Solubility Equilibria
Ionic Equation

- IA cations and the ammonium ion combined with any anion make soluble salts.
- Salts of nitrate and acetate are soluble

ex

$NaCl_{(aq)} + AgNO_{3(aq)} \rightleftharpoons AgCl_{(s)} + NaNO_{3(aq)}$

total ionic equation

$\underline{Na^+_{(aq)}} + Cl^-_{(aq)} + Ag^+_{(aq)} + \underline{NO_{3(aq)}^-} \rightleftharpoons AgCl_{(s)} + \underline{Na^+_{(aq)}} + \underline{NO_{3(aq)}^-}$

spectator ions

Net ionic equation

$Cl^-_{(aq)} + Ag^+_{(aq)} \rightleftharpoons AgCl_{(s)}$

- all to Cl, Br, I are soluble except those of Ag, Pb and Hg

$BaCl + KSO_4 \longrightarrow KCl + BaSO_4$

Saturated Solutions

- mixture in which max amt. of solute dissolved
- @ equilibrium
  - rate of dissolution & rate of reforming are =
- characterize by solute concentration
- ionic solutes → the product of the ion concentration

Two ways to characterize the solubility

1
- Solute that is generally soluble = $\frac{\text{total mass of sol.}}{\text{amt of solvent}}$
  - 2000g/L sucrose in water
- molar solubility (moles/L)

2
- ion product
- used for "insoluble" substances
- characterize the small concentrations of ions in a solution.

Ion Product

- product of the concentration of the ions that is found to be a fixed value
- value at the saturation conditions (equilibrium) is the solubility product (K_{sp})
- K_{sp}
  - constant specific to the compound &
  - temp dependent
  - aqueous solution

$K_{sp} = [Ag^+][Cl^-]$ for AgCl

$K_{sp} = [Pb^{2+}][Cl^-][Cl^-] = [Pb^{2+}][Cl^-]^2$ for $PbCl_2$

- use for not very soluble compounds otherwise its # of moles or grams that will dissolve

From Ksp to Solubility

- tells how many moles of something will dissolve

Ex. AgCl : 1 mol of Ag - 1 mol of Cl

$$K_{sp} = [Ag^+][Cl^-] = (x)(x) = x^2$$ molar solubility

$PbCl_2$: 1 mol of Pb - 2 mol of Cl

$$K_{sp} = [Pb^{2+}][Cl^-]^2 = (x)(2x)^2 = 4x^3$$

Ag_2CO_3 : 1 mol of CO_3 - 2 mol of Ag

$K_{sp} = 32 mg/L$ $\quad x = 1.16 \times 10^{-4}$ moles

$$K_{sp} = [Ag^+]^2[CO_3^{2-}] = (2.32 \times 10^{-4})^2(1.16 \times 10^{-4}) = 6.2 \times 10^{-12}$$

- molar solubility is $\sqrt[n]{\ }$ of Ksp $\sqrt[n]{K_{sp}}$ n = # of ions formed

Calc. Molar Solubility

CuBr is ~~0.5M~~ molar solubility

8.4e-8 molar

NaBr = 0.5M

$K_{sp} = 4.2 \times 10^{-8}$

CuBr

Na^+ Br^-

1:1 ration

[.5] of Br

r	$CuBr_{(s)} \rightarrow$	$Cu^+_{(aq)}$ +	$Br^-_{(aq)}$
i		∅	.5M
c		x	x
e		x	.5+x

$$4.2e{-8} = [Cu^+][Br^-]$$

$$x = 8.4e{-8} = [Cu^+]$$

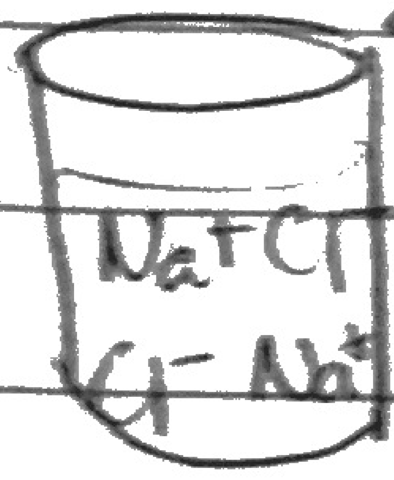
$$4.2e{-8} = (x)(.5+x)$$

$$4.2e{-8} = .5x$$

Common Ion Effect

- existence of ions affects solubility

← add AgCl



its not as soluble because Cl^- already in it

- Cl^- is a common ion
- possible to manipulate solubility by changing concentration of common ions

Precipitation

- when a solid solute spontaneously form from a solution

$$BaCl_{2(aq)} + 2NaNO_{3(aq)} \rightleftharpoons Ba(NO_3)_{2(s)} + 2NaCl_{(aq)}$$

$$Ba^{2+} + 2Cl^- + 2Na^+ + 2NO_3^- \rightleftharpoons Ba(NO_3)_{2(s)} + 2Na^+ + 2Cl^-$$

$$Ba^{2+} + 2NO_3^- \rightleftharpoons Ba(NO_3)_2$$

- only happen @ certain concentrations
- ion product - if a system is in equilibrium or not
- @ equilibrium ion product (Q) is Ksp (solubility product)

↑ concentrations $Q_{sp} > K_{sp}$ then precipitate will form

↓ " $Q_{sp} < K_{sp}$ then no " " "

$Q_{sp} = K_{sp}$ @ equilibrium

Super-Saturated Solutions - solution w/ a concentration higher than the saturation concentration