In the lab you mix 50.0 mL of 0.250 M Ca(NO3)2 with 50.0 mL of 0.500 M NaF in a coffee cup calorimeter to form a CaF2 precipitate. The initial temperature of each solution is 23 degrees celsius. Assuming that the final solution has a total mass of 100.0 g and a specific heat of 4.18 J/g degrees celsius, calculate the final temperature you expect for the solution. Assume no heat is lost to the calorimeter.

Ca^{2+} (aq) + $2F^-$(aq) --> CaF_2 (s)
$\Delta H° = -115$ kJ/mol

Solution:

First you need to determine which reactant is the limiting one. That is, which reactant will run out first. It will determine how much heat is given off.

moles Ca^{2+} = M Ca^{2+} x L Ca^{2+} = (0.400)(0.0500) = 0.0200 moles Ca^{2+}
moles F^- = M F^- x L F^- = (0.800)(0.0500) = 0.0400 moles F^-

The balanced equation tells us that it takes 2 moles of F^- to react with 1 mole of Ca^{2+}, and that's exactly what we have: 0.0400 moles F^- / 0.0200 moles Ca^{2+} = 2/1. So both reactants will run out at the same time.

The equation also tells us that 1 mole of Ca^{2+} (or 2 moles of F^-) will produce -11.5 kJ of heat. So how much heat will 0.0200 moles of Ca^{2+} produce?

0.0200 moles Ca^{2+} x (-11.5 kJ heat / 1 mole Ca^{2+}) = 0.230 kJ heat = 230 J heat

This amount of heat was absorbed by the water, causing the water temperature to increase.

Heat gained by water = (mass H_2O)(specific heat H_2O)(T_f - T_i)
230 J = (100 g H_2O)(4.18 J / g C)(T_f - 23.0)
$230 = 418T_f - 9614$
$9844 = 418T_f$
$T_f = 23.55$ C.

Answer:T_f=23.55 C.