Answer Key
CHEMISTRY LAB: Heat of Solution of Ammonium Chloride

Data Table

initial temp water: measured to tenths place with units
mass of ammonium chloride: measured to hundredths with units
final temperature solution: measured to tenths place with units

Post-Lab Questions

this number depends on initial temp. measured and recorded in data table

1. $50.00\,\cancel{mL} \times \frac{0.997799\,g}{1\,\cancel{mL}} = \boxed{49.89\,g\ H_2O}$ 4 significant figures

2. $q = C_s\,m\,\Delta T$

$C_s\,(\text{liquid}) = 4.184\ J/g°C$

$m = \underbrace{49.89\,g + 4.55\,g}_{\text{both numbers depend on measurements}} = 54.44\,g$

$\Delta T = 15.5°C - 22.0°C = -6.5°C$

$q = (4.184\ J/\cancel{g°C})(54.44\,\cancel{g})(-6.5\,\cancel{°C}) = \boxed{-1500\ J}$ 2 significant figures

3. $4.55\,\cancel{g}\ NH_4Cl \times \frac{1\ \text{mole}}{53.4915\,\cancel{g}} = \boxed{0.0851\ \text{moles}\ NH_4Cl}$ 3 significant figures

4. $\Delta H = \frac{-q}{mol}$ $\quad q = -1500\ J \times \frac{1\,kJ}{1000\,J} = -1.5\,kJ$

$\Delta H = \frac{-(-1.5\,kJ)}{0.0851\ mol} = \boxed{+18\ kJ/mole}$ 2 significant figures

5. $\boxed{NH_4Cl\,(s) \rightarrow NH_4^+\,(aq) + Cl^-\,(aq) \quad \Delta H = +18\ kJ/mole}$

6. enthalpy (kJ/mole) in an intensive property $\boxed{A}$

7. positive enthalpy, decrease in temperature
heat flows from surrounding to system $\boxed{B}$

8. positive enthalpy, endothermic $\boxed{A}$

9. separating the intermolecular forces in the solvent (ΔH_1) and the solute (ΔH_2) costs energy (positive enthalpy). Forming new intermolecular forces between them (ΔH_3) is thermodynamically favorable (negative enthalpy). Dissolving ammonium chloride has positive enthalpy, so $\Delta H_1 + \Delta H_2$ must be larger $\boxed{A}$

10. $NH_4Cl\,(s) \rightleftharpoons NH_4^+\,(aq) + Cl^-\,(aq)$ $\quad K_{sp} = [NH_4^+][Cl^-]$
more soluble = larger K_{sp} $\boxed{A}$

11. percent error $= \dfrac{|\text{error}|}{\text{accepted}} \times 100\%$

$\text{error} = 15 - 14.7 = 0.3\ \text{kJ}$
$\text{percent error} = \dfrac{0.3}{14.7} \times 100\% = \boxed{2\%\ \text{error}}$
1 significant figure

12. less NH_4Cl ends up in calorimeter than accounted for in the calculation of enthalpy, ΔT would be smaller $\boxed{B}$

13. less H_2O in the calorimeter than accounted for in the calculation of enthalpy, ΔT would be larger $\boxed{A}$

14. Since the student decided to make the change, we can assume that these amounts were accounted for in the calculation of enthalpy $\boxed{C}$

15. Again, we can assume the student accounted for the changed amounts in the calculation $\boxed{C}$

16. less NH_4Cl in solution, smaller ΔT $\boxed{B}$

17. $m = \frac{\text{moles } NH_4Cl}{\text{kg } H_2O}$ $\qquad 49.89 \text{ g } H_2O \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.4989 \text{ kg } H_2O$

$m = \frac{0.0851 \text{ moles}}{0.4989 \text{ kg } H_2O}$ } numbers depend on measurement

$m = \boxed{0.171 \text{ mol/kg}}$ 3 significant figures

18. [tie breaking question]

$\Delta T_F = K_F\ m\ i$ $\qquad K_F = 1.853 \text{ K} \cdot \text{kg/mol}$

$m = 0.171 \text{ mol/kg}$ {from previous}

i = van't Hoff factor.

NH_4Cl dissociates to 2 ions

$\Delta T_F = (1.853 \text{ K} \cdot \cancel{\text{kg}}/\cancel{\text{mol}})(0.171\ \cancel{\text{mol}}/\cancel{\text{kg}})(2)$

$\Delta T_F = 0.634°C$

$T_F = T_{F(\text{pure water})} - \Delta T_F = 0.000 - 0.634°C = \boxed{-0.634°C}$

19. intensive $\boxed{A}$

20. physical $\boxed{A}$