I.   **MYTH BUSTERS: Soda Freezing Demonstration.**

        Sometimes when you purchase a bottle of soda from a vending machine, the soda will freeze when you open the bottle. One explanation found on the www is that although the soda is cooled to below 0 °C in the vending machine, the soda remains a liquid because the dissolved gas depresses the freezing point. However, as soon as the soda is opened, it freezes because the concentration of the dissolved solute is decreased when the CO_2 gas bubbles out of solution. Let's do some calculations to check this explanation.

   2 bottles of club soda: one is on ice (0 °C), the other is in a salt/ice bath (-8 °C).

1. Describe the two bottles from a macro perspective before and after the bottles are opened.
   Before opened, both bottles look the same: completely liquid.
   After opened, there is a difference: the bottle in the salt bath is now a slushy, whereas the bottle in just the ice bath is a liquid. Both have bubbles being released from the liquid.

2. Using the concepts of Henry's Law and Freezing Point Depression, prove that the 'myth" is correct/incorrect. The following questions will help you determine the molar concentration of solute in the carbonated water, which will allow you to calculate the freezing point depression, which will help confirm your explanation for the observed effect.

a. Use the concept of Henry's Law to explain the observed "degassing" of the soda upon opening the bottle.
   Henry's Law states that the amount of gas (mole fraction) dissolved in a liquid is dependent on the pressure of the gas above the liquid. When a bottle of soda is sealed at the factory, it is done with extra pressure of CO_2. When the seal on the bottle is broken, the excess CO_2 is released (thus the psst sound of CO_2 being released) and it is only when the seal is broken that bubbles start to appear. Only once the excess CO_2 is released above the liquid can the CO_2 begin coming out of solution.

   It is possible that the gas dissolved in the bottle keeps the soda from freezing at -8°C, but as soon as the pressure is released and the gas comes out of solution, the gas that was depressing the freezing point is gone and the liquid freezes.

b. Assume the partial pressure of CO_2 inside the bottle above the liquid is 2 atm. The Henry's Law constant for CO_2 is 0.117 M * atm^{-1}. Determine the molar concentration of the gas, CO_2

$C_{gas} = k_H P_{gas}$

From the question we know: $k_H = 0.117$ M*atm^{-1} and $P_{gas} = 2$ atm

$C_{gas} = (0.117$ M*atm$^{-1})(2$ atm$)$

$C_{gas} = 0.234$ M

c. Determine the freezing point depression of the soda, given that $K_f = 1.86$ °C/m

Assume that all we have is CO_2 in pure water

$\Delta T = -iK_f m$

$i = 1$ (CO_2 is a molecular solute)

$K_f = 1.86$ °C/m (from table)

m = moles solute/kg solvent

$C_{gas} = 0.234$ moles solute/L solution

Assuming the solution is *mostly* pure water, we can predict that the density of solution is around 1 g/mL (density of pure water). Then, let's assume we have 1 L of solution:

1L*(1000 mL/L)*(1 g/mL) = 1000 g (or 1 kg *of solution*).

In 1 L of solution, there are 0.234 moles of CO_2:

0.234 moles CO_2*(44 g/mol) = 10.296 g of CO_2

Now, out of the 1000 g of total solution (solute + solvent), 1000 g - 10.296 g must be from the pure water solvent. This is 989.704 g of water or 0.989704 kg.

The molality of this solution is then:

(#moles CO_2)/(mass H_2O) = 0.234 mol/0.989704 kg = 0.236 m

The Freezing point depression is then:

$\Delta T = -iK_f m = -1*(1.86$ °C/m$)(0.236$ m$) = -0.44$°C

d. Does your answer support your explanation for why one of the soda's froze upon opening and the other did not? Explain.

No. Based on the small amount of solute dissolved in the solvent there is not enough CO_2 to keep the soda solution from freezing in the -8°C salt bath.

II. Revisit Solutions/Solubility "Cold Solution"
Sodium Acetate Dissolves in Water and the Solution Gets Cold.

1. Describe what happens when sodium acetate dissolves in water using a macroscopic description and a microscopic description including a picture.
   The solid crystals dissolve in water, seeming to "disappear." As this happens, the solution gets cold.

   Microscopically, the sodium acetate splits into sodium and acetate ions in the water in equal ratios (1:1). Water molecules surround the ions.
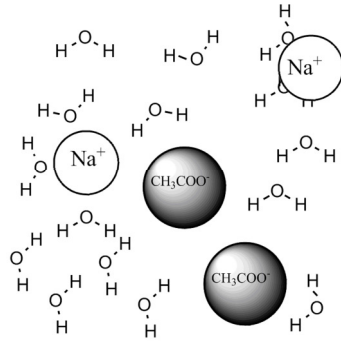
2. Describe what happens thermodynamically when sodium acetate dissolves in water (include ΔS, ΔH, ΔG)
   ΔG Since the sodium acetate does dissolve in water spontaneously, ΔG is negative
   ΔS Since sodium acetate was in a crystal lattice and is now dispersed in water molecules, there are more microstates in the solution. That is, ΔS has increased and is positive.
   ΔH Since we observed that the solution got cold as the sodium acetate dissolved, we know that this is an endothermic process (heat pulled from surroundings and into the system) and ΔH is positive.

3. When one crystal of sodium acetate is added to a super saturated solution of sodium acetate what happens to the solution?
   The solution is so saturated that the single crystal begins a cascading effect of most of the sodium acetate crystallizing or crashing out of solution.

4. Do you expect that the test tube should feel warm or cool? Explain, thermodynamically, your choice of warm or cool?
   This solid is then hot, which we should expect. If the dissolving process was endothermic, the crystallization process (the reverse of dissolving) should be exothermic. As stable bonds are made in the crystal, heat is released.

5.   The sodium acetate concentration in the supersaturated tube was 20 molar. The Ksp for $NaCH_3CO_2$ = 25. In this case is Q > Ksp or < Ksp at room temperature.   Does this calculated value of Q fit with your observation?

$[NaC_2H_3O_2]$ = 20

$Q = [Na^+][C_2H_3O_2^-]$

$Q = [20][20] = 400$ (each ion is 20 M because the ions each exist in a 1:1 ratio with the original salt).

$400 > 25$

$Q > K_{sp}$

Yes, this does fit with the observation as the sodium acetate will fall out of solution with a single crystal to initiate the reaction. It is somewhat surprising that this super saturated solution can exist at all.

III. Now let's think about free energy and vapor pressure lowering.

1. Consider two covered beakers: one containing 100 mL of pure water the other with 100 mL of 0.3 M KNO_3 solution. Which has the lower free energy? Which has the lower vapor pressure?

The solution has the lower free energy as well as the lower vapor pressure. Colligative properties include boiling point elevation. The more solute in a solution, the higher the boiling point. The higher the boiling point, the lower the vapor pressure.

2. Given the vapor pressure of pure water is 25 Torr at room temperature, calculate the vapor pressure of the solution.

$P_{solution} = X_{solvent}P°$

$M_{KNO3}*V_{KNO3}$ = 0.3 M KNO_3*(0.1 L) = 0.03 moles KNO_3

So the total moles of solute is moles K^+ + NO_3^-, giving 0.06 moles

$V_{H20}*d_{H20}*mm_{H20}$=(100 mL H_2O)*(1g/1 mL)*(1 mole/18 g) = 5.56 moles H_2O

Total moles = moles K^+ + moles NO_3^- + molesH_2O = 5.62 moles

$P_{solution} = X_{solvent} P°$= (5.56/5.62)*(25 torr)

$P_{solution}$ = 24.73 Torr

3. The two beakers are placed together in a vacuum bell jar in which all the gas has been evacuated, and they are uncovered. What, if anything, will happen?
The volumes of both the solution and the pure solvent will decrease at first as the liquids evaporate to come to equilibrium. As equilibrium is reached, the volume of the solution will continue to increase until all the liquid from the pure solvent has evaporated and recondensed in the beaker with the solution
4.   Will the pure water or the solution evaporate faster? If so, which one.
The pure water will evaporate faster.

5. What will be the final result of this phenomena (evaporation will stop, evaporation will continue until the free energy is the same in both beakers, evaporation will never start)?

Evaporation will continue until the free energy is the same in both beakers. Since one beaker contains pure water, there is no way to add solute to it to decrease the free energy of the pure solvent. The pure water will always be at a higher free energy than the solution. So, the liquid from the pure solvent will evaporate and recondense in the beaker with the solution until all the liquid is in the solution beaker.

6. What will the final concentration be in each of the beakers?

The concentrations are striving to be equal. However, since there is no solute in the beaker with the pure water, there will only be one solution, of 0.15 M KNO_3.
(0.03 moles KNO_3)/(0.1 L + 0.1L) = (0.03 moles)/(0.2L) = 0.15M KNO_3

7. Draw what you think the picture will look like at the end of the process

IV. Think about the common ion effect in the case of suppressing the dissolution of a toxic metal.

1. Your friend is serving a home made salsa out of a hand crafted pottery dish that was painted with a paint containing lead II carbonate. Lead II carbonate is a white substance that is added to paints to help the colors stay mixed. You caution your friend that according to the EPA a safe concentration of lead in solution is 0.015 mg Pb^{2+}/Liter solution. Your friend says not to worry because he added enough sodium carbonate salt to the salsa to bring that concentration up to .15 molar. Your friend said that this amount of sodium carbonate should limit the lead dissolving into the salsa due to the common ion effect. Given that the Ksp for $PbCO_3 = 7.4 \times 10^{-14}$, determine the concentration of lead ion in the salsa and report whether or not it is below the EPA's standards.

a. Convert the allowed Pb ion concentration into molar units.
0.015 mg Pb^{2+} in one liter
1.5×10^{-5} g *(1 mole/207.2 g) = 7.24×10^{-8} moles Pb^{2+}
7.24×10^{-8} M

b. Determine the molar concentration of Pb ion in the salsa.
R $PbCO_3 \rightarrow Pb^{2+} + CO_3^-$
I some   0   0.15 M
C -x   +x   +x
E some-x   +x   0.15 M + x

Ksp = 7.4×10^{-14} = $[Pb^{2+}][CO_3^{-2}]$
7.4×10^{-14} = [x][0.15 + x]
Since this is a common ion problem, we will ignore x in the CO_3^- concentration. We

will also assume that x is MUCH smaller (by a factor of 100) than 0.15*.

$7.4 \times 10^{-14} = 0.15x$

$x = 4.93 \times 10^{-13}$

$[Pb^{2+}] = 4.93 \times 10^{-13}$

*If you need to prove this to yourself, after solving for x, go back and look at how small x is. In this case, it would now be 0.150000000005. This is not significantly different.

c. Is the concentration of Pb ion in the solution more or less than the concentration of ion allowed by the EPA?

With the common ion, the Pb ion in solution is LESS than that allowed by the EPA.

7.24×10^{-8} M (EPA) $> 4.93 \times 10^{-13}$ M (our solution)