


12.2.5 Solubility product constant

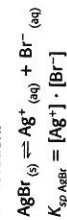
 Textbook, p. 336 to 338

1. a) Can the solubility of magnesium hydroxide ($\text{Mg}(\text{OH})_2$) and that of lead chloride (PbCl_2) be compared by only comparing their solubility product constants? Justify your answer.
Yes, they can be compared this way, since the two ionic compounds are composed of ions in the same proportions, that is, the compounds give one cation for two anions.
- b) If, at 25°C , the solubility product constants are 5.61×10^{-12} for the magnesium hydroxide and 1.6×10^{-9} for the lead chloride, which of the two compounds has the greatest solubility in water at 25°C ? Justify your answer.
The lead chloride, since its solubility product constant is greater.
2. a) Can the solubility of silver bromide (AgBr) and that of iron hydroxide ($\text{Fe}(\text{OH})_2$) be compared by only comparing their solubility product constants? Justify your answer.
No, they cannot be compared this way, since the two ionic compounds are not composed of ions in the same proportions. Silver bromide produces one cation for one anion, while iron hydroxide produces one cation for two anions.
- b) At 25°C , the solubility product constant of silver bromide is 5.35×10^{-13} , while that of iron hydroxide is 4.87×10^{-17} . Which of the two compounds is the most soluble in water at 25°C ?

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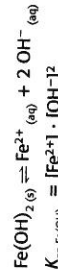
$$K_{sp, \text{AgBr}} = 5.35 \times 10^{-13}$$

$$\text{Solubility}_{\text{AgBr}} = ?$$



$$s = \sqrt{K_{sp, \text{AgBr}}}$$

$$= \sqrt{5.35 \times 10^{-13}} = 7.31 \times 10^{-7} \text{ mol/L}$$



$$s = \sqrt[3]{\frac{K_{sp, \text{Fe}(\text{OH})_2}}{4}}$$

$$= \sqrt[3]{\frac{4.87 \times 10^{-17}}{4}} = 2.30 \times 10^{-6} \text{ mol/L}$$

Answer: The compound that is the most soluble in water at 25°C is the iron hydroxide since its solubility is greater than that of silver bromide even though its solubility product constant is lower.

3. At 20°C , the solubility of mercury sulfate (Hg_2SO_4) is $0.271 \text{ g}/100 \text{ mL}$. What is the value of its solubility product constant?

Data:

$$\text{Solubility}_{\text{Hg}_2\text{SO}_4} = 0.271 \text{ g}/100 \text{ mL}$$

$$M_{\text{Hg}_2\text{SO}_4} = 497.241 \text{ g/mol}$$

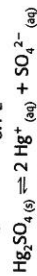
$$V = 100 \text{ mL} = 0.1 \text{ L}$$

$$K_{sp} = ?$$

Calculation:

$$n = \frac{m}{M} = \frac{0.271 \text{ g}}{497.241 \text{ g/mol}} = 5.450 \times 10^{-4} \text{ mol}$$

$$C = \frac{n}{V} = \frac{5.450 \times 10^{-4} \text{ mol}}{0.1 \text{ L}} = 5.450 \times 10^{-3} \text{ mol/L}$$



$$K_{sp} = [\text{Hg}^+]^2 \cdot [\text{SO}_4^{2-}]$$

$$= (2s)^2 \cdot s = 4s^3$$

$$= 4(5.450 \times 10^{-3})^3 = 6.476 \times 10^{-7}$$

Answer: The solubility product constant of mercury sulfate in water at 20°C is 6.5×10^{-7} .

4. Determine the mass of barium iodate ($\text{Ba}(\text{IO}_3)_2$) that must be dissolved in 300 mL of water to obtain a saturated solution of barium iodate with a solubility product constant of 4.01×10^{-9} at 25°C , if the volume of the solute is not taken into account.

Data:

$$M_{\text{Ba}(\text{IO}_3)_2} = 487.129 \text{ g/mol}$$

$$V = 300 \text{ mL} = 0.3 \text{ L}$$

$$K_{sp, \text{Ba}(\text{IO}_3)_2} = 4.01 \times 10^{-9}$$

$$m = ?$$

Calculation:



$$K_{sp} = [\text{Ba}^{2+}] \cdot [\text{IO}_3^-]^2$$

$$= s \cdot (2s)^2 = 4s^3$$

$$s = \sqrt[3]{\frac{K_{sp}}{4}}$$

$$= \sqrt[3]{\frac{4.01 \times 10^{-9}}{4}} = 1.001 \times 10^{-3} \text{ mol/L}$$

$$n = \frac{m}{M} \Rightarrow m = n \cdot M$$

$$m = 1.001 \times 10^{-3} \text{ mol} \cdot 487.129 \text{ g/mol}$$

$$= 0.4876 \text{ g}$$

$$\frac{0.4876 \text{ g}}{1.0 \text{ L}} = \frac{m}{0.3 \text{ L}}$$

$$m = \frac{0.4876 \text{ g} \cdot 0.3 \text{ L}}{1.0 \text{ L}} = 0.14628 \text{ g}$$

Answer: The mass of barium iodate at 25°C to be dissolved in 300 mL of water is 0.15 g .

12.2.4 Acidity constant and basicity constant

Textbook, p. 329 to 335

1. A chlorous acid (HClO_2) solution has a concentration of 0.2 mol/L and an acidity constant value of 1.1×10^{-2} . What is the pH of this acid solution? What is the ionization percentage of this acid at this concentration?

Data:

$[\text{HClO}_2]_i = 0.2 \text{ mol/L}$	$\text{HClO}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{ClO}_2^-(\text{aq})$
$K_a = 1.1 \times 10^{-2}$	
pH = ?	
ionization % = ?	
Concentration (mol/L)	
Initial (C_i)	0.2 - - 0 0
Change (ΔC)	-x - - +x +x
Equilibrium (C_{eq})	0.2 - x - - x x
$C_i + \Delta C = C_{\text{eq}}$	

$$x = [\text{H}_3\text{O}^+]_{\text{eq}} = [\text{ClO}_2^-]_{\text{eq}} \quad [\text{HClO}_2]_{\text{eq}} = 0.2 - x$$

$$K_a = \frac{[\text{HClO}_2]}{[\text{H}_3\text{O}^+] \cdot [\text{ClO}_2^-]} = \frac{x^2}{0.2 - x} = 1.1 \times 10^{-2}$$

$$x^2 + 0.011x - 0.0022 = 0 \quad x = -0.0527 \text{ or } 0.0417$$

The value of $x = -0.0527$ must be rejected since it is negative.

$$x = [\text{H}_3\text{O}^+]_{\text{eq}} = [\text{ClO}_2^-]_{\text{eq}} = 0.0417 \text{ mol/L}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(0.0417) = 1.3799$$

$$\text{ionization \%} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HClO}_2]_i} \cdot 100 = \frac{0.0417 \text{ mol/L}}{0.2 \text{ mol/L}} \cdot 100 = 20.85\%$$

Answer: The pH of the solution is 1.38 and its ionization percentage is 20.9%.

2. A solution of cyanic acid (HOCN) with a concentration of 0.01 mol/L has an ionization percentage of 17.04%.

- a) What is the pH of this acid solution?

Data:

$$[\text{HOCN}]_i = 0.01 \text{ mol/L} \quad \text{ionization \%} = \frac{[\text{H}_3\text{O}^+]_{\text{eq}}}{[\text{HOCN}]_i} \cdot 100$$

$$\text{ionization \%} = 17.04\%$$

$$\text{pH} = ? \quad [\text{H}_3\text{O}^+]_{\text{eq}} = \frac{\text{ionization \%} \cdot [\text{HOCN}]_i}{100} = \frac{17.04 \cdot 0.01 \text{ mol/L}}{100}$$

$$= 1.704 \times 10^{-3} \text{ mol/L}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.704 \times 10^{-3}) = 2.7685$$

Answer: The pH of the solution is 2.77.

- b) What is the value of the acidity constant of the cyanic acid?

Data:

$[\text{HOCN}]_i = 0.01 \text{ mol/L}$	$\text{HOCN}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OCN}^-(\text{aq})$
$[\text{H}_3\text{O}^+]_{\text{eq}} = 1.704 \times 10^{-3} \text{ mol/L}$	
$K_a = ?$	
Concentration (mol/L)	
Initial (C_i)	0.01 - - 0 0
Change (ΔC)	-x - - +x +x
Equilibrium (C_{eq})	0.01 - x - - x x
$C_i + \Delta C = C_{\text{eq}}$	

$$x = [\text{H}_3\text{O}^+]_{\text{eq}} = [\text{OCN}^-]_{\text{eq}} = 1.704 \times 10^{-3} \text{ mol/L}$$

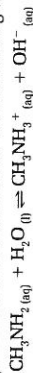
$$[\text{HOCN}]_{\text{eq}} = 0.01 - x = 0.01 \text{ mol/L} - 1.704 \times 10^{-3} \text{ mol/L} = 8.296 \times 10^{-3} \text{ mol/L}$$

$$K_a = \frac{[\text{H}_3\text{O}^+] \cdot [\text{OCN}^-]}{[\text{HOCN}]} = \frac{(1.704 \times 10^{-3})^2}{8.296 \times 10^{-3}} = 3.50 \times 10^{-4}$$

$$K_a = \frac{[\text{HOCN}]}{8.296 \times 10^{-3}} = 3.50 \times 10^{-4}$$

Answer: The value of the acidity constant of the cyanic acid is 3.5×10^{-4} .

3. Methanamine (CH_3NH_2) is a base that dissociates in water according to the following equation:



What are the initial concentration and ionization percentage of a methanamine solution with a pH of 9.16, if the value of its basicity constant is 4.60×10^{-4} ?

Data:

$$\text{pH} = 9.16$$

$$K_b = 4.60 \times 10^{-4}$$

$$K_{\text{water}} = 1.00 \times 10^{-14}$$

$$[\text{CH}_3\text{NH}_2]_i = ?$$

$$\text{ionization \%} = ?$$

Calculation:

$$[\text{H}_3\text{O}^+]_{\text{eq}} = 10^{-\text{pH}} = 10^{-9.16} = 6.9183 \times 10^{-10} \text{ mol/L}$$

$$[\text{H}_3\text{O}^+]_{\text{eq}} \cdot [\text{OH}^-]_{\text{eq}} = K_{\text{water}}$$

$$[\text{OH}^-]_{\text{eq}} = \frac{K_{\text{water}}}{[\text{H}_3\text{O}^+]_{\text{eq}}} = \frac{1.00 \times 10^{-14}}{6.9183 \times 10^{-10} \text{ mol/L}} = 1.4454 \times 10^{-5} \text{ mol/L}$$

Concentration (mol/L)	$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$
Initial (C_i)	C_i - - 0 0
Change (ΔC)	-x - - +x +x
Equilibrium (C_{eq})	$C_i - x$ - - x x
$C_i + \Delta C = C_{\text{eq}}$	

$$x = [\text{OH}^-]_{\text{eq}} = [\text{CH}_3\text{NH}_3^+]_{\text{eq}} = 1.4454 \times 10^{-5} \text{ mol/L}$$

$$[\text{CH}_3\text{NH}_2]_{\text{eq}} = C_i - x$$

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+] \cdot [\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{x^2}{C_i - x}$$

$$C_i = \frac{x^2}{K_b} + x = \frac{(1.4454 \times 10^{-5})^2}{4.60 \times 10^{-4}} + 1.4454 \times 10^{-5} = 1.4908 \times 10^{-5} \text{ mol/L}$$

$$\text{ionization \%} = \frac{[\text{OH}^-]_{\text{eq}}}{[\text{CH}_3\text{NH}_2]_i} \cdot 100 = \frac{1.4454 \times 10^{-5} \text{ mol/L}}{1.4908 \times 10^{-5} \text{ mol/L}} \cdot 100 = 96.955\%$$

Answer: The ionization percentage of methanamine is 96.96%.

Chemistry

Killer and Not-so-Difficult Ksp Problems From Various Tests

1. a) Calculate the solubility product constant(Ksp) for manganese(II) hydroxide, $\text{Mn}(\text{OH})_2$, if 100.0 mL of a saturated solution of manganese(II) hydroxide was found to contain 3.28×10^{-4} grams of manganese(II) hydroxide dissolved in it.

Manganese has a charge of +2

		$\text{Mn}(\text{OH})_2(\text{s})$ =	$\text{Mn}^{+2}(\text{aq})$	+ 2 $\text{OH}^{-}(\text{aq})$
Mole /L	I		0	0
	C	3.28×10^{-4} grams/(88.95g/mole)/0.100 L= $3.687464868 \times 10^{-5}$	$3.687464868 \times 10^{-5}$	$3.687464868 \times 10^{-5}$ *2 = $7.374929736 \times 10^{-4}$
	E		$3.687464868 \times 10^{-5}$	$7.374929736 \times 10^{-4}$

$$\begin{aligned} K_{sp} &= [\text{Mn}^{+2}] [\text{OH}^{-}]^2 \\ &= [3.687464868 \times 10^{-5}] [7.374929736 \times 10^{-4}]^2 \\ &= 2.01 \times 10^{-13} \end{aligned}$$

- b) You are concerned about disposing a saturated solution of $\text{Mn}(\text{OH})_2$, because of Mn^{+2} 's effects on wildlife and humans. What could be done to lower the concentration of Mn^{+2} ?

Raise the pH (raise concentration of OH^{-})

↳ OH^{-} will combine with Mn^{2+} and form $\text{Mn}(\text{OH})_2$.

2. The watery portion of a 100.0 ml saturated solution of what was supposed to be cadmium arsenate, $\text{Cd}_3(\text{AsO}_4)_2$ is evaporated. 0.050 grams of $\text{Cd}_3(\text{AsO}_4)_2$ (s) were left behind after evaporation. Prove that it was not really $\text{Cd}_3(\text{AsO}_4)_2$. The solubility product constant (K_{sp}) for cadmium arsenate is 2.20×10^{-33} .

		$\text{Cd}_3(\text{AsO}_4)_2$ (s)	$=$	3 Cd^{+2} (aq)	$+ 2 \text{ AsO}_4^{-3}$ (aq)
Mole /L	I			0	0
	C		x	3x	2x
	E			3x	2x

$$K_{sp} = [\text{Cd}^{+2}]^3 [\text{AsO}_4^{-3}]^2$$

$$2.20 \times 10^{-33} = [3x]^3 [2x]^2$$

$$108 x^5 = 2.20 \times 10^{-33}$$

$$x = (2.20 \times 10^{-33} / 108)^{0.2}$$

~~$$x = 1.1529 \times 10^{-7} \text{ moles/L } \text{Cd}_3(\text{AsO}_4)_2$$~~

~~$$x = 1.1529 \times 10^{-7} \text{ moles/L } \text{Cd}_3(\text{AsO}_4)_2 (615.07 \text{ g/mole})(0.100 \text{ L}) = 7.0 \times 10^{-6} \text{ g}$$~~

So there should have been only 7.0×10^{-6} g, not 0.050 grams

3. a) Calculate the solubility product constant (K_{sp}) for $PbCl_{2(s)}$, if 50.0 mL of a saturated solution of lead(II) chloride was found to contain 0.2207 g of lead(II) chloride dissolved in it.

		$PbCl_{2(s)}$	$Pb^{+2}_{(aq)}$	$+ 2 Cl^{-}_{(aq)}$
	I		0	0
Mole /L	C	$\frac{0.2207 \text{ g}}{278.11 \text{ g/mole} / 0.050 \text{ L}} = 0.01587141778$	0.01587141778	$2 * 0.01587141778 = 0.03174283556$
	E		0.01587141778	0.03174283556

$$\begin{aligned}
 K_{sp} &= [Pb^{+2}] [Cl^{-}]^2 \\
 &= [0.01587141778] [0.03174283556]^2 \\
 &= 1.60 \times 10^{-5}
 \end{aligned}$$

- b) You are concerned about disposing a saturated solution of $PbCl_2$ because of Pb^{+2} effects on wildlife and humans. What could be done to lower the concentration of Pb^{+2} ?

Add chloride

- c) What could happen in the environment that would increase the concentration of Pb^{+2} ?

precipitation of chloride by another ion; more $PbCl_2$ dissolving

4. If 55 mg of lead (II) sulfate, PbSO_4 , is placed in 250.0 mL of pure water, how much PbSO_4 will remain undissolved?
The solubility product constant (K_{sp}) for PbSO_4 is 2.53×10^{-8} .

		$\text{PbSO}_{4(s)} =$	$\text{Pb}^{+2} \text{ (aq)}$	$+ \text{SO}_4^{-2} \text{ (aq)}$
Mole /L	I		0	0
	C	x	x	x
	E		x	x

$$K_{sp} = [\text{Pb}^{+2}] [\text{SO}_4^{-2}]$$

$$2.53 \times 10^{-8} = x^2$$

$$x = 1.590597372 \times 10^{-3} \text{ mole/L}$$

$$x = 1.590597372 \times 10^{-3} \text{ mole/L} (303.26 \text{ g/mole})(0.250 \text{ L}) = 0.01205911398 \text{ g}$$

$$= 12.1 \text{ mg}$$

$$\text{Undissolved} = 55 - 12.1 \text{ mg} = 43 \text{ mg}$$