Using Kitchen Chemistry and the Concept of Saturation to Help Students Understand Igneous Phase Diagrams

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An ability to read and use simple igneous phase diagrams is a major goal of most undergraduate courses in petrology. Many students have difficulty attaining this goal because phase diagrams are an unfamiliar kind of graph and they are described in most textbooks with an unfamiliar language – liquidus, solidus, eutectic, peritectic, etc. I have found that students respond well to class activities that use concepts, materials, and terminology from their life or educational experiences outside of geology. These activities let them apply knowledge that they already have to help them understand new ideas. Particularly useful in this regard is the concept of the saturation of a liquid with a dissolving phase, such as sugar in coffee. Indeed, I have yet to meet a college student who is not familiar with the term "saturation" and its meaning, at least in terms of sugar or salt in water.

In the following pages I have collected some phase diagrams that are helpful for discussions, problems, and other activities involving "kitchen chemistry." They can be used in many ways, depending on your class setting and the time available. Along with each diagram, I include some suggestions of how they might be used based on my own teaching experiences. Each year in my petrology course, I try to find new activities or variations to keep it interesting for me and, therefore, for the students. I encourage you to do the same and hope that you will send me new activities and kitchen phase diagrams that you discover.

1. Ice liquidus.

I described some activities using the H_2O-NaCl phase diagram in a previous article (Brady, 1992). See also Petersen (1997) and Darling (2000). The H_2O-NaCl system is particularly useful for many reasons, most importantly because it is safe to work with, reaction kinetics are fast even at low temperatures, and students have considerable experience with it. Students all know that halite can be dissolved in water, making the water saltier, until the water is saturated with halite. They also know that more halite can be dissolved in hot water than cold. However, they don't know that salty water can be saturated with ice. I use the example of salty puddles on sidewalks after using salt to melt ice in the winter. I ask them what happens to salt or ice (snowflakes) added to those puddles. Students will say that halite will "dissolve" in (unsaturated) salty water at –5°C, but that ice will "melt" in the same solution. To convince them that ice dissolves, we determine the ice saturation curve as a class activity. This can be done in a qualitative way or more quantitatively as a complete lab exercise.

For the qualitative activity, students mix solutions of different salinities, add ice in excess (saturating the solution), and measure the temperature. Saltier solutions will be saturated at lower temperatures. For the more quantitative activity, students decant some of the ice-saturated solution into pre-weighed beakers or plastic cups and weigh again. The contents of

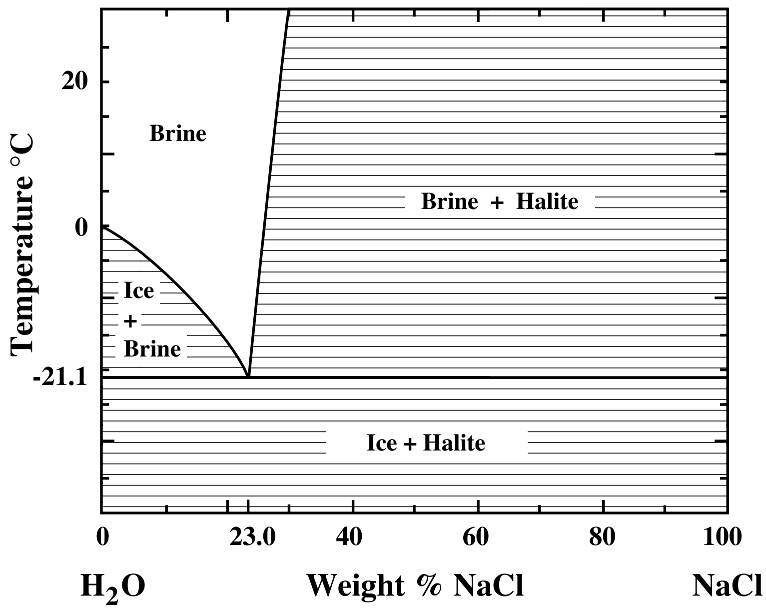

Figure 1. Saturation diagram for the system H_2O-NaCl modified to ignore the presence of hydrohalite (NaCl.2H_2O).

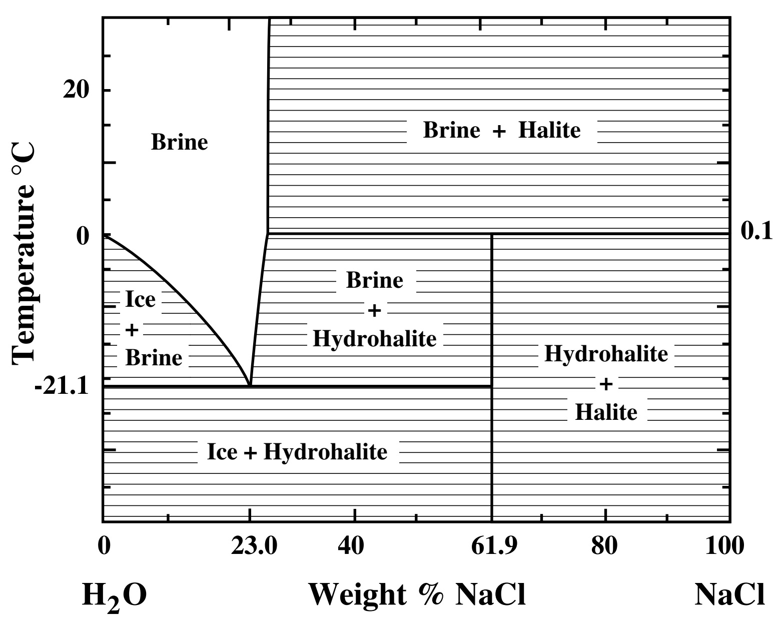

Figure 2. Saturation diagram for the system H2O-NaCl with the incongruent melting behavior of hydrohalite (NaCl·2H_2O) shown. Data taken from Stephens and Stephen (1963).

the cups are allowed to evaporate in a hood or a warming oven overnight or over the next few days. When only salt remains in the cups, they are weighed again and the weight percent NaCl for each saturated solution is calculated. These data, along with the measured temperatures are used to construct the ice-saturated liquidus on an H_2O-NaCl phase diagram. Other halides systems can be used instead of NaCl for variety or so that different groups have different projects. KCl works quite well. $MgCl_2$ and $CaCl_2$ seem to be more difficult to work with, but can also be used. I often ask the students to find saturation data in the library for comparison with their experimental results.

I have included here the H_2O-NaCl phase diagram fit to published saturation data. Because this and most halide-water systems have a halide-hydrate that is stable only at low temperatures, each has an incongruently melting phase. To simplify the initial discussion, I commonly start with Figure 1, which ignores the presence of hydrohalite ($NaCl{\cdot}2H_2O$). The more accurate version is given as Figure 2.

2. Partial melting of sucrose-ice rocks (Popsicles®)

Food is always a good draw for college students and I rarely get complaints when I distribute Popsicles® in class for partial melting experiments (no biting or chewing). In principle, the extracted liquids will all have the eutectic sweetness until the sucrose is consumed and only ice remains. Then the liquids should taste a lot like water (see Figure 3). Homemade "quiescently frozen confections" using fruit juice work best. The commercial varieties are rarely just sugar and water, so read the label before you buy. Do not attempt saturation experiments with sugar and water like those described for halite, because reaction rates are slow and sugar syrup is very messy.

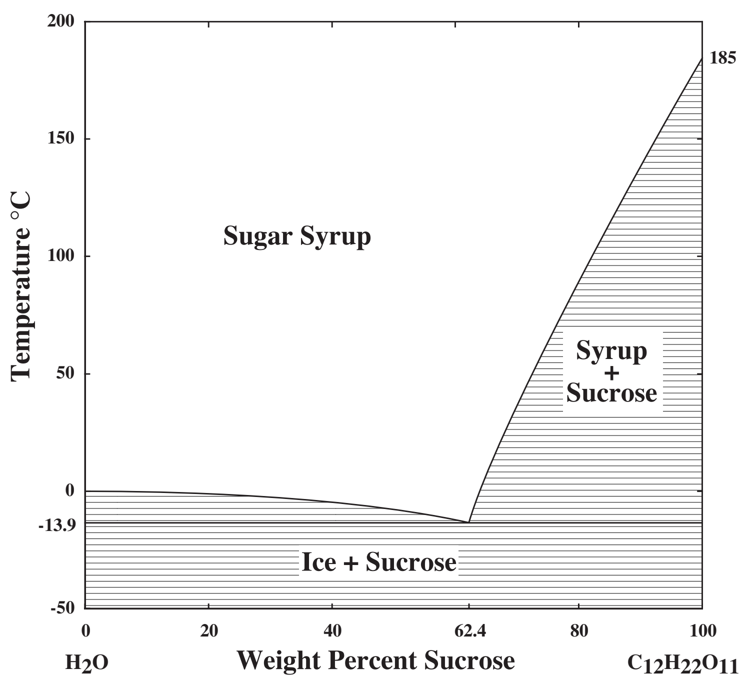

Figure 3. Water-sucrose saturation diagram drawn from data reported by Bates et al. (1928).

3. Freezing of hard cider

Unpasteurized apple cider will ferment due to yeast on the apples until the concentration of alcohol in the cider becomes so high that the yeast are killed. This will occur when the alcohol content reaches 12% to 14% by weight. However, a barrel of fermented cider left outdoors in New England in the winter will undergo fractional crystallization. Pure (stoichiometric) ice will crystallize from the liquid, growing inward from the cold barrel wood. The remaining liquid will become enriched in C_2H_5OH, turning into hard cider that accumulates in the center of the barrel. I have tried this experiment in a class using small beakers of ethanol-water mixtures. We made alcohol slush, rather than getting the ice crystals crystals to grow on the sides of the beakers. Decanting was necessary to determine the weight percent solid and fluid composition at various points on the ice saturation curve (see Figure 4). Perhaps a barrel is needed! Liquid compositions can be determined by density or refractive index, if careful measurements are made. This may be a better for a class discussion or homework problem than a lab activity.

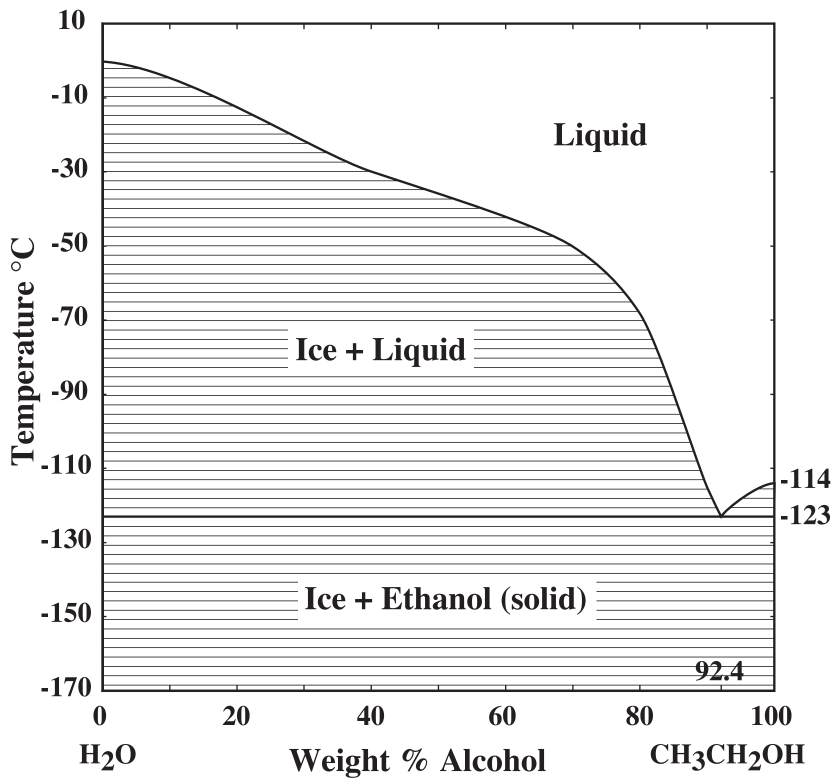

Figure 4. Water-ethanol liquidus diagram drawn from data reported by Timmermans (1960, p. 182).

4. Fractional boiling (distillation)

I haven't found a good low-temperature experiment to demonstrate the melting behavior of solid solutions. However, many of the features important for melting solid solution phases can be explored by boiling ethanol-water solutions. Because college students over the age of 21 have some interest in ethanol-water solutions, most pay careful attention during discussions or activities involving this system. One simple activity is to mix alcohol and water in different proportions and to determine the temperature at which the solution begins to boil (see Figure 5). Be sure to add something to the beaker (small rocks!) to provide a nucleation site for the gas bubbles. Alternatively, the boiling temperature of one mixture of alcohol and water can be monitored while the solution changes composition due to fractional boiling. If the solution is mostly alcohol, with a higher weight percent alcohol than the minimum (96% alsohol) in the saturation curves, the maximum temperature reached will be only 78.3°C. If the solution is more water rich, the maxumum temperature reached will be 100°C. Because ethanol is highly flammable, this experiment should be done in a hood with caution (no nearby ignition sources – use a hotplate not a Bunsen burner).

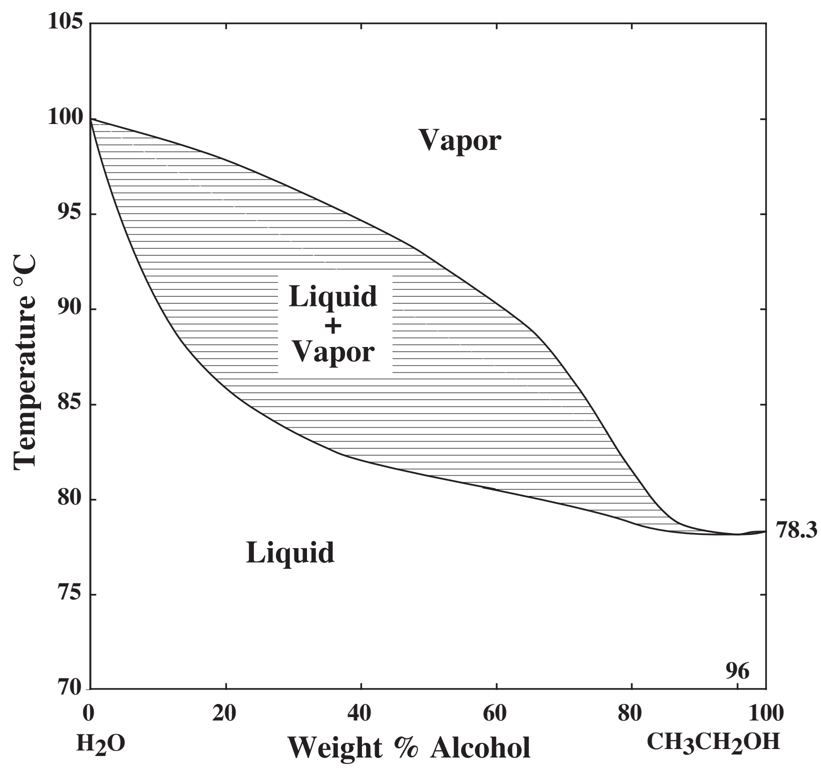

Figure 5. Water-ethanol phase diagram drawn from data reported by Timmermans (1960, p. 173, 179). The minimum on the diagram is at 78.1°C

5. Dew point

Students also have some familiarity with the concept of saturation in the context of relative humidity and the dew point. I use this example to build upon their growing knowledge of phase diagrams and to show them how nicely phase diagrams can organize observational data. To bring out the important details of the air-water system, Figure 6 shows only the air-rich portion of the diagram. Saturation curves are shown for water in air and for ice in air. These curves give the water content of air at 100% humidity. Percent relative humidity of unsaturated air at each temperature is the weight percent water in the air divided by the weight percent water in saturated air at the same temperature times 100. A dashed curve is shown on Figure 6 for 50% relative humidity (a humidity isopleth). I have used this diagram only for discussion or problem sets. It is particularly good for emphasizing the idea that a solution (air) can be saturated by lowering the temperature (reaching the dew point).

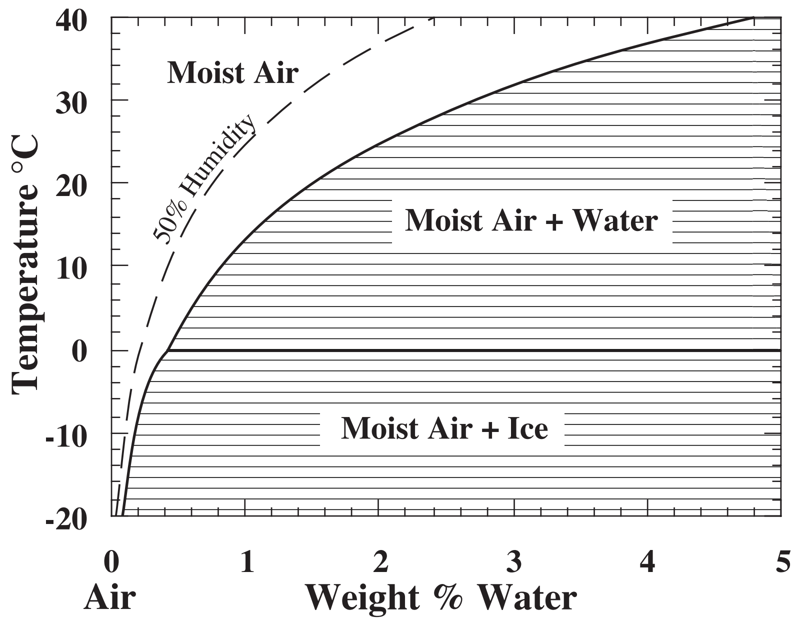

Figure 6. The air-rich portion of the air-water phasediagram at 1 bar pressure drawn from data reported in Gedzelman (1980, Table 10.1).

6. Crystallization boiling

Ever leave a soda can in the freezer too long? The explosion that may have occurred is the result of vapor saturation and then vapor exsolution during crystallization of the liquid. A similar process occurs during the crystallization of ice cubes that contain dissolved air, resulting in air bubbles trapped inside the ice cube. I don't yet have a good low temperature phase diagram to use for experiments or discussions of simultaneous crystallization and boiling. Perhaps a workshop participant can suggest one and/or some good expeiments to try. The forsterite-water system at 1.0 GPa is a good high P-T analog (Hess, 1989, p. 22).

References Cited

Bates, F., Phelps, F.P., and Snyder, C.F. (1928) Saccarimetry, the properties of commercial sugars and their solution, in Washburn, E.W., Editor, International Critical Tables, New York, McGraw-Hill Book Company.

Brady, J.B. (1992) Does ice dissolve or does halite melt? — A low-temperature liquidus experiment for petrology classes: Journal of Geological Education, 40, 116-118.

Darling, R.S. (2000), Modeling magmatic phase equilibria with synthetic H2O-KCl fluid inclusions, Journal of Geoscience Education, 48, 198-202.

Gedzelman, S.D. (1980) The science and wonders of the atmosphere, New York, John Wiley and Sons, 535 p.

Hess, P.C. (1989) Origins of igneous rocks. Cambridge, Harvard University Press, 336p.

Petersen, E.U. (1997), Phase diagrams in vivo, in Brady, J.B., Mogk, D.W., and Perkins, D., III (Editors), Teaching mineralogy,147-158.

Stephens, H., and Stephen, T. (1963) Solubilities of inorganic and organic compounds. Volume I. New York, MacMillan Company, 108-109.

Timmermans, J. (1960) The physico-chemical constants of binary systems in concentrated solutions, New York, Interscience Publishers, v. 4.