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:

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $\quad K_a = \dfrac{[NH_3]\,[H_3O^+]}{[NH_4^+]}$

$HOCl + H_2O \rightleftharpoons OCl^- + H_3O^+$ $\quad K_a = \dfrac{[OCl^-]\,[H_3O^+]}{[HOCl]}$

$K_a(HCOOH)$ $\quad K_a = \dfrac{[HCOO^-]\,[H_3O^+]}{[HCOOH]}$

Equilibrium expressions

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

For the equation $Ca(NO_3)_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2\ NO_3^{-}{}_{(aq)}$

The equilibrium expression is: $K_c = \dfrac{[Ca^{2+}][NO_3^{-}]^2}{[Ca(NO_3)_2]}$ $\qquad \dfrac{\textbf{[products]}}{\textbf{[reactants]}}$

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 $Ca(NO_3)_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)}$

The solubility constant is: $K_s = [Ca^{2+}][NO_3^-]^2$

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

$MgS \rightarrow Mg^{2+} + S^{2-}$ **$K_s = [Mg^{2+}] [S^{2-}]$**

$CaF_2 \rightarrow Ca^{2+} + 2 F^-$ **$K_s = [Ca^{2+}] [F^-]^2$**

$Ba(OH)_2$ **$Ba(OH)_2 \rightarrow Ba^{2+} + 2 OH^-$** **$K_s = [Ba^{2+}] [OH^-]^2$**

magnesium oxide

$MgO \rightarrow Mg^{2+} + O^{2-}$ **$K_s = [Mg^{2+}] [O^{2-}]$**

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

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

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 (SrF_2)has a solubility of 0.012 g per 100 mL of water. What is its solubility in mol.L^{-1}? M(SrF_2) = 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 - $NaCl$, $CaSO_4$

A_2B/AB_2 - Na_2SO_4, Ag_2CO_3 $MgCl_2$, $Cu(NO_3)_2$

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.

$A_xB_y \rightarrow x\,A + y\,B$ $K_s = [A]^x[B]^y$

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 9.1×10^{-9} mol.L^{-1}, calculate the solubility constant, K_s.

$$AgI_{(s)} \rightarrow Ag^{+}_{(aq)} + I^{-}_{(aq)}$$

$$K_s = [Ag^+]\,[I^-]$$

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

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

K_s – no units!

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

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

$$CaF_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2\,F^{-}_{(aq)}$$

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

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

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

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

Calculating K_s from solubility

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

$Ag_2O_{(s)} \rightarrow 2\ Ag^+_{(aq)} + O^{2-}_{(aq)}$ $\quad K_s = [Ag^+]^2\ [O^{2-}]$

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

$[Ag^+] = 2 \times 0.00108$ mol.L^{-1} $= 0.00216$ mol.L^{-1}

$K_s = 0.00216^2 \times 0.00108$
$= 5.04 \times 10^{-9}$

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

$BaSO_{4(s)} \rightarrow Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$ $\quad K_s = [Ba^{2+}]\ [SO_4^{2-}]$

$[SO_4^{2-}] = 1.05 \times 10^{-5}$ mol.L^{-1}

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

$K_s = (1.05 \times 10^{-5}$ 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 $mol.L^{-1}$ 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 (Ag_2CrO_4) = 332 g mol^{-1}

(a) Write the K_s expression for $Ag_2CrO_4(s)$.

TWO (a)	$K_s = [Ag^+]^2[CrO_4^{2-}]$

(b) (i) Calculate the solubility of $Ag_2CrO_4(s)$, and hence give the $[Ag^+]$ and $[CrO_4^{2-}]$ in the solution.

(ii) Determine the $K_s(Ag_2CrO_4)$.

2013 Exam Q2

(b)(i)

$$n(Ag_2CrO_4) = \frac{1.44\times10^{-3}}{332}$$

$$= 4.33\times10^{-6} \text{ mol in 50 mL}$$

$$[Ag_2CrO_4] = \frac{4.33\times10^{-6}}{50\times10^{-3}}$$

$$= 8.67\times10^{-5} \text{ mol L}^{-1}$$

$$[Ag^+] = 8.67\times10^{-5}\times2 = 1.73\times10^{-4} \text{ mol L}^{-1}$$

$$[CrO_4^{\ 2-}] = 8.67\times10^{-5} \text{ mol L}^{-1}$$

(ii)

$$K_s = (1.73\times10^{-4})^2\,(8.67\times10^{-5})$$

$$= 2.61\times10^{-12}$$

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.

$PbSO_{4(s)} \rightarrow Pb^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$ $\quad K_s = [Pb^{2+}]\,[SO_4^{2-}]$

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

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

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

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

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

Calculating solubility from K_s

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

$$Ag_2CrO_{4(s)} \rightarrow 2\ Ag^+_{(aq)} + CrO_4^{2-}{}_{(aq)} \quad K_s = [Ag^+]^2\ [CrO_4^{2-}]$$

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

$$(2x)^2\ x = K_s$$

$$4x^3 = 9.0 \times 10^{-12}$$

$$x = \sqrt[3]{\left(\frac{9.0 \times 10^{-12}}{4}\right)}$$

$$x = 1.31 \times 10^{-4}$$

The solubility of Ag_2CrO_4 is 1.31 x 10^{-4} mol.L^{-1}

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

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

2012 Exam

QUESTION TWO

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

(a) (i) Write the equation for $Fe(OH)_2$ dissolving in water.

$Fe(OH)_2(s) \rightarrow Fe^{2+}(aq) + 2OH^-(aq)$ (states not required, allow $\rightleftharpoons$)

(ii) Write the expression for $K_s(Fe(OH)_2)$.

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

Let s be the solubility: y (in mol L^{-1}) of iron(II) hydroxide in water at 25°C.

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

$[OH^-] = 2s$

$K_s = s \times (2s)^2$ $\quad 4.10 \times 10^{-15} = 4s^3 \quad s = 1.01 \times 10^{-5}$ mol L^{-1}

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

2014 Exam

QUESTION TWO

A flask contains a saturated solution of $PbC1_2$ in the presence of undissolved $PbC1_2$.

(a) (i) Write the equation for the dissolving equilibrium in a saturated solution of $PbC1_2$.

$$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$$

(ii) Write the expression for $K_s(PbC1_2)$.

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

(iii) Calculate the solubility (in mol L^{-1}) of le… and give the $[Pb^{2+}]$ and $[Cl^-]$ in the solution.

$K_s(PbC1_2) = 1.70 \times 10^{-5}$ at 25°C

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

$K_s = 4x^3$

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

$$= \sqrt[3]{\frac{1.70\times10^{-5}}{4}}$$

$$= 1.62\times10^{-2} \text{ mol L}^{-1}$$

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

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

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 x 10^{-2} 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 x 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

$CuCO_{3(s)} \rightarrow Cu^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$

$K_s = [Cu^{2+}][CO_3^{2-}]$

$K_s = x^2$

AB

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

$Ag_2CO_{3(s)} \rightarrow 2\ Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

$K_s = [Ag^+]^2[CO_3^{2-}]$

$K_s = 4x^3$

A_2B

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

$CuI_{2(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ I^-_{(aq)}$

$K_s = [Cu^{2+}][I^-]^2$

$K_s = 4x^3$

AB_2