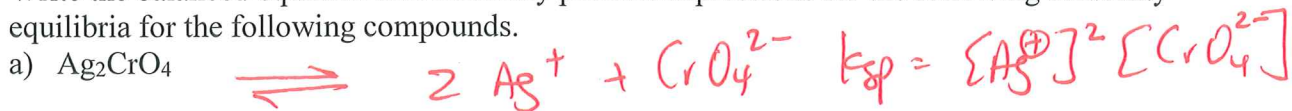
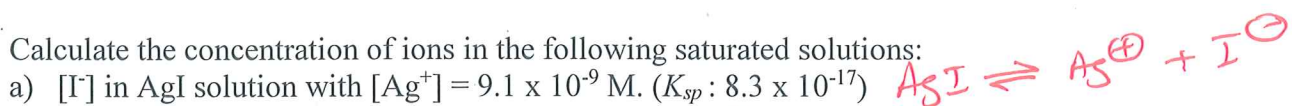


- 1) Write the balanced equation and solubility product expressions for the following solubility equilibria for the following compounds.



- 2) Calculate the concentration of ions in the following saturated solutions:



$$8.3 \times 10^{-17} = [\text{I}^-] (9.1 \times 10^{-9})$$

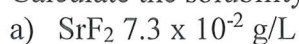
$$[\text{I}^-] = \frac{8.3 \times 10^{-17}}{9.1 \times 10^{-9}} = \boxed{9.1 \times 10^{-9} \text{ M}}$$



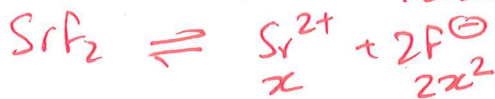
$$K_{sp} = [\text{Al}^{3+}] [\text{OH}^-]^3$$

$$[\text{Al}^{3+}] = \frac{K_{sp}}{[\text{OH}^-]^3} = \frac{1.8 \times 10^{-33}}{(2.9 \times 10^{-9})^3} = \boxed{7.4 \times 10^{-8} \text{ M}}$$

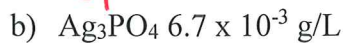
- 3) Calculate the solubility products ( $K_{sp}$ ) for the following compounds:



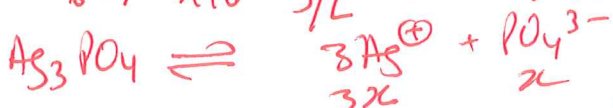
$$7.3 \times 10^{-2} \text{ g/L} \times \frac{1 \text{ mol SrF}_2}{125.6 \text{ g}} = 5.8 \times 10^{-4} \text{ mol/L } (x)$$



$$K_{sp} = (x)(2x)^2 = 4x^3 = 4(5.8 \times 10^{-4})^3 = \boxed{7.8 \times 10^{-10}}$$

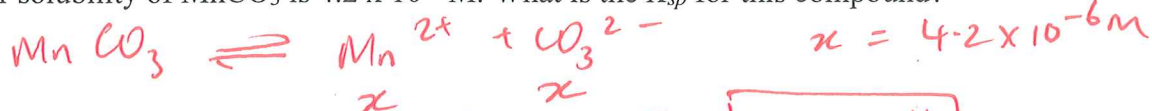


$$6.7 \times 10^{-3} \text{ g/L} \times \frac{1 \text{ mol}}{418.7 \text{ g}} = 1.6 \times 10^{-5} \text{ mol/L } (x)$$



$$K_{sp} = (3x)^3(x) = 27x^4 = 27(1.6 \times 10^{-5})^4 = \boxed{1.8 \times 10^{-18}}$$

- 4) The molar solubility of  $\text{MnCO}_3$  is  $4.2 \times 10^{-6} \text{ M}$ . What is the  $K_{sp}$  for this compound?



$$K_{sp} = (x)(x) = x^2 = (4.2 \times 10^{-6})^2 = \boxed{1.8 \times 10^{-11}}$$

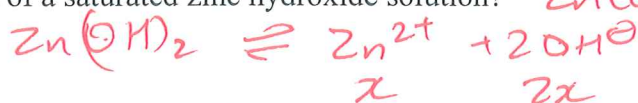
- 5) The solubility of an ionic compound MX (mol mass = 346 g/mol) is  $4.63 \times 10^{-3}$  g/L. What is the  $K_{sp}$  for this compound?

$$4.63 \times 10^{-3} \frac{\text{g}}{\text{L}} \times \frac{1 \text{ mol}}{346 \text{ g}} = 1.34 \times 10^{-5} = s$$



$$K_{sp} = (s)(s) = (1.34 \times 10^{-5})^2 = \boxed{1.8 \times 10^{-10}}$$

- 6) What is the pH of a saturated zinc hydroxide solution?  $Zn(OH)_2$ .  $K_{sp} = 1.8 \times 10^{-14}$ .



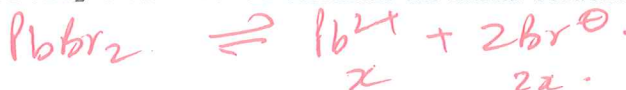
$$1.8 \times 10^{-14} = (x)(2x)^2$$

$$x = 1.65 \times 10^{-5} \equiv [OH^{-}]$$

$$pOH = 4.48 \quad pH = 14 - 4.48 = \boxed{9.52}$$

- 7) The solubility product of  $PbBr_2$  is  $8.9 \times 10^{-6}$ . Determine the molar solubility in:

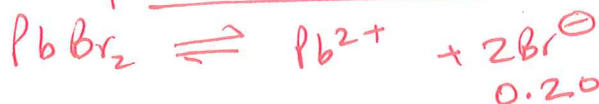
a) pure water,



$$8.9 \times 10^{-6} = (x)(2x)^2$$

$$\boxed{x = 1.3 \times 10^{-2} \text{ M}}$$

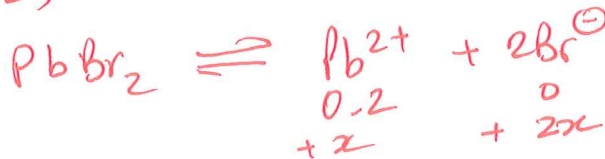
b) 0.20 M KBr



$$8.9 \times 10^{-6} = (x)(0.20 + 2x)^2 \quad \text{ignore.}$$

$$\frac{8.9 \times 10^{-6}}{(0.2)^2} = x = \boxed{2.2 \times 10^{-4} \text{ M}}$$

c) 0.20 M  $Pb(NO_3)_2$ .



$$8.9 \times 10^{-6} = (0.2 + x)(2x)^2 \quad \text{ignore.}$$

$$\frac{8.9 \times 10^{-6}}{0.2} = (2x)^2 \quad \boxed{x = 3.3 \times 10^{-3} \text{ M}}$$

- 8) If 20.0 mL of 0.10 M  $\text{Ba}(\text{NO}_3)_2$  is added to 50.0 mL of 0.10 M  $\text{Na}_2\text{CO}_3$ , will  $\text{BaCO}_3$  precipitate?  $K_{sp} = 8.1 \times 10^{-9}$ .

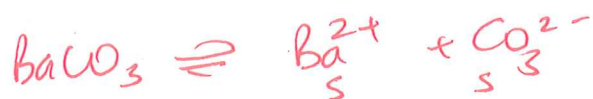
$$20 \text{ mL} \times \frac{0.1 \text{ mol}}{1000 \text{ mL}} = 2 \times 10^{-3} \text{ mol Ba}^{2+}$$

$$50 \text{ mL} \times \frac{0.1 \text{ mol}}{1000 \text{ mL}} = 5 \times 10^{-3} \text{ mol CO}_3^{2-}$$

} added.  
total vol.  
20 + 50 = 70 mL

$$\frac{2 \times 10^{-3}}{0.07 \text{ L}} = 2.9 \times 10^{-2} \text{ M Ba}^{2+}$$

$$\frac{5 \times 10^{-3}}{0.07 \text{ L}} = 7.1 \times 10^{-2} \text{ M CO}_3^{2-}$$



$$Q_{sp} = (2.9 \times 10^{-2}) (7.1 \times 10^{-2}) = 2.1 \times 10^{-3}$$

$$Q_{sp} > K_{sp} \quad | \quad \text{yes ppt forms.}$$

- 9) Which of the following ionic compounds will be more soluble in acid solution than pure water: a)  $\text{BaSO}_4$ ,  $\text{PbCl}_2$ ,  $\text{Fe}(\text{OH})_3$ ,  $\text{CaCO}_3$ .

a)  $\text{BaSO}_4$   
↓  
little bit more in acid b/c  $\text{SO}_4^{2-}$  conj base of strong acid so weak base.

b)  $\text{PbCl}_2$   
unaffected.  
 $\text{Cl}^-$  conj base of strong acid

c)  $\text{Fe}(\text{OH})_3$   
more sol. in acid  
 $\text{OH}^-$  more sol. in acid

d)  $\text{CaCO}_3$   
more sol. in  $\text{H}^+$   
 $\text{CO}_3^{2-}$  is weak base