Dissolved Oxygen in Surface Water.

Let’s use thermodynamics to calculate how much oxygen gas (O$_2$) can be dissolved in water at 25°C and 1 bar pressure. Let’s try to get our final answer in parts per million (ppm) by weight (mass).

**Step 1:**
In the space below, write a reaction between oxygen in the atmosphere and oxygen in surface water:

\[ \text{Step 2:} \]

\[ \Delta G_{O_2(g)} = 0.0 \text{ KJ/mol} \]
\[ \Delta G_{O_2(aq)} = 16.4 \text{ KJ/mol} \]

Using the Gibbs Free Energies for oxygen gas and dissolved (aqueous) oxygen gas listed above, calculate the $\Delta G$ of the reaction in Step 1. Express your final answer in J/mole.

**Step 3:**
Calculate the $\log K_{eq}$ for your reaction using the equation discussed in class. Express your final answer with a value for $K_{eq}$ (i.e. $K_{eq} = 10^x$).

**Step 4:**
Write an expression for the $K_{eq}$ using activities of products and reactants.
Step 5:
Using $K_{eq}$ from Step 3, and what you know is the $aO_2 \,(g)$ in our atmosphere, calculate the $aO_2 \,(aq)$ dissolved in surface water:

Step 6:
For “ideal” solutes, activity ($a$) is the same as molality ($m$; moles of solute per kilogram of solvent). Express your answer in Step 5 as molality.

Step 7:
The mass of $O_2$ is 32 grams per mole. Based on your answer in Step 6, how many grams of $O_2 \,(aq)$ are found in surface water?

Step 8:
Lastly, convert the value obtained in Step 7 to ppm (by weight).

How many ppm $O_2$ can be dissolved in surface water? ______________

What do you think happens to the equilibrium amount of $O_2 \,(aq)$ at $T = 5^\circ C$ (i.e. cold water)?

What impact may this have on respiring organisms like fish?