Answer on Question #62609 - Chemistry - General Chemistry

Question:

6. a. When a 3.88 g sample of solid ammonium nitrate was dissolved in 60.0 g of water in a coffee cup calorimeter, the temperature dropped from 23.0°C to 18.4°C. Calculate ΔH (in kJ per mole of NH4NO3) for the solution process. Assume that the specific heat of the solution is the same as that of pure water (Hint: include the mass of NH4NO3 in the total mass of the solution).

Solution:

1) Let's find the total amount of heat (Q, J) absorbed in the experiment:
   $Q = c*m*(T-T_0)$, where c – specific heat of the solution (J/(g*K)), . m – mass of the solution (g), T_0 and T – initial and final temperature of the solution (K).
   Note that temperature difference expressed in °C is numerically equal to one expressed in K. So we can use temperature given in Centigrade.
   The specific heat of the water is 4.19 J/(g*K) = 4.19 J/(g*°C).
   Do the calculation:
   $Q = 4.19\ J/(g*°C) * (3.88+60.0)\ g * (18.4°C - 23.0°C) = -1231.2\ J.$
2) Let's find the amount of heat referred to one mole of substance:
   Molar mass of NH_4NO_3 (M) = 14.0 + 4*1.0 + 14.0 + 3*16.0 g/mol = 80.0 g/mol.
   Then $\Delta H = Q*M/m_{NH4NO3} = -1231.2\ J * 80.0\ g/mol / 3.88\ g = -25{,}385.6\ J/mol = -25.4\ kJ/mol$

Answer:

The heat of dissolving of NH_4NO_3 in water is -25.4 kJ/mol (minus means that heat is absorbed during dissolving).