This print-out should have 10 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

What is K_{sp} for Ag_3PO_4, if its molar solubility is 2.7×10^{-6} mol/L?

1. 5.3×10^{-23}

2. 4.8×10^{-22}

3. 2.0×10^{-17}

4. 5.3×10^{-16}

5. 7.3×10^{-12}

6. 1.7×10^{-14}

7. 1.4×10^{-21} **correct**

Explanation:

$S = 2.7 \times 10^{-6}$ mol/L

The solubility equilibrium is

$$Ag_3PO_4(s) \rightleftharpoons 3\,Ag^+(aq) + PO_4^{3-}(aq)$$

$[Ag^+] = 3\,S = 8.1 \times 10^{-6}$ mol/L
$[PO_4^{3-}] = S = 2.7 \times 10^{-6}$ mol/L

$$\begin{aligned} K_{sp} &= [Ag^+]^3[PO_4^{3-}] \\ &= \left(8.1 \times 10^{-6}\right)^3 (2.7 \times 10^{-6}) \\ &= 1.43489 \times 10^{-21} \end{aligned}$$

002 10.0 points

What is the molar solubility of CuBr in 0.5 M NaBr? The K_{sp} is 4.2×10^{-8}.

1. 2.05×10^{-4}

2. 4.20×10^{-8}

3. 4.20×10^{-7}

4. 8.40×10^{-8} **correct**

5. 3.48×10^{-3}

Explanation:

003 10.0 points

A solution is 0.01 M $BaCl_2$ and 0.02 M $SrCl_2$. Which cation can be selectively precipitated first with a concentrated Na_2SO_4 solution? K_{sp} is 1.5×10^{-9} for $BaSO_4$, and 7.6×10^{-7} for $SrSO_4$.

1. Ba^{+2} **correct**

2. Both will precipitate at the same time.

3. Sr^{+2}

Explanation:

Before addition of Na_2SO_4 ...

$[Ba^{2+}] = 0.01$ M $[Sr^{2+}] = 0.02$ M

Cl^- is a spectator ion to be ignored.

Each of the two cation concentrations listed above have a corresponding concentration of SO_4^{2-} that will cause each of these cations to precipitate (shown as '↓') as the sulfate salt. We must calculate these two SO_4^{2-} concentrations and note which of the two is smaller (since we'll get to that one first as we increase the SO_4^{2-} concentration from zero).

$$SrSO_4 \rightarrow Sr^{2+} + SO_4^{2-}$$

$$\begin{aligned} K_{sp} &= [Sr^{2+}]\,[SO_4^{2-}] \\ 7.6 \times 10^{-7} &= (0.02\text{ M})\,[SO_4^{2-}] \end{aligned}$$

$[SO_4^{2-}] = 3.8 \times 10^{-5}$ M for $SrSO_4\downarrow$

$$BaSO_4 \rightarrow Ba^{2+} + SO_4^{2-}$$

$$\begin{aligned} K_{sp} &= [Ba^{2+}]\,[SO_4^{2-}] \\ 1.5 \times 10^{-9} &= (0.01\text{ M})\,[SO_4^{2-}] \end{aligned}$$

$[SO_4^{2-}] = 1.5 \times 10^{-8}$ M for $BaSO_4\downarrow$

The concentration of sulfate ion that will cause $BaSO_4$ to precipitate is smaller than the concentration that will cause $SrSO_4$ to precipitate. Therefore, Ba^{+2} will be precipitated first.

004 10.0 points

Rank following salts from least to most soluble:

BiI	$K_{sp} = 7.7 \times 10^{-19}$
$Cd_3(AsO_4)_2$	$K_{sp} = 2.2 \times 10^{-33}$
$AlPO_4$	$K_{sp} = 9.8 \times 10^{-21}$
$CaSO_4$	$K_{sp} = 4.9 \times 10^{-5}$

1. $BiI < Cd_3(AsO_4)_2 < CaSO_4 < AlPO_4$

2. $AlPO_4 < BiI < Cd_3(AsO_4)_2 < CaSO_4$ **correct**

3. $CaSO_4 < AlPO_4 < BiI < Cd_3(AsO_4)_2$

4. $Cd_3(AsO_4)_2 < CaSO_4 < AlPO_4 < BiI$

Explanation:

Molar solubility can be approximated by taking the n^{th} root of the K_{sp} where n is the number of ions in the salt. Doing so results in approximate molar solubilities of 10^{-10}, 10^{-7}, 10^{-11} and 10^{-3} for bismuth iodide, cadmium arsenate, aluminum phosphate and calcium sulfate, respectively. Arranging these from least to greatest produces: $AlPO_4 < BiI < Cd_3(AsO_4)_2 < CaSO_4$.

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005 10.0 points

What is the molar solubility of $PbCl_2$ in an aqueous solution of 0.5 M NaCl? The K_{sp} of $PbCl_2$ is 1.14×10^{-5}.

1. 4.56×10^{-4}

2. 2.28×10^{-5}

3. 1.14×10^{-5}

4. 2.28×10^{-4}

5. 4.56×10^{-5} **correct**

Explanation:

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006 10.0 points

What is the molarity of a $FeSO_4$ solution if 25.0 mL of it reacts with 38.0 mL of 0.1214 M $KMnO_4$ solution?

$$MnO_4^- + 8\,H^+ + 5\,Fe^{2+} \rightarrow Mn^{2+} + 5\,Fe^{3+} + 4\,H_2O$$

1. 0.185 M

2. 0.0798 M

3. 0.923 M **correct**

4. 0.399 M

5. 0.426 M

Explanation:

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007 10.0 points

What is the mass in grams of NH_3 titrated to the endpoint of a reaction with 10 mL of 0.02 N H_2SO_4?

$$H_2SO_4 + 2\,NH_3 \rightarrow SO_4^{-2} + 2\,NH_4^+$$

1. 0.0068 g

2. 0.0034 g **correct**

3. 0.0017 g

4. 0.0002 g

5. 0.0001 g

Explanation:

$V_{H_2SO_4} = 10$ mL $N_{H_2SO_4} = 0.02$ N

$MW_{H_2SO_4} = 98$ g/mol $MW_{NH_3} = 17$ g/mol

$$\text{EqW} = \frac{\text{MW}}{\text{number reactive sites in molecule}}, \text{ so}$$

$$\text{EqW}_{H_2SO_4} = \frac{98 \text{ g/mol}}{2 \text{ eq/mol}} = 49 \text{ g/eq } H_2SO_4$$

$$\text{EqW}_{NH_3} = \frac{17 \text{ g/mol}}{1 \text{ eq/mol}} = 17 \text{ g/eq } NH_3$$

$$\text{eq acid} = \text{eq base}$$

$$? \text{ eq } H_2SO_4 = 0.02 \text{ N} \times \frac{\text{eq}}{\text{N} \cdot \text{L soln}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 10 \text{ mL} = 0.0002 \text{ eq } H_2SO_4 = 0.0002 \text{ eq } NH_3$$

$$? \text{ g } NH_3 = 0.0002 \text{ eq } NH_3 \times \frac{17 \text{ g } NH_3}{\text{eq } NH_3} = 0.0034 \text{ g } NH_3$$

008 10.0 points

An animal cell assumes its normal volume when it is placed in a solution with a total solute molarity of 0.3 M. If the cell is placed in a solution with a total solute molarity of 0.1 M,

1. the escaping tendency of water in the cell increases.

2. water enters the cell, causing expansion. **correct**

3. water leaves the cell, causing contraction.

4. no movement of water takes place.

Explanation:

009 10.0 points

A decrease in temperature usually (decreases, increases, does not change) the solubility of salts in water.

1. increases

2. does not change

3. decreases **correct**

Explanation:

Most salts are less soluble at lower temperature.

010 10.0 points

Rank the following compounds from most soluble to least soluble. Assume that all bonds except the OH are ionic. (You can estimate this ranking without using a calculator.)

Compound	K_{sp}
Bi_2S_3	1.0×10^{-97}
$Fe(OH)_2$	1.6×10^{-14}
PbI_2	2.6×10^{-13}
HgS	1.6×10^{-52}

1. $Fe(OH)_2 > PbI_2 > HgS > Bi_2S_3$

2. $HgS > PbI_2 > Fe(OH)_2 > Bi_2S_3$

3. $PbI_2 > Fe(OH)_2 > Bi_2S_3 > HgS$ **correct**

4. $Bi_2S_3 > Fe(OH)_2 > HgS > PbI_2$

5. $PbI_2 > Fe(OH)_2 > HgS > Bi_2S_3$

Explanation: