Reading: Faure 10.1 to 10.3

Goals:
- Calculate supersaturation/undersaturation of salts.
- Consider how addition of salts that are made of weak acids plus strong bases affects pH.
- Introduce a common equation for estimating activity coefficients.

1. Solubility of Salts (Faure 10.1)
   a. In the mass action expression, we set the activity of solids equal to unity.
   Why?
      i. Alternatively: Why doesn’t the AMOUNT of solid matter?
      ii. At first glance- the surface area might seem to be relevant.
      iii. But common sense tells us that adding more crystals to an undersaturated solution only speeds up the approach to equilibrium, it will not change the concentration at which the crystals stop dissolving.
      iv. Surface area is NOT important… it cancels out because it affects forward and reverse reaction rates equally
   b. Review terminology: Saturated/Undersaturated/Supersaturated
   c. Saturation index, log Q/K (Q = IAP) is convenient for expressing the saturation state of a solution
   d. Precipitation of a metastable phase.
      i. Sometimes, the precipitation rate is extremely slow for the most stable phase and much faster for a metastable one.
      ii. Concentrations may build up beyond the saturation/equilibrium point for the first phase and continue increasing until the faster one precipitates.
      iii. Examples: Gypsum vs. Anhydrite; Quartz and Amorphous silica, Hematite vs. amorphous HFO (hydrous ferric oxide – Fe(OH)_3•nH_2O
   e. The common ion effect: Example: Gypsum and Barite together.
      i. Barite is extremely insoluble (very small K_{sp}, precipitates even when sulfate and Ba concentrations are quite low)
      ii. This keeps sulfate concentration low
      iii. This in turn causes gypsum to dissolve.
      iv. So, because gypsum and barite have a common ion (sulfate), they compete for sulfate and, because barite competes more strongly, gypsum will dissolve, and barite will precipitate, until the Ba^{2+} in solution is depleted to a very low concentration.

2. Hydrolysis: What happens when we put ions into water and those ions grab H^+ or OH^- ?(Faure 10.2)
   a. FeCl_3 in water:
      i. Fe(III) grabs OH^- ions (Cl^- does not)
ii. Drives pH down
b. Na$_3$PO$_4$ solid dissolved in water:
   i. PO$_4^{3-}$ grabs H$^+$, drives pH up. (Na$^+$ has no effect)
c. What other salts also do this?
   i. Mixtures of weak acids and strong bases (most silicate minerals).
      1. when these dissolve, H$^+$ consumed, pH tends to increase
   ii. Mixtures of strong acids and weak bases
      1. when these dissolve, OH$^-$ consumed, pH tends to decrease

3. Solubility of neutral gas molecules in water (or magma)
a. Example: $N_2(g) \rightleftharpoons N_2(aq)$, $K = (N_2(aq))/p_{N_2}$
b. Equilibrium constant is known as a “henry’s law” constant
c. We often talk about pCO$_2$ in water; think of this as the partial pressure of CO$_2$ in a gas in equilibrium with the solution
   i. Recall: Partial pressure means the contribution of the one gas to the total pressure, e.g., CO$_2$ is only about 0.035% of the atmosphere (by volume), so pCO$_2$ in water at the surface of a lake $\approx 3.5 \times 10^{-4}$ atm.
d. Igneous petrologists often use $f_{O_2}$ (oxygen fugacity) to describe the oxidation state of magmas, this is quite similar to p$_{O_2}$. 