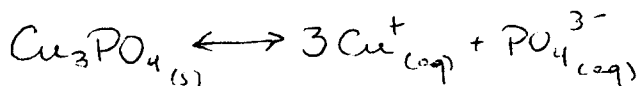


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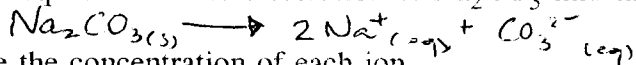
### Solubility Open-ended Review

1. Write an equation that describes the equilibrium present in a saturated solution of  $\text{Cu}_3\text{PO}_4$ .



2. 53 g of  $\text{Na}_2\text{CO}_3$  are dissolved in sufficient water to make 5.0 L of solution.

- A. Write the equation for the dissolution of  $\text{Na}_2\text{CO}_3$  into its aqueous ions.



- B. Calculate the concentration of each ion.

$$\text{FW Na}_2\text{CO}_3 = 2(23.0) + 12.0 + 3(16.0) = 106 \text{ g/mol}$$

$$53 \text{ g} \div 106 \text{ g/mol} = 0.50 \text{ mol}$$

$$[\text{Na}_2\text{CO}_3] = 0.50 \text{ mol} / 5.0 \text{ L} = 0.10 \text{ M}$$

$$[\text{Na}^+] = 2(0.10 \text{ M}) = 0.20 \text{ M}$$

$$[\text{CO}_3^{2-}] = 0.10 \text{ M}$$

- C. Describe the changes in entropy and enthalpy as the  $\text{Na}_2\text{CO}_3$  dissolves.

entropy increases as ions are formed.

enthalpy increases as the dissolving process is endothermic

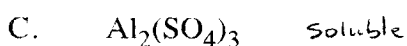
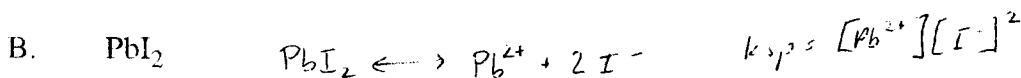
- D. When the solution was prepared, some doubt existed that  $\text{Na}_2\text{SO}_4$  might have been used by mistake. Describe a suitable precipitation test that will confirm the presence of  $\text{CO}_3^{2-}$  ions in the solution.

Add any cation except: Alkali ions,  $\text{H}^+$ ,  $\text{NH}_4^+$  which ppt neither

or  $\text{Ag}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Pb}^{2+}$  which ppt both

possible answers include  $\text{Al}^{3+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Cu}^{2+}$  etc.

3. Write an equilibrium expression and an equation that describes the equilibrium for only those salts that have low solubility:

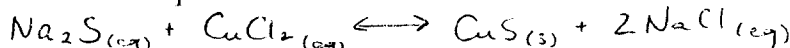


4. A 1.0 M solution of sodium sulphide is added to a 1.0 M solution of copper II chloride resulting in the formation of a precipitate.

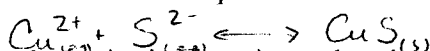
- A. Write the name and formula of the precipitate.

Copper II sulphide       $\text{CuS}$

- B. Write the full equation for the reaction.

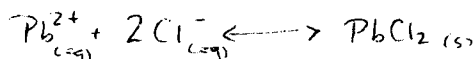


- C. Write the net ionic equation for the reaction.

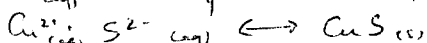
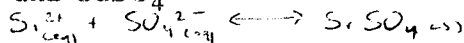


5. Write balanced net ionic equations showing the formation of each precipitate formed when equal volumes of the following 0.50 M solutions are mixed:

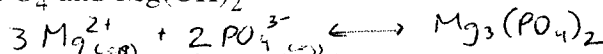
- A.  $\text{MgCl}_2$  and  $\text{Pb}(\text{NO}_3)_2$



- B.  $\text{SrS}$  and  $\text{CuSO}_4$



- C.  $(\text{NH}_4)_3\text{PO}_4$  and  $\text{Mg}(\text{OH})_2$



6. Calculate the concentration of each ion in the following saturated solutions:

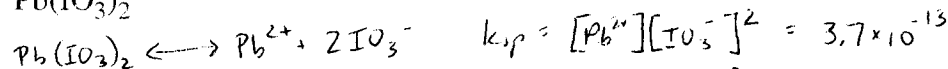
- A.  $\text{Al}(\text{OH})_3$        $K_{sp} = 3.0 \times 10^{-33}$        $\text{Al}(\text{OH})_3 \rightleftharpoons \text{Al}^{3+} + 3\text{OH}^-$        $K_{sp} = [\text{Al}^{3+}][\text{OH}^-]^3$

Let  $x = \text{solubility}$        $3.0 \times 10^{-33} = x(3x)^3$        $[\text{Al}^{3+}] = 3.3 \times 10^{-9} \text{ M}$

then  $[\text{Al}^{3+}] = x$        $x = 3.25 \times 10^{-9} \text{ M}$        $[\text{OH}^-] = 9.7 \times 10^{-9} \text{ M}$

$[\text{OH}^-] = 3x$

- B.  $\text{Pb}(\text{IO}_3)_2$



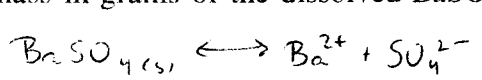
Let  $x = \text{solubility}$        $3.7 \times 10^{-13} = x(2x)^2$        $[\text{Pb}^{2+}] = 4.5 \times 10^{-5} \text{ M}$

then  $[\text{Pb}^{2+}] = x$        $x = 4.52 \times 10^{-5}$        $[\text{IO}_3^-] = 9.0 \times 10^{-5} \text{ M}$

$[\text{IO}_3^-] = 2x$

key.

7. A suspension of barium sulphate is used to improve the quality of X-rays in the digestive system. If the patient is required to drink 0.400 L of this suspension, calculate the actual mass in grams of the dissolved BaSO<sub>4</sub>.



$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

Let  $x$  = solubility of BaSO<sub>4</sub>  
then  $[\text{Ba}^{2+}] = [\text{SO}_4^{2-}] = x$

$$1.1 \times 10^{-10} = x^2$$

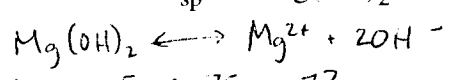
$$x = 1.05 \times 10^{-5} \text{ M}$$

$$\text{mols BaSO}_4 = 1.05 \times 10^{-5} \text{ M} \times 0.400 \text{ L} = 4.20 \times 10^{-6} \text{ mols}$$

$$\text{Fw BaSO}_4 = 137.3 + 32.1 + 4(16.0) = 233.4 \text{ g/mol}$$

$$\text{mass} = 4.20 \times 10^{-6} \text{ mol} \times 233.4 \text{ g/mol} = 9.8 \times 10^{-4} \text{ g}$$

8. Calculate the K<sub>sp</sub> for Mg(OH)<sub>2</sub> if the solubility of magnesium hydroxide is 7.6 mg/L.



$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$$

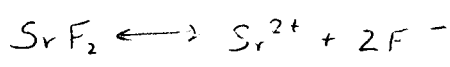
$$\text{Fw Mg(OH)}_2 = 24.3 + 2(16.0) + 2(1.0) = 58.3 \text{ g/mol}$$

$$[\text{Mg}^{2+}] = 7.6 \text{ mg/L} \div 58.3 \text{ g/mol} \times 1000 \text{ mg/g} = 1.30 \times 10^{-4} \text{ M}$$

$$[\text{OH}^-] = 2(1.30 \times 10^{-4} \text{ M}) = 2.61 \times 10^{-4} \text{ M}$$

$$K_{sp} = (1.30 \times 10^{-4})(2.61 \times 10^{-4})^2 = 8.8 \times 10^{-12}$$

9. What maximum [F<sup>-</sup>] exists in a solution in which the [Sr<sup>2+</sup>] = 4.4 x 10<sup>-3</sup> M?

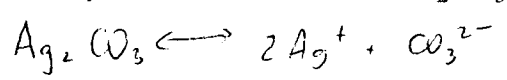


$$K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 = 4.3 \times 10^{-9}$$

$$[\text{F}^-] = 9.9 \times 10^{-4} \text{ M}$$

$$[\text{F}^-]^2 = \frac{4.3 \times 10^{-9}}{4.4 \times 10^{-3}} = 9.77 \times 10^{-7}$$

10. Show by calculation if a precipitate forms when 10.0 mL of 0.010 M AgNO<sub>3</sub> are mixed with an equal volume of 0.10 M Na<sub>2</sub>CO<sub>3</sub>.



$$K_{sp} = [\text{Ag}^+]^2[\text{CO}_3^{2-}] = 8.5 \times 10^{-12}$$

$$[\text{Ag}^+] = \frac{10.0 \text{ mL} \times 0.010 \text{ M}}{20.0 \text{ mL}} = 0.0050 \text{ M}$$

$$[\text{CO}_3^{2-}] = \frac{10.0 \text{ mL} \times 0.10 \text{ M}}{20.0 \text{ mL}} = 0.050 \text{ M}$$

$$Q_{IP} = (0.0050)^2(0.050) = 1.25 \times 10^{-6}$$

$Q_{IP} > K_{sp}$  ∴ a ppt forms.

11. A solution may contain Ba<sup>2+</sup> and/or Al<sup>3+</sup>. Describe a procedure to confirm the presence or absence of these ions.

① Add K<sub>2</sub>SO<sub>4</sub> to ppt Ba<sup>2+</sup>

② Add K<sub>2</sub>SO<sub>3</sub> to ppt Al<sup>3+</sup>

12.

How many moles of PbI<sub>2</sub> would dissolve in water in which the [Pb<sup>2+</sup>] = 5.0 x 10<sup>-6</sup> M?

- no volume given omit question.

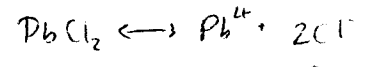
13. A 25.0 mL sample of saturated PbCl<sub>2</sub> solution is titrated to the endpoint with 48.1 mL of 0.015 M AgNO<sub>3</sub> solution. Calculate the K<sub>sp</sub> of PbCl<sub>2</sub>.

$$\text{mols Ag}^+ = 48.1 \text{ mL} \times 0.015 \text{ M} = 0.722 \text{ mmol}$$

$$\text{mols Ag}^+ = \text{mols Cl}^-$$

$$[\text{Cl}^-] = \frac{0.722 \text{ mmol}}{25.0 \text{ mL}} = 2.89 \times 10^{-2} \text{ M}$$

$$[\text{Pb}^{2+}] = \frac{1}{2} (2.89 \times 10^{-2} \text{ M}) = 1.44 \times 10^{-2} \text{ M}$$



$$K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

$$K_{sp} = (1.44 \times 10^{-2})(2.89 \times 10^{-2})^2$$

$$= 1.2 \times 10^{-5}$$