

Solubility calculations

ESS 312 - May 10, 2010

In class we considered a reaction at equilibrium:



and derived the simple relationship between the equilibrium constant and the standard-state Gibbs Free Energy difference between products and reactants:

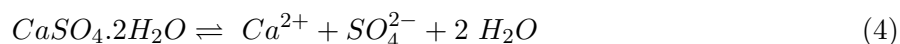
$$\ln\left(\frac{a_C^c \cdot a_D^d}{a_A^a \cdot a_B^b}\right) = \ln(K_{eq}) = \frac{-\Delta G^0}{RT} \quad (2)$$

or equivalently:

$$K_{eq} = e^{-\left(\frac{\Delta G^0}{RT}\right)} \quad (3)$$

For a dissolution reaction, this remarkable result relates the *ion activity product* (IAP) in a saturated solution to the measurable salt concentrations, and to the Gibbs Free Energy of formation ΔG_f^0 of the substances in solution.

For example, for a saturated solution of gypsum ($CaSO_4 \cdot 2H_2O$), the dissolution reaction is:



To calculate the equilibrium constant at 298K and 1 bar use the following data:

substance	ΔG_f^0 kJ/mol
$CaSO_4 \cdot 2H_2O$	-1796.4
H_2O	-237.2
Ca^{2+}	-553.8
SO_4^{2-}	-742.0

Now add these up to get ΔG^0 for the reaction:

$$\Delta G^0 = -742.0 - 553.8 - 2(237.2) - (-1796.4) = 26.3 \text{ kJ/mol} \quad (5)$$

Remembering that $R = 0.008314 \text{ kJ}/(K \text{ mol})$:

$$K_{eq} = e^{-\left(\frac{\Delta G^0}{RT}\right)} = e^{-\frac{26.3}{0.008314 \times 298.15}} = e^{-10.61} = 2.47 \times 10^{-5} = 10^{-4.61} \quad (6)$$

So ... now you know how to calculate a solubility product from tables of ΔG_f^0 .