

CHEM 3.6 (5 credits)

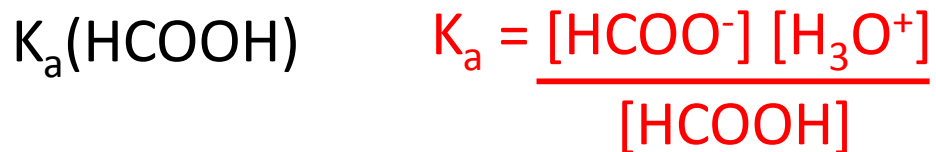
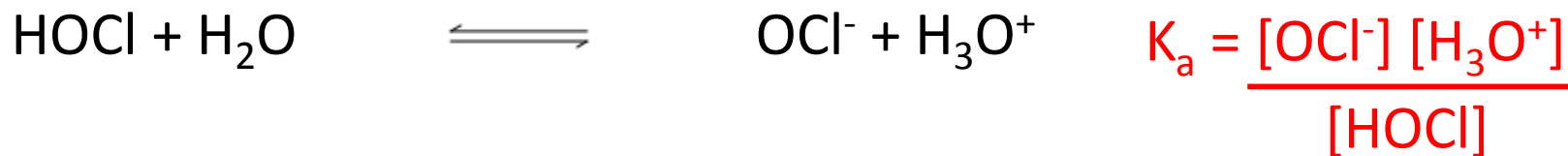
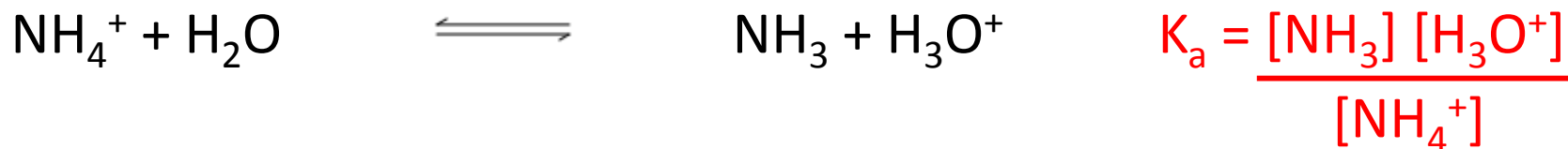
Demonstrate understanding of equilibrium principals
in aqueous systems

- sparingly soluble ionic solids
- acidic and basic solutions

- concentrations of dissolved species
- K_s calculations
- common ion effect
- predicting precipitation/dissolution
- K_a and pK_a calculations
- concentration of species present
- pH and conductivity
- titration curves and selection of indicators

Do now:

Write K_a expressions for the following acids:



Equilibrium expressions

You learnt last year in 2.6 how to write and calculate equilibrium expressions. Quick recap:

For the equation $\text{Ca}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Ca}^{2+}_{(\text{aq})} + 2 \text{NO}_3^{-}_{(\text{aq})}$

The equilibrium expression is: $K_c = \frac{[\text{Ca}^{2+}][\text{NO}_3^{-}]^2}{[\text{Ca}(\text{NO}_3)_2]}$ [products]
[reactants]

Key words

Solvent

Solute

Saturated solution

Solubility

Solubility product

Solubility constant (K_s)

K_s is a new constant this year – the solubility product or solubility constant

Just like K_c except we don't include solids (as they aren't soluble!)

For the equation $\text{Ca}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2 \text{NO}_3^{-}(\text{aq})$

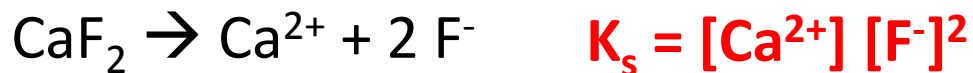
The solubility constant is: $K_s = [\text{Ca}^{2+}][\text{NO}_3^{-}]^2$

K_s is constant so long as the temperature is constant.

The greater the value of K_s the better the solubility. We are dealing with sparingly soluble ionic compounds this year K_s values will be small

Do now:

Write K_s expressions for the following ionic solids



magnesium oxide



mol.L⁻¹ and g.L⁻¹

Solubilities of ions in solution will be given to you as either mol.L⁻¹ or g.L⁻¹.

We need to be able to convert between these two units

$$\frac{\text{g}}{\text{L}} \times \frac{\text{mol}}{\text{g}} = \frac{\text{mol}}{\text{L}}$$

1/molar mass

workbook pg 192 Q3

The solubility of Mg(OH)₂ is 1.16 x 10⁻² g.L⁻¹, what is its solubility in mol.L⁻¹? **1.98 x 10⁻⁴ mol.L⁻¹**

Stronium flouride (SrF₂) has a solubility of 0.012 g per 100 mL of water. What is its solubility in mol.L⁻¹? M(SrF₂) = 125.6 g.mol⁻¹

0.120 g.L⁻¹

9.6 x 10⁻⁴ mol.L⁻¹

Workbook

pg 190 - 192

Calculations with K_s

There are two types of ionic solids – depending on their formula.

AB - NaCl, CaSO₄

A₂B/AB₂ - Na₂SO₄, Ag₂CO₃ MgCl₂, Cu(NO₃)₂

We are expected to be able to calculate K_s if given the concentration of ions for a sparingly soluble ionic salt

OR

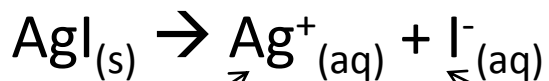
calculate the solubility (and/or concentration of ions in solution) of a sparingly soluble ionic salt if given K_s .



In both cases always write the dissolving equation and the K_s expression first

Calculating K_s from solubility

The solubility of AgI in pure water is $9.1 \times 10^{-9} \text{ mol.L}^{-1}$, calculate the solubility constant, K_s .



$$K_s = [\text{Ag}^+] [\text{I}^-]$$

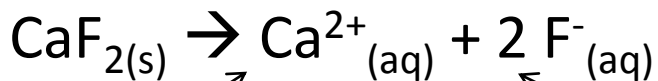
$[\text{Ag}^+] = 9.1 \times 10^{-9} \text{ mol.L}^{-1}$

$[\text{I}^-] = 9.1 \times 10^{-9} \text{ mol.L}^{-1}$

K_s – no units!

$$\begin{aligned} K_s &= (9.1 \times 10^{-9})^2 \\ &= 8.281 \times 10^{-17} \end{aligned}$$

The solubility of CaF_2 in pure water is $2.3 \times 10^{-4} \text{ mol.L}^{-1}$, calculate the solubility constant, K_s .



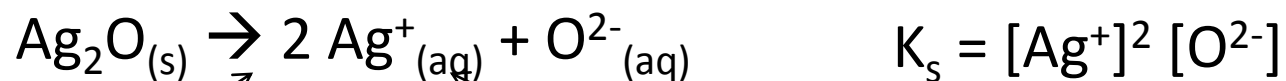
$$K_s = [\text{Ca}^{2+}] [\text{F}^-]^2$$

$[\text{Ca}^{2+}] = 2.3 \times 10^{-4} \text{ mol.L}^{-1}$ $[\text{F}^-] = 2 \times 2.3 \times 10^{-4} \text{ mol.L}^{-1} = 4.6 \times 10^{-4} \text{ mol.L}^{-1}$

$$\begin{aligned} K_s &= 2.3 \times 10^{-4} \times (4.6 \times 10^{-4})^2 \\ &= 4.87 \times 10^{-11} \end{aligned}$$

Calculating K_s from solubility

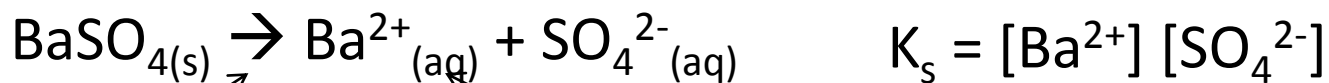
The solubility of Ag_2O in pure water is $0.00108 \text{ mol.L}^{-1}$, calculate the solubility constant, K_s .



$$[\text{O}^{2-}] = 0.00108 \text{ mol.L}^{-1}$$

$$[\text{Ag}^+] = 2 \times 0.00108 \text{ mol.L}^{-1} = 0.00216 \text{ mol.L}^{-1} \quad K_s = 0.00216^2 \times 0.00108 = 5.04 \times 10^{-9}$$

The solubility of BaSO_4 in pure water is $1.05 \times 10^{-5} \text{ mol.L}^{-1}$, calculate the solubility constant, K_s .



$$[\text{SO}_4^{2-}] = 1.05 \times 10^{-5} \text{ mol.L}^{-1}$$

$$[\text{Ba}^{2+}] = 1.05 \times 10^{-5} \text{ mol.L}^{-1}$$

$$K_s = (1.05 \times 10^{-5} \text{ mol.L}^{-1})^2 = 1.10 \times 10^{-10}$$

Do now:

- (a) Write the equation for PbI_2 dissolving in water
- (b) Write an expression for the solubility constant for PbI_2
- (c) If the solubility of PbI_2 in water is $0.00164 \text{ mol.L}^{-1}$ what is the concentration of Pb^{2+} and I^- in solution?
- (d) What is K_s for PbI_2 ?

2013 Exam Q 2

QUESTION TWO

In an experiment, a saturated solution was made by dissolving 1.44×10^{-3} g of Ag_2CrO_4 in water, and making it up to a volume of 50.0 mL.

$$M(\text{Ag}_2\text{CrO}_4) = 332 \text{ g mol}^{-1}$$

- (a) Write the K_s expression for $\text{Ag}_2\text{CrO}_4(s)$.

TWO (a)	$K_s = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]$
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- (b) (i) Calculate the solubility of $\text{Ag}_2\text{CrO}_4(s)$, and hence give the $[\text{Ag}^+]$ and $[\text{CrO}_4^{2-}]$ in the solution.

- (ii) Determine the $K_s(\text{Ag}_2\text{CrO}_4)$.

2013 Exam Q2

(b)(i)

$$n(\text{Ag}_2\text{CrO}_4) = \frac{1.44 \times 10^{-3}}{332}$$
$$= 4.33 \times 10^{-6} \text{ mol in 50 mL}$$

$$[\text{Ag}_2\text{CrO}_4] = \frac{4.33 \times 10^{-6}}{50 \times 10^{-3}}$$
$$= 8.67 \times 10^{-5} \text{ mol L}^{-1}$$

$$[\text{Ag}^+] = 8.67 \times 10^{-5} \times 2 = 1.73 \times 10^{-4} \text{ mol L}^{-1}$$

$$[\text{CrO}_4^{2-}] = 8.67 \times 10^{-5} \text{ mol L}^{-1}$$

(ii)

$$K_s = (1.73 \times 10^{-4})^2 (8.67 \times 10^{-5})$$
$$= 2.61 \times 10^{-12}$$

2013 Exam Q 2

For achieved:

- K_s correct in (a)
- Working correct OR correct answer but minimal working in (b) (i)
- Correct ratio of $[Ag^+]$ and $[CrO_4^{2-}]$ in (b) (i)
- Use of $4s^3$ in (b) (ii)

For merit:

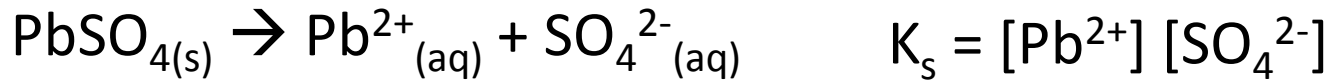
- Correct calculation of concentration of Ag_2CrO_4 in (b) (i)

For excellence:

- Correct calculation of concentration of Ag^+ ions and CrO_4^{2-} ions in (b) (i) AND K_s in (b) (ii)

Calculating solubility from K_s

The solubility constant of PbSO_4 is 1.6×10^{-8} . Calculate the solubility of PbSO_4 in water and the concentrations of the ions in solution.



If the solubility of PbSO_4 is x then we can substitute into the K_s equation

$$\begin{aligned}x^2 &= 1.6 \times 10^{-8} \\x &= \sqrt{1.6 \times 10^{-8}} \\x &= 2.82 \times 10^{-4}\end{aligned}$$

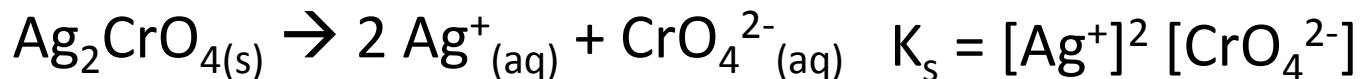
The solubility of PbSO_4 is $2.82 \times 10^{-4} \text{ mol.L}^{-1}$

The concentration of Pb^{2+} ions in solution is $2.82 \times 10^{-4} \text{ mol.L}^{-1}$

The concentration of SO_4^{2-} ions in solution is $2.82 \times 10^{-4} \text{ mol.L}^{-1}$

Calculating solubility from K_s

The solubility constant of Ag_2CrO_4 is 9.0×10^{-12} . Calculate the solubility of Ag_2CrO_4 in water and the concentrations of the ions in solution.



If the solubility of Ag_2CrO_4 is x then we can substitute into the K_s equation

$$\begin{aligned}(2x)^2 x &= K_s \\ 4x^3 &= 9.0 \times 10^{-12} \\ x &= \sqrt[3]{\frac{9.0 \times 10^{-12}}{4}} \\ x &= 1.31 \times 10^{-4}\end{aligned}$$

The solubility of Ag_2CrO_4 is $1.31 \times 10^{-4} \text{ mol.L}^{-1}$

The concentration of Ag^+ ions in solution is $2.62 \times 10^{-4} \text{ mol.L}^{-1}$

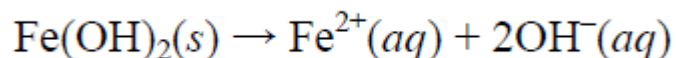
The concentration of CrO_4^{2-} ions in solution is $1.31 \times 10^{-4} \text{ mol.L}^{-1}$

2012 Exam

QUESTION TWO

Iron(II) hydroxide, $\text{Fe}(\text{OH})_2$, has a K_s of 4.10×10^{-15} at 25°C .

- (a) (i) Write the equation for $\text{Fe}(\text{OH})_2$ dissolving in water.



(states not required,
allow \rightleftharpoons)

- (ii) Write the expression for $K_s(\text{Fe}(\text{OH})_2)$.

$$K_s = [\text{Fe}^{2+}] [\text{OH}^-]^2$$

Let s be the solubility:

$$[\text{Fe}^{2+}] = s$$

$$[\text{OH}^-] = 2s$$

$$K_s = s \times (2s)^2$$

y (in mol L^{-1}) of iron(II) hydroxide in water at 25°C .

$$4.10 \times 10^{-15} = 4s^3 \quad s = 1.01 \times 10^{-5} \text{ mol L}^{-1}$$

$$\text{Solubility of } \text{Fe}(\text{OH})_2(s) = 1.01 \times 10^{-5} \text{ mol L}^{-1}$$

2014 Exam

QUESTION TWO

A flask contains a saturated solution of PbCl_2 in the presence of undissolved PbCl_2 .

- (a) (i) Write the equation for the dissolving equilibrium in a saturated solution of PbCl_2 .



- (ii) Write the expression for $K_s(\text{PbCl}_2)$.

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

$$[\text{Pb}^{2+}] = x \quad [\text{Cl}^-] = 2x$$

$$K_s = 4x^3$$

$$x = \sqrt[3]{\frac{K_s}{4}}$$

$$[\text{Pb}^{2+}] = 1.62 \times 10^{-2} \text{ mol L}^{-1}$$

$$[\text{Cl}^-] = 3.24 \times 10^{-2} \text{ mol L}^{-1}$$

- (iii) Calculate the solubility (in mol L^{-1}) of Pb^{2+} and Cl^- in the solution.

$$K_s(\text{PbCl}_2) = 1.70 \times 10^{-5} \text{ at } 25^\circ\text{C}$$

$$= \sqrt[3]{\frac{1.70 \times 10^{-5}}{4}}$$

$$= 1.62 \times 10^{-2} \text{ mol L}^{-1}$$

and give the

Do now:

(a) Write the equation for PbBr_2 dissolving in water

(b) Write an expression for the solubility product of PbBr_2 dissolving in water

(c) If the solubility of PbBr_2 in water at room temperature is $2.65 \times 10^{-2} \text{ mol.L}^{-1}$ what is the concentration of Pb^{2+} and Br^- in solution?

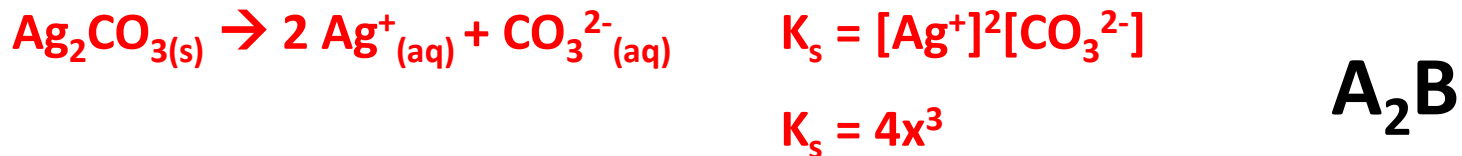
(d) How would you calculate the solubility of PbBr_2 if $K_s = 7.62 \times 10^{-6}$ at 0°C ?

Do now:

- a) Write the equation for CuCO_3 dissolving and the K_s expression.
b) If x = solubility, substitute in to the K_s expression to solve for x



- c) Write the equation for Ag_2CO_3 dissolving and the K_s expression.
d) If x = solubility, substitute into the K_s expression to solve for x



- e) Write the equation for CuI_2 dissolving and the K_s expression.
f) If x = solubility, substitute into the K_s expression to solve for x

