

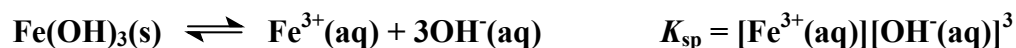
Marks
2

- Calculate the molar solubility of $\text{Fe}(\text{OH})_3$ in a $\text{pH} = 5.0$ buffer solution. The solubility product constant of $\text{Fe}(\text{OH})_3$ is $4 \times 10^{-38} \text{ M}^4$.

Using $\text{pOH} = -\log_{10}([\text{OH}^-(\text{aq})])$ and $\text{pH} + \text{pOH} = 14.0$:

$$\text{pOH} = (14.0 - 5.0) = 9.0 \text{ and} \\ [\text{OH}^-(\text{aq})] = 1 \times 10^{-9} \text{ M}$$

The solubility equilibrium and product are :



Hence,

$$[\text{Fe}^{3+}(\text{aq})] = \frac{K_{\text{sp}}}{[\text{OH}^-(\text{aq})]^3} = \frac{(4 \times 10^{-38})}{(1.0 \times 10^{-9})^3} = 4 \times 10^{-11} \text{ M}$$

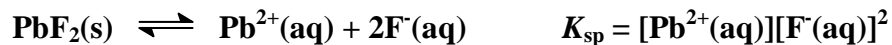
As $\text{Fe}(\text{OH})_3(\text{s})$ dissolves to give 1 $\text{Fe}^{3+}(\text{aq})$, this is also the molar solubility.

Answer: $4 \times 10^{-11} \text{ M}$

Marks
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- The molar solubility of lead(II) fluoride, PbF_2 , is found to be $2.6 \times 10^{-3} \text{ M}$ at 25°C . Calculate the value of K_{sp} for this compound at this temperature.

The solubility equilibrium and constant for $\text{PbF}_2(\text{s})$ are,



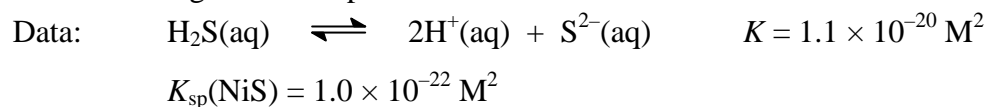
As one moles of $\text{Pb}^{2+}(\text{aq})$ and two moles of $\text{F}^{-}(\text{aq})$ are produced for every mole of $\text{PbF}_2(\text{s})$ which dissolves, $[\text{Pb}^{2+}(\text{aq})] = 2.6 \times 10^{-3} \text{ M}$ and $[\text{F}^{-}(\text{aq})] = (2 \times 2.6 \times 10^{-3}) = 5.2 \times 10^{-3} \text{ M}$. Hence,

$$K_{\text{sp}} = (2.6 \times 10^{-3}) \times (5.2 \times 10^{-3})^2 = 7.0 \times 10^{-8}$$

$$K_{\text{sp}} = 7.0 \times 10^{-8}$$

Marks
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- The pH of a solution can be controlled by adding small amounts of gaseous HCl. Assuming no change in volume, calculate what the pH of the solution must be to just dissolve 1.00 g of NiS suspended in 1.0 L of water.



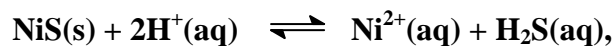
For the dissolution $\text{NiS}(\text{s}) \rightleftharpoons \text{Ni}^{2+}(\text{aq}) + \text{S}^{2-}(\text{aq})$,

$$K_{\text{sp}} = [\text{Ni}^{2+}(\text{aq})][\text{S}^{2-}(\text{aq})].$$

For the acid dissociation, $\text{H}_2\text{S}(\text{aq}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{S}^{2-}(\text{aq})$,

$$K = \frac{[\text{H}^+(\text{aq})]^2[\text{S}^{2-}(\text{aq})]}{[\text{H}_2\text{S}(\text{aq})]}$$

Combining these two equilibria gives, for the overall reaction,



$$K_{\text{eq}} = \frac{[\text{Ni}^{2+}(\text{aq})][\text{H}_2\text{S}(\text{aq})]}{[\text{H}^+(\text{aq})]^2} = \frac{K_{\text{sp}}}{K} = \frac{1.0 \times 10^{-22}}{1.1 \times 10^{-20}} = 9.1 \times 10^{-3}$$

The formula mass of NiS is (58.69 (Ni) + 32.07 (S)) $\text{g mol}^{-1} = 90.76 \text{ g mol}^{-1}$. The amount in 1.00 g is therefore,

$$\text{number of moles} = \frac{\text{mass}(\text{g})}{\text{molar mass}(\text{g mol}^{-1})} = \frac{1.00 \text{ g}}{90.76 \text{ g mol}^{-1}} = 0.0110 \text{ mol}$$

In the dissolution $\text{NiS}(\text{s}) \rightleftharpoons \text{Ni}^{2+}(\text{aq}) + \text{S}^{2-}(\text{aq})$, so if 0.0110 mol of NiS(s) completely dissolves in 1.0 L, $[\text{Ni}^{2+}(\text{aq})] = 0.011 \text{ M}$. From the overall reaction, each mole of NiS which dissolves produces one mole of $\text{H}_2\text{S}(\text{aq})$ so $[\text{H}_2\text{S}(\text{aq})] = 0.011 \text{ M}$. Hence,

$$[\text{H}^+(\text{aq})]^2 = \frac{(0.011 \text{ M})(0.011 \text{ M})}{9.1 \times 10^{-3}} \quad \text{or} \quad [\text{H}^+(\text{aq})] = 0.12 \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}^+(\text{aq})] = -\log_{10}(0.12) = 0.94$$

pH = 0.94

Marks
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- A champagne bottle is filled with 750 mL of wine, leaving 10.0 mL of air at atmospheric pressure when it is sealed with a cork. After fermentation, the pressure inside the bottle is 6.0 atm at 20 °C. Assume that the gas produced is entirely CO₂ and that its solubility in the wine is the same as in water. What mass of CO₂ has been produced by the fermentation?

Data: The mole fraction solubility of CO₂ in water is 7.1×10^{-4} at 293 K and 1.0 atm.

The molar mass of H₂O is (16.00 (O) + 2 × 1.008 (H)) g mol⁻¹ = 18.016 g mol⁻¹. Assuming that the wine is entirely water with a density of 1.0 g mL⁻¹, the bottle contains 750 g of water or:

$$\text{number of moles of water} = \frac{\text{mass}}{\text{molar mass}} = \frac{750 \text{ g}}{18.016 \text{ g mol}^{-1}} = 41.67 \text{ mol}$$

The mole fraction of CO₂ in water, X_{CO_2} , is given by:

$$X_{\text{CO}_2} = \frac{n_{\text{CO}_2(\text{aq})}}{n_{\text{CO}_2(\text{aq})} + n_{\text{H}_2\text{O}(\text{l})}} = \frac{n_{\text{CO}_2(\text{aq})}}{n_{\text{CO}_2(\text{aq})} + (41.67 \text{ mol})} = 7.1 \times 10^{-4}$$

Hence, the number of moles of CO₂ in the wine *before* fermentation (1.0 atm) is given by:

$$\begin{aligned} n_{\text{CO}_2(\text{aq})} &= (7.1 \times 10^{-4})(n_{\text{CO}_2(\text{aq})} + 41.67) \\ &= 7.1 \times 10^{-4} n_{\text{CO}_2(\text{aq})} + (7.1 \times 10^{-4} \times 41.67) \end{aligned}$$

$$n_{\text{CO}_2(\text{aq})}(1.0 - 7.1 \times 10^{-4}) = (7.1 \times 10^{-4} \times 41.67)$$

$$n_{\text{CO}_2(\text{aq})} = 0.0296 \text{ mol}$$

After fermentation, the pressure is 6.0 atm so $n_{\text{CO}_2(\text{aq})} = 6.0 \times 0.0296 \text{ mol}$. The number of moles of CO₂ produced by the fermentation and dissolved in the wine is therefore:

$$n_{\text{CO}_2(\text{aq})} = (6.0 - 1.0) \times 0.0296 \text{ mol} = 0.148 \text{ mol}$$

The increase in air pressure of 5.0 atm is due to extra CO₂(g). As 1 atm = 101.3 kPa, $P = (5.0 \times 101.3) \text{ kPa} = 506.5 \text{ kPa}$. The volume of air = 10.0 mL = 0.0100 L = $1.00 \times 10^{-5} \text{ m}^3$. Using the ideal gas equation, $PV = nRT$, the number of moles of CO₂(g) is:

$$n_{\text{CO}_2(\text{g})} = \frac{PV}{RT} = \frac{(5.065 \times 10^5 \text{ Pa})(1.00 \times 10^{-5} \text{ m}^3)}{(8.314 \text{ m}^3 \text{ Pa K}^{-1} \text{ mol}^{-1})(20+273)\text{K}} = 0.00208 \text{ mol}$$

Overall:

$$n_{\text{CO}_2} = n_{\text{CO}_2(\text{aq})} + n_{\text{CO}_2(\text{g})} = (0.148 + 0.002) \text{ mol} = 0.150 \text{ mol}$$

The molar mass of CO₂ is (12.01 (C) + 2 × 16.00 (O)) g mol⁻¹ = 44.01 g mol⁻¹. Hence, the mass of CO₂ produced by fermentation is given by:

$$\begin{aligned} \text{mass} &= \text{number of moles} \times \text{molar mass} \\ &= (0.150 \text{ mol}) \times (44.01 \text{ g mol}^{-1}) = 6.6 \text{ g} \end{aligned}$$

Answer: 6.6 g

ANSWER CONTINUES ON THE NEXT PAGE

After the bottle has been opened and all of the bubbles have been released, what volume of CO₂ has escaped? Assume all the CO₂ produced escapes.

When the cork is released, the pressure returns to 1.0 atm. The amount of CO₂ that will remain dissolved is therefore, from above, $n_{\text{CO}_2(\text{aq})} = 0.0296$ mol.

The amount of CO₂ which escapes is therefore:

$$n_{\text{CO}_2(\text{g})} = (0.150 - 0.0296) \text{ mol} = 0.120 \text{ mol}$$

At 1.0 atm = 101.3 kPa, this will occupy a volume:

$$V = \frac{nRT}{P} = \frac{(0.120 \text{ mol})(8.314 \text{ m}^3 \text{ Pa K}^{-1} \text{ mol}^{-1})((20+273)\text{K})}{(1.013 \times 10^5 \text{ Pa})}$$

$$= 2.9 \times 10^{-3} \text{ m}^3 = 2.9 \text{ L}$$

$$\text{Answer: } 2.9 \times 10^{-3} \text{ m}^3 = 2.9 \text{ L}$$

- The solubility of BaF_2 in water is 1.30 g L^{-1} . Calculate the solubility product for BaF_2 .

Marks
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The molar mass of BaF_2 is $(137.34 \text{ (Ba)} + 2 \times 19.00 \text{ (F)}) \text{ g mol}^{-1} = 175.34 \text{ g mol}^{-1}$.
As 1.30 g dissolves in one litre, this corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.30 \text{ g}}{175.34 \text{ g mol}^{-1}} = 0.00741 \text{ mol}$$

As BaF_2 dissolves to give $\text{Ba}^{2+}(\text{aq}) + 2\text{F}^{-}(\text{aq})$, dissolution of 0.00741 mol in one litre will produce $[\text{Ba}^{2+}(\text{aq})] = 0.00741 \text{ M}$ and $[\text{F}^{-}(\text{aq})] = 2 \times 0.00741 \text{ M} = 0.0148 \text{ M}$. Hence:

$$K_{\text{sp}} = [\text{Ba}^{2+}(\text{aq})][\text{F}^{-}(\text{aq})]^2 = (0.00741)(0.0148)^2 = 1.6 \times 10^{-6}$$

Answer: 1.6×10^{-6}

- A mixture of NaCl (5.0 g) and AgNO_3 (5.0 g) was added to 1.0 L of water. What are the concentrations of $\text{Ag}^{+}(\text{aq})$, $\text{Cl}^{-}(\text{aq})$ and $\text{Na}^{+}(\text{aq})$ ions in solution after equilibrium has been established? $K_{\text{sp}}(\text{AgCl}) = 1.8 \times 10^{-10}$.

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The molar mass of NaCl is $(22.99 \text{ (Na)} + 35.45 \text{ (Cl)}) \text{ g mol}^{-1} = 58.44 \text{ g mol}^{-1}$.
Hence:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{5.0 \text{ g}}{58.44 \text{ g mol}^{-1}} = 0.0856 \text{ mol}$$

As NaCl dissolves to give $\text{Na}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq})$, dissolution of this amount in one litre will give $[\text{Na}^{+}(\text{aq})] = 0.0856 \text{ M}$ and $[\text{Cl}^{-}(\text{aq})] = 0.0856 \text{ M}$.

The molar mass of AgNO_3 is $(107.87 \text{ (Ag)} + 14.01 \text{ (N)} + 3 \times 16.00 \text{ (O)}) \text{ g mol}^{-1} = 169.88 \text{ g mol}^{-1}$. Hence:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{5.0 \text{ g}}{169.88 \text{ g mol}^{-1}} = 0.0294 \text{ mol}$$

As AgNO_3 dissolves to give $\text{Ag}^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq})$, dissolution of this amount in one litre will give $[\text{Ag}^{+}(\text{aq})] = 0.0294 \text{ M}$.

Precipitation of $\text{AgCl}(\text{s})$ follows $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$. As 0.0294 mol of Ag^{+} ions and 0.0856 mol of Cl^{-} ions are present, the former is limiting and so 0.0294 mol of $\text{AgCl}(\text{s})$ will form leaving $(0.0856 - 0.0294) \text{ mol} = 0.0562 \text{ mol}$ of Cl^{-} ions. Hence, after precipitation, $[\text{Cl}^{-}(\text{aq})] = 0.0562 \text{ M}$.

$\text{AgCl}(\text{s})$ has a very low solubility and dissolves to give $\text{Ag}^{+}(\text{aq})$ and $\text{Cl}^{-}(\text{aq})$ with

$$K_{\text{sp}} = [\text{Ag}^{+}(\text{aq})][\text{Cl}^{-}(\text{aq})] = 1.8 \times 10^{-10}$$

Hence:

$$[\text{Ag}^{+}(\text{aq})] = K_{\text{sp}} / [\text{Cl}^{-}(\text{aq})] = (1.8 \times 10^{-10}) / (0.0562) \text{ M} = 3.2 \times 10^{-9} \text{ M}$$

$$[\text{Ag}^{+}(\text{aq})] = 3.2 \times 10^{-9} \text{ M}$$

$$[\text{Cl}^{-}(\text{aq})] = 0.0562 \text{ M}$$

$$[\text{Na}^{+}(\text{aq})] = 0.0856 \text{ M}$$