1a) As I said in class, there are two ways to approach this aspect of the problem: calculate the molality and from that value, get the molar mass, or, by substituting algebraically for molality, get the molar mass directly. Either way works. I’ll do the first one.

We have 25.0 mL (=25.0 g = 0.025 kg) of water, and we have a \( K_f = 1.858 \, ^\circ C/m \), so we can get molality:

\[
m = \frac{0.201 \, ^\circ C}{1.858 \, ^\circ C/m} = 0.108 \, m.
\]

Molality is moles solute/kg solvent, so we also know that

\[
0.108 \, m = \frac{\text{moles}}{0.025 \, \text{kg}} \quad \text{and} \quad \text{moles} = 2.7 \times 10^{-3},
\]

which is what 0.243 g represents, so

\[
M_m = \frac{0.243 g}{0.0027 \, \text{moles}} = 89.85 \, g/mol.
\]

b) If we have 53.31% C and 11.18% H, with the balance being oxygen, we have, therefore, 35.51% O. Since we have the molar mass, let’s do the molecular formula first, then get the empirical. Of the 89.85 g of compound, 47.90 g is C, 10.05 is H, and 31.91 g is O, which leads to mole ratios of: 3.988:9.97:1.994, which is nearly 4:10:2, and a molecular formula of \( C_4H_{10}O_2 \). The empirical formula thus is \( C_2H_5O \).

c) There are several possible answers for both parts of this problem; here are a few for each. (I will evaluate yours on their own merits.)

Hydrogen-bonding:

Non-hydrogen-bonding:
2) Recall that the definition of lattice energy is the energy required to do the following reaction:

\[
\text{KCl}(s) \rightarrow \text{K}^+(g) + \text{Cl}^-(g) \quad \Delta H = +717 \text{ kJ/mol. Call this equation A}
\]

The heat of solution is defined by the following equation:

\[
\text{KCl}(s) \rightarrow \text{K}^+(aq) + \text{Cl}^-(aq) \quad \Delta H = +18.0 \text{ kJ/mol. Call this equation B}
\]

We can see (I hope) that this is simply a Hess’ Law problem somewhat disguised. Given the heat for \(\text{Cl}^-(g) \rightarrow \text{Cl}^-(aq)\) \(\Delta H = -338 \text{ kJ/mol. (From 12.105)}\) (Call this equation C), we are asked to get the heat for \(\text{K}^+(g) \rightarrow \text{K}^+(aq)\). Call this equation D. We can see that A + C + D = B and therefore the heats for those reactions following the same relation. Therefore, \(+717 + (-338) + D = 18.0\) (all in kJ/mol) or, solving, we get \(379 + D = 18\) or \(D = -361 \text{ kJ/mol}\). When you compare with Problem 12.105 (which I suspect you all did to see if your method worked), you find that the corresponding value for \(\text{Na}^+\) is -445 kJ/mol. Comparison shows that potassium’s is smaller than sodium’s, which can be explained by smaller ions have larger (i.e. more negative) hydration energies.

3) We are given a molality of 0.119 \(m\). of aqueous formic acid, and a freezing point of -0.229 °C. This last datum tells us that \(\Delta T_f = 0.229\)°C. Since we are in water, we can use \(K_f\) for water which is 1.858 °C/\(m\). We are also told that formic acid is partially dissociated in solution, so we should know to use: \(\Delta T_f = K_f m_i\), and find the value of \(i\). Doing just this yields a value of \(i = 1.0369\). If formic acid did not dissociate, \(i\) would equal 1; since it does, the value to the right of the decimal point represents the fraction of the formic acid which dissociates. the percent dissociation = 3.69%.

4a) This is a problem similar to the first one, but just presented differently. Here we are given an osmotic pressure, and some masses and volumes, and asked to use them. Let’s do it.

We can get the molarity from the osmotic pressure equation, and from that, the number of moles of solute we have (represented by 10.0 mg) and then, finally, the molar mass. We can then use this datum to calculate the freezing point depression, remembering that \(K_f = 1.858 \text{ °C/m. The osmotic pressure equation, } \Pi = MRT, \text{ leads us to, upon rearrangement: } M = \frac{\Pi}{RT} = \frac{.340 \text{torr}}{\sqrt{760 \text{torr/atm}}} = 1.829 \times 10^{-2} \text{ M;} \text{ molarity =}
moles/liter so: \( 1.829 \times 10^{-5} \) = \( \frac{\text{moles}}{0.030L} \), and \( \text{moles} = 5.486 \times 10^{-7} \). This number of moles is represented by 10.0 mg (=0.010 g), so the molar mass, \( M_m = \frac{0.010g}{5.486 \times 10^{-7} \text{moles}} = 18230 \text{ g/mol} \).

b) We have 30.0 mL of a solution with density 0.997 g/mL, so we have a total solution mass of 29.91 g, of which 0.010 g is solute, so we have 29.90 g (=0.02990 kg) of solvent, with the same number of moles of solute (= 5.486 x 10^{-7}). Therefore, we have a molality of: \( m = \frac{5.486 \times 10^{-7} \text{moles}}{0.02990 \text{kg}} = 1.835 \times 10^{-5} \), which leads to a freezing point depression of: \( \Delta T_f = K_f m = 1.858^\circ \text{C/m}(1.835 \times 10^{-5} m) = 3.41 \times 10^{-5}^\circ \text{C} \). (So you can see that this freezing point would be very difficult to measure!)

5) One might be tempted to say “no” just from the manner in which the question is phrased. Reflection, however, should lead one to say “yes.” Here are two reasons why: i) you have equal masses of each component in the solution, so which one is taken to be solvent is somewhat arbitrary; ii) the mole fraction of the individual components is the same, regardless of the masses of the components, since these masses must be equal. (It can be shown via algebra that the mole fraction of either component is merely a ratio of the molar masses of the components arranged in the proper manner.)