}

7. 2.0×10^{-9}

8. 1.0×10^{-3}

Explanation:

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:

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:

$M = 0.28$ M $\qquad P = 3.5\%$

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

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

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

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:

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:

$V_{HNO_3} = 20$ mL $\qquad [HNO_3] = 0.20$ M

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

The balanced equation for this neutralization reaction is

$$2\,HNO_3 + Ba(OH)_2 \rightarrow Ba(NO_3)_2 + 2\,H_2O$$

We determine the moles of HNO_3 used:

$$\begin{aligned} ?\text{ mol } HNO_3 &= 0.020\text{ L soln} \\ &\quad \times \frac{0.20\text{ mol } HNO_3}{1\text{ L soln}} \\ &= 0.0040\text{ mol } HNO_3 \end{aligned}$$

Using the mole ratio from the chemical equation we calculate the moles $Ba(OH)_2$ needed to react with 0.0040 mol of HNO_3:

$$\begin{aligned} ?\text{ mol } Ba(OH)_2 &= 0.0040\text{ mol } HNO_3 \\ &\quad \times \frac{1\text{ mol } Ba(OH)_2}{2\text{ mol } HNO_3} \\ &= 0.0020\text{ mol } Ba(OH)_2 \end{aligned}$$

There are 0.0020 moles $Ba(OH)_2$ 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 → salt + water.

031 10.0 points

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

1. three

2. 1.5 **correct**

3. four

4. eight

5. 0.5

6. two

7. six

8. one

Explanation:

$n_{\rm HCl} = 3$ mol

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

$$2\,\mathrm{HCl} + \mathrm{Ca(OH)_2} \rightarrow \mathrm{CaCl_2} + 2\,\mathrm{H_2O}$$

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

032 10.0 points

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

Correct answer: 2.11563 M.

Explanation:

$V_{\rm RbOH} = 29.1$ mL $\qquad V_{\rm HBr} = 22.51$ mL

[HBr] = 2.735 M $\qquad$ [RbOH] = ?

$$\mathrm{RbOH} + \mathrm{HBr} \longrightarrow \mathrm{RbBr} + \mathrm{H_2O}$$

$$\left(\frac{2.735\,\mathrm{mol\ HBr}}{\mathrm{L}}\right)(22.51\ \mathrm{mL})\left(\frac{\mathrm{L}}{1000\,\mathrm{mL}}\right) = 0.0615649\ \mathrm{mol\ HBr}$$

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

033 10.0 points

For the neutralization reaction involving HNO_3 and LiOH, how much of 2.10 M HNO_3 is needed to neutralize 22.2 L of a 4.66 M LiOH solution? The molar mass of LiOH is 23.95 g/mol. The molar mass of HNO_3 is 63.1 g/mol. The density of the HNO_3 solution is 1.06 g/mL. The density of the 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:
The reaction is
$HNO_3(aq) + LiOH(aq) \rightarrow LiNO_3(aq) + H_2O(\ell)$
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_3 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_3:

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

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034 10.0 points

An aqueous solution is prepared with 2 moles of HCl and 1 mole of $Ca(OH)_2$. The resulting solution contains mainly of

1. water and Cl^-, H^+, and Ca^{2+} ions.

2. water and Cl^- and Ca^{2+} ions. **correct**

3. water and Cl^-, H^+, OH^-, and Ca^{2+} ions.

4. water and Cl^-, OH^-, and Ca^{2+} ions.

Explanation:
$2\,HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2\,H_2O(aq)$
1 mole of $Ca(OH)_2$ reacts with 2 moles of HCl, so there will be no $Ca(OH)_2$ nor HCl left. The $CaCl_2(aq)$ will exist as $Ca^{2+}(aq)$ and $Cl^-(aq)$. The H^+ from the HCl and the OH^- from the $Ca(OH)_2$ have all reacted. Only a miniscule amount of H^+ and OH^- 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:
$[CH_3COOH] = 0.4$ M　　percent = 1.3%
First calculate the concentration of acetic acid that is ionized:

$$CH_3COOH \rightarrow CH_3COO^- + H^+$$

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

Thus

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

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036 10.0 points

Determine the total ionic equation for the reaction between HBr(aq) and $Ba(OH)_2$(aq).

1. $2\,H^+ + 2\,OH^- \rightarrow 2\,H_2O$

2. $2\,Br^- + Ba^{2+} \rightarrow BaBr_2$

3. $2\,HBr + Ba(OH)_2 \rightarrow BaBr_2 + 2\,H_2O$

4. $2\,H^+ + 2\,Br^- + Ba^{2+} + 2\,OH^- \rightarrow Ba^{2+} + 2\,Br^- + 2\,H_2O$ **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^+ and OH^-, is barium bromide ($BaBr_2$), which is soluble. The formula unit equation is

$$Ba(OH)_2(aq) + 2\,HBr(aq) \rightarrow BaBr_2(aq) + 2\,H_2O(\ell).$$

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

$$\left[Ba^{2+}_{(aq)} + 2\,OH^{-}_{(aq)}\right] + 2\left[H^{+}_{(aq)} + Br^{-}_{(aq)}\right] \rightarrow \left[Ba^{2+}_{(aq)} + 2\,Br^{1-}_{(aq)}\right] + 2\,H_2O_{(\ell)}$$

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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:

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038 10.0 points

Identify the products of the chemical equation

$$3\,LiOH + H_3PO_4 \rightarrow$$

1. $3\,LiH + (OH)_3PO_4$

2. $Li_3PO_4 + 3\,H_2O$ **correct**

3. $3\,H + 3\,O_2 + H_3Li_3$

4. $Li_3P + 2\,H_2O + H_3O_5$

Explanation:

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

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039 10.0 points

What are the products of the following reaction?

$$Sr(OH)_2 + 2\,HNO_3 \rightarrow$$

1. $Sr(NO_2)_2 + 2\,H_2O_2$

2. $Sr(NO_3)_2 + 2\,H_2O$ **correct**

3. $SrNO_3 + H_2O$

4. $SrH_2 + HNO_5$

Explanation:

$Sr(OH)_2$ is a base and HNO_3 an acid; they create a salt and water.

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040 10.0 points

Aqueous ammonia can be used to neutralize sulfuric acid (H_2SO_4) and nitric acid (HNO_3) to produce two salts extensively used as fertilizers. They are

1. $(NH_4)_2SO_4$ and NH_4NO_3, respectively. **correct**

2. NH_4SO_4 and NH_4NO_3, respectively.

3. NH_4SO_3 and NH_4OH, 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.

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041 10.0 points

Identify the salt that is produced from the acid-base neutralization reaction between potassium hydroxide and acetic acid (CH_3COOH).

1. potassium cyanide

2. potassium acetate **correct**

3. potassium formate

4. potassium amide

Explanation:

The balanced equation is

$$CH_3CO_2H(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + CH_3CO_2^-(aq) + H_2O(\ell)$$

Potassium acetate (CH_3CO_2K) is the salt.

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042 10.0 points

What volume of 0.585 M $Ca(OH)_2$ 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:

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

$[HCl] = 1.51$ M

The balanced equation for this neutralization reaction is

$$2\,HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2\,H_2O$$

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 $Ca(OH)_2$ needed to react with this amount of HCl:

$$?\text{ mol } Ca(OH)_2 = 23.86\text{ mol HCl} \times \frac{1\text{ mol } Ca(OH)_2}{2\text{ mol HCl}} = 11.93\text{ mol } Ca(OH)_2$$

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

$$?\text{ L } Ca(OH)_2 = 11.93\text{ mol } Ca(OH)_2 \times \frac{1\text{ L soln}}{0.585\text{ mol } Ca(OH)_2} = 20.4\text{ L } Ca(OH)_2$$

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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:

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

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

The base is NaOH. To neutralize, mol H^+ = mol OH^-.

$$n_{H^+} = \frac{0.012\text{ mol}}{\text{L}}(25\text{ mL HCl}) \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{1\text{ mol } H^+}{1\text{ mol HCl}} = 0.0003\text{ mol } H^+ = n_{OH^-} = n_{NaOH}$$

$$M_{NaOH} = \frac{\text{mol}}{\text{L}} = \frac{0.0003\text{ mol NaOH}}{0.04\text{ L}} = 0.0075\text{ M NaOH}$$

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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:

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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 H^+ for every mole of OH^-. Therefore the balanced equation is

$$HCl + NaOH \rightarrow NaCl + H_2O$$

$$\begin{aligned} ?\ \text{mol NaOH} &= 3\ \text{mol HCl} \times \frac{1\ \text{mol NaOH}}{1\ \text{mol HCl}} \\ &= 3\ \text{mol NaOH} \end{aligned}$$