Solubility Equilibrium

Section 7.6

Homework
Solubility and K_{sp}
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Types of Equilibrium

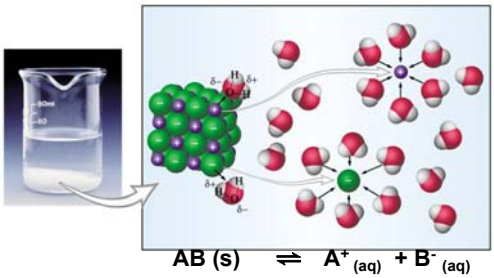
Chemical:
Equilibrium between forward and reverse reactions

Solubility:
Equilibrium between dissolving & crystallizing

Phase:
Equilibrium between changes of state

The dissolving process is the result of intermolecular attraction:

Polar ends of H_2O ------ Oppositely-charged ions



$AB_{(s)} \rightleftharpoons A^+_{(aq)} + B^-_{(aq)}$

Example. Dissolution equations:

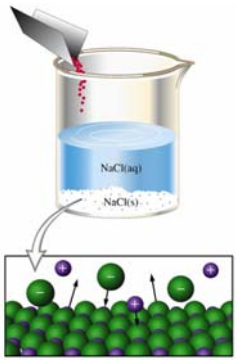
Pay attention:
- ionic charges
- molar ratios

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

b) $AlCl_{3\,(s)} \rightarrow Al^{3+}_{(aq)} + 3\,Cl^-_{(aq)}$

c) $Mg(OH)_{2\,(s)} \rightarrow Mg^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

Dissolution is a reversible process.



$AgI_{(s)} \underset{PRECIPITATION}{\overset{DISSOLUTION}{\rightleftharpoons}} Ag^+_{(aq)} + I^-_{(aq)}$

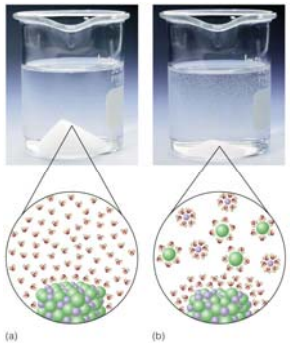
SOLID — DISSOLVED IONS

At equilibrium, the solution is **SATURATED**.

Contains the maximum amount of solute that it can dissolve

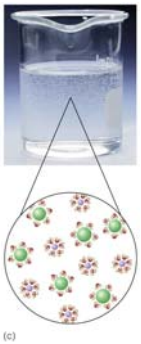
Unsaturated Solution

$rate_{dissolving} > rate_{precipitation}$



Saturated Solution

$rate_{dissolving} = rate_{precipitation}$



If you try to add more solute to a saturated solution,

$rate_{precipitation}$ ↑

∴ a solid precipitate will form.

Ever add too much hot chocolate mix to your water?

HIGH solubility
Lots of solute can dissolve

When saturated, equilibrium lies far right

LOW solubility
Very little solute can dissolve

When saturated, equilibrium lies far left

The saturation point depends on the identity of the ionic compound.

Anions	Cations: high solubility ≥ 0.1 mol/L at SATP	Cations: low solubility < 0.1 mol/L at SATP
F^-	most	Li^+, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Fe^{2+}, Hg_2^{2+}, Pb^{2+}
Cl^-, Br^-, I^-	most	Ag^+, Pb^{2+}, Tl^+, Hg_2^{2+}, Hg^+, Cu^+
S^{2-}	Group 1, Group 2, NH_4^+	most
OH^-	Group 1, NH_4^+, Sr^{2+}, Ba^{2+}, Tl^+	most
SO_4^{2-}	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}
CO_3^{2-}, PO_4^{3-}, SO_3^{2-}	Group 1, NH_4^+	most
$C_2H_3O_2^-$	most	Ag^+
NO_3^-	all	none
IO_3^-	NH_4^+, K^+, Na^+	most

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The solubility of a compound can be expressed quantitatively.

Mass solubility:

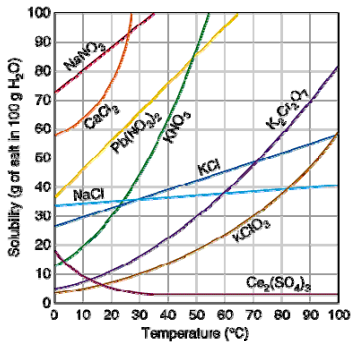
$$\frac{\text{maximum } \textbf{mass} \text{ of solute (g)}}{\text{100 mL of solution (100 mL)}} \qquad \frac{45.8 \text{ g } AlCl_3}{\text{100 mL of solution}}$$

Molar solubility:

$$\frac{\text{maximum } \textbf{moles} \text{ of solute (mol)}}{\text{L of solution (L)}} \qquad \frac{3.43 \text{ mol } AlCl_3}{\text{L of solution}}$$

…is the molar concentration of the solution when saturated

Solubility is temperature dependent!



Learning Checkpoint

- For a solid solute in an aqueous solution, the processes of dissolution and precipitation are occurring simultaneously.
  - Solid solute ⇌ Aqueous ions
- When $rate_{dissolution} = rate_{precipitation}$, the system is at equilibrium. At this point, the solution is saturated.
  - Once the solution is saturated, the addition of more ions will shift the equilibrium to the LEFT. This causes a precipitate to form.

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

Since dissolution is an equilibrium process, we can write an equilibrium law expression:

$$K_{sp} = [Ag^+]\ [I^-]$$

Where did [AgI] go??

K_{sp}: "Solubility Product Constant": The equilibrium constant for a dissolution process

Example. K_{sp} expressions

Dissolution Equation	K_{sp} expression
$AlCl_{3\,(s)} \rightleftharpoons Al^{3+}{}_{(aq)} + 3\,Cl^{-}{}_{(aq)}$	$K_{sp} = [Al^{3+}]\,[Cl^{-}]^3$
$Mg(OH)_{2\,(s)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2\,OH^{-}{}_{(aq)}$	$K_{sp} = [Mg^{2+}]\,[OH^{-}]^2$

Where are the exponents coming from?

Calculating K_{sp}

Example 1. *From ion concentrations*

$$BaF_{2\,(s)} \rightleftharpoons Ba^{2+}{}_{(aq)} + 2\,F^{-}{}_{(aq)}$$

GIVEN:
At equilibrium, -- 9.13×10^{-3} M 1.83×10^{-2} M

Strategy
1. Write K_{sp} expression
2. Sub in values and solve!

Solution

$K_{sp} = [Ba^{2+}]\,[F^-]^2$

$= (9.13 \times 10^{-3})\,(1.83 \times 10^{-2})^2$

$K_{sp} = 3.06 \times 10^{-6}$

Example 2. *Calculating K_{sp} from solubility value*

The solubility of MgF_2 at 20°C is 0.00172 g/100 mL. Determine the K_{sp} of MgF_2.

$$MgF_{2\,(s)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2\,F^{-}{}_{(aq)}$$

Strategy
1. Convert mass solubility to a molar solubility.
2. Use mole ratios to find concentrations of each ion.
3. Write K_{sp} expression
4. Sub in values and solve!

$$MgF_{2\,(s)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2\,F^{-}{}_{(aq)}$$

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Strategy
1. Convert mass solubility to a molar solubility.
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3. Write K_{sp} expression
4. Sub in values and solve!

❶ $[MgF_2] = \frac{0.00172\ g}{100\ mL} \times \frac{1\ mol}{62.31\ g} \times \frac{1000\ mL}{1\ L} = 2.76 \times 10^{-4}$ mol/L

From mole ratios...

❷ $[Mg^{2+}] = [\,MgF_2] = 2.76 \times 10^{-4}$ mol/L

$[F^-] = 2\,[\,MgF_2] = 2(2.76 \times 10^{-4}\ mol/L) = 5.52 \times 10^{-4}$ mol/L

❸ $K_{sp} = [Mg^{2+}]\,[F^-]^2 = (2.76 \times 10^{-4})\,(5.52 \times 10^{-4})^2$

$K_{sp} = 8.35 \times 10^{-11}$

You can find a table of reference K_{sp} values on **page 725** (Appendix B4)

B4 Solubility Product Constants (K_{sp})

Table 1 Solubility Product Constants at 25 °C

Name	Formula	K_{sp}
barium carbonate	$BaCO_3$(s)	2.6×10^{-9}
barium chromate	$BaCrO_4$(s)	1.2×10^{-10}
barium sulfate	$BaSO_4$(s)	1.1×10^{-10}
calcium carbonate	$CaCO_3$(s)	5.0×10^{-9}
calcium hydroxide	$Ca(OH)_2$	5.5×10^{-6}
calcium oxalate	CaC_2O_4(s); $CaOOCCOO$(s)	2.3×10^{-9}
calcium phosphate	$Ca_3(PO_4)_2$(s)	2.1×10^{-33}
calcium sulfate	$CaSO_4$(s)	7.1×10^{-5}
copper(I) chloride	$CuCl$(s)	1.7×10^{-7}
copper(I) iodide	CuI(s)	1.3×10^{-12}
copper(II) iodate	$Cu(IO_3)_2$(s)	6.9×10^{-8}
copper(II) sulfide	CuS(s)	6.0×10^{-37}
iron(II) hydroxide	$Fe(OH)_2$(s)	4.9×10^{-17}

Example 3. *Calculating solubility from K_{sp}*

Calculate the molar solubility of zinc hydroxide at 25°C, where $K_{sp} = 7.7 \times 10^{-17}$.

Strategy
1. Write dissociation equation.
2. Set up K_{sp} expression.
3. Sub in values and solve!

❶ $Zn(OH)_{2\,(s)} \rightleftharpoons Zn^{2+}{}_{(aq)} + 2\,OH^{-}{}_{(aq)}$

	$Zn(OH)_{2\,(s)}$ ⇌	$Zn^{2+}{}_{(aq)}$	$2\,OH^{-}{}_{(aq)}$
I	-	0	0
C	-	+ x	+ 2x
E	-	x	2x

x represents the mol/L of Zn^{2+} from the dissolving of $Zn(OH)_2$; In other words, the molar solubility of $Zn(OH)_2$

❷ $K_{sp} = [Zn^{2+}]\,[OH^-]^2$ *sub in* ❸ $7.7 \times 10^{-17} = (x)(2x)^2$

$7.7 \times 10^{-17} = 4x^3$

$\sqrt[3]{\frac{7.7 \times 10^{-17}}{4}} = x = 2.7 \times 10^{-6}$ mol/L

Learning Checkpoint

- The equilibrium constant for a dissolution equation has a special name: K_{sp} – The Solubility Product Constant
  - calculated the same way as a "regular" K
  - since solids aren't included in K expressions, ends up as $K_{sp} = [ion^+]^m [ion^-]^n$
- To calculate the value of K_{sp}, use ion concentrations of the saturated solution.
  - since saturated solution is at solubility equilibrium

Homework

Solubility and K_{sp}

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