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Topic 6.10: Enthalpy of Solution

Worksheet · AP Chemistry

UNIT 6 · THERMOCHEMISTRY

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How to Use This Worksheet

Work in order. Section A checks the definitions and the sign rules. Section B practises finding ΔHsol per mole from a measured heat, predicting temperature changes with q = mcΔT, and adding lattice and hydration enthalpies. Section C is exam style: judging claims and two full scenarios.

Reference facts. ΔHsol is the enthalpy change when one mole of solute dissolves to form a dilute solution, reported from the system's point of view: positive for endothermic dissolving, negative for exothermic dissolving. Useful values in kJ/mol: sodium chloride +3.9, potassium chloride +17.2, potassium nitrate +34.9, ammonium nitrate +25.7, ammonium chloride +14.8, potassium iodide +20.3, sodium hydroxide −44.5, calcium chloride −81.3, lithium chloride −37.0, magnesium sulfate −91.2, calcium hydroxide −16.7. Lattice enthalpies are always positive and hydration enthalpies are always negative: sodium chloride +787 and −783, potassium chloride +717 and −700, sodium hydroxide +900 and −944, potassium iodide +632 and −612, lithium chloride +853 and −890. The cycle is ΔHsol = ΔHlattice + ΔHhydration. Solubility data: potassium nitrate dissolves 31.6 g per 100 g of water at 20 °C and 63.9 g at 40 °C; calcium hydroxide dissolves 0.173 g per 100 g of water at 20 °C and 0.077 g at 60 °C; sodium chloride dissolves 35.7 g per 100 g of water at 0 °C and 39.1 g at 100 °C. The specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$.

Three habits protect every answer. Name the system before you assign any sign. Check every energy ledger balances: gain on one side, equal loss on the other. Read what a hand or thermometer feels as a report about the surroundings, not the system.

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Do not look at the answer key until you have attempted every question.

Section A: Recall and Concepts

A1. (a) Define the enthalpy of solution, ΔHsol, including the amount of solute. (b) State the system and the surroundings when a salt dissolves in a beaker of water. (c) Write the dissolving equation for sodium chloride, including state symbols. (d) State the value of ΔHsol for sodium chloride and what its sign tells you about the heat flow.

A2. (a) State the sign of ΔHsol for an exothermic dissolving process, from the system's point of view, and explain the sign. (b) State the sign of ΔHsol for an endothermic dissolving process, from the system's point of view, and explain the sign. (c) Name one salt from the reference facts that warms its solution and one that cools its solution, giving the ΔHsol value for each. (d) State which type of dissolving powers a cold pack, and explain why.

A3. (a) State the sign of the lattice enthalpy step and explain why it has that sign. (b) State the sign of the hydration enthalpy step and explain why it has that sign. (c) Write the equation that combines the two steps to give ΔHsol. (d) For sodium chloride, the lattice enthalpy is +787 kJ/mol and the hydration enthalpy is −783 kJ/mol. Calculate ΔHsol and compare your answer with the measured value of +3.9 kJ/mol.

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A4. (a) State the qualitative rule linking the sign of ΔHsol to how solubility changes with temperature. (b) Potassium nitrate has ΔHsol = +34.9 kJ/mol. State whether its solubility increases or decreases as the temperature rises, and explain why. (c) Calcium hydroxide has ΔHsol = −16.7 kJ/mol. State whether its solubility increases or decreases as the temperature rises, and use the reference solubility data to support your answer. (d) Sodium chloride has ΔHsol = +3.9 kJ/mol. State what the reference data show about its solubility change with temperature, and explain why the change is so small.

A5. (a) State what happens to the heat exchanged when the mass of salt dissolved is doubled, and what happens to ΔHsol in kJ/mol. (b) Rank these three salts from the one that warms a solution most per mole to the one that warms it least: magnesium sulfate, calcium chloride, sodium hydroxide. Justify the ranking with the reference values. (c) Rank these three salts from the one that cools a solution most per mole to the one that cools it least: potassium nitrate, ammonium nitrate, ammonium chloride. Justify the ranking. (d) A student says potassium nitrate cools water more than ammonium chloride because it is more soluble. State whether this reasoning is correct, and give the correct reason.

Section B: Core Practice

B1. A student dissolves 8.00 g of sodium hydroxide (0.200 mol) in water and the solution warms up. The dissolving releases 8.90 kJ of heat. (a) Calculate ΔHsol for sodium hydroxide, per mole, with its sign. (b) State the system and the surroundings. (c) Write the energy ledger for this dissolving: the change for the system and the change for the surroundings. (d) State what happens to the average kinetic energy of the water particles. (e) Explain why the beaker feels hot.

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B2. A student dissolves 20.22 g of potassium nitrate (0.200 mol) in 200 g of water and the solution cools. The dissolving absorbs 6.98 kJ of heat. (a) Calculate ΔHsol for potassium nitrate, per mole, with its sign. (b) State the system and the surroundings. (c) Calculate the temperature fall of the 200 g of water, showing your working. (d) State what happens to the average kinetic energy of the water particles. (e) The student repeats the experiment with 40.44 g of potassium nitrate in the same 200 g of water. State the heat absorbed and the new temperature fall.

B3. A student dissolves 5.55 g of anhydrous calcium chloride (0.0500 mol) in 200 g of water. (a) Calculate the heat released, in kJ and in J, using ΔHsol = −81.3 kJ/mol. (b) Calculate the temperature rise of the 200 g of water, showing your working. (c) State the system and the surroundings. (d) Explain why calcium chloride is used in hot packs, using the size of its ΔHsol. (e) The student dissolves the same 5.55 g in 400 g of water instead. State the new temperature rise and explain why it differs.

B4. Use the reference lattice and hydration enthalpies. (a) For potassium iodide, the lattice enthalpy is +632 kJ/mol and the hydration enthalpy is −612 kJ/mol. Calculate ΔHsol and compare it with the measured value of +20.3 kJ/mol. (b) For lithium chloride, the lattice enthalpy is +853 kJ/mol and the hydration enthalpy is −890 kJ/mol. Calculate ΔHsol and compare it with the measured value of −37.0 kJ/mol. (c) State which of the two salts warms its solution and which cools its solution, using your answers. (d) Explain, in terms of the two steps, why lithium chloride warms its solution while potassium iodide cools its solution. (e) A student claims the lattice enthalpy of lithium chloride must be negative because its ΔHsol is negative. Correct the student.

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B5. A student dissolves 10.70 g of ammonium chloride (0.200 mol) in 100 g of water and the solution cools. (a) Calculate the heat absorbed, in kJ and in J, using ΔHsol = +14.8 kJ/mol. (b) Calculate the temperature fall of the 100 g of water, showing your working. (c) State the system and the surroundings in this experiment. (d) Compare this fall with B2, where 0.200 mol of potassium nitrate cooled 200 g of water by 8.35 °C. Give two reasons for the difference. (e) State whether ammonium chloride or potassium nitrate cools the same mass of water more per mole, and justify your answer.

B6. Use the reference solubility data for potassium nitrate: 31.6 g dissolve in 100 g of water at 20 °C and 63.9 g at 40 °C. (a) Calculate how many times more soluble potassium nitrate is at 40 °C than at 20 °C. (b) Calculate the mass of potassium nitrate that dissolves in 250 g of water at 40 °C. (c) Calculate the mass that dissolves in 250 g of water at 20 °C, and how much more dissolves at 40 °C. (d) Explain why the solubility rises so steeply with temperature, using the sign of ΔHsol. (e) A saturated potassium nitrate solution at 40 °C is cooled to 20 °C. State what happens to the dissolved potassium nitrate, and explain why.

Section C: Exam Style

C1. Judge each claim, supporting each verdict with the definition or sign rule it depends on. (a) "ΔHsol = +25.7 kJ/mol for ammonium nitrate means the surroundings gained 25.7 kJ of enthalpy per mole." (b) "The lattice enthalpy is negative, because breaking a crystal apart releases the energy stored in it." (c) "Calcium hydroxide becomes more soluble when heated, because all salts become more soluble when heated." (d) "Dissolving 0.400 mol of sodium hydroxide releases twice as much heat as dissolving 0.200 mol, but ΔHsol stays at −44.5 kJ/mol." (e) "Magnesium sulfate warms water more than calcium chloride per mole, because −91.2 kJ/mol is more negative than −81.3 kJ/mol."

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C2. A cold pack contains ammonium nitrate and water. Squeezing the pack mixes them and the pack turns cold: $NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$. In a test, dissolving 16.01 g of ammonium nitrate (0.200 mol) in 100 g of water absorbs 5.14 kJ of heat. (a) Calculate ΔHsol for ammonium nitrate, per mole, with its sign. (b) Calculate the temperature fall of the 100 g of water, showing your working. (c) Identify the system and the surroundings when the pack is held against an injured ankle. (d) Explain, in terms of heat flow and particle kinetic energy, why the ankle feels cold. (e) State the energy change of the surroundings for the 0.200 mol test, with its sign, and show that the total energy change of system plus surroundings is zero.

C3. A student compares two salts for a hot pack: magnesium sulfate, ΔHsol = −91.2 kJ/mol, and calcium chloride, ΔHsol = −81.3 kJ/mol. The student dissolves 6.02 g of magnesium sulfate (0.0500 mol) in 50 g of water. (a) Calculate the heat released, in kJ and in J. (b) Calculate the temperature rise of the 50 g of water, showing your working. (c) Calculate the heat that 0.0500 mol of calcium chloride would release in the same test, and state which salt warms the water more. (d) Explain why magnesium sulfate has the more negative ΔHsol, using the charges of its ions and the hydration step. (e) The student says magnesium sulfate is the better hot pack salt, so it must also be the more soluble salt. State whether this conclusion follows, and explain why.

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Answer Key

A1. (a) ΔHsol is the enthalpy change when one mole of solute dissolves in enough solvent to form a dilute solution. (b) The system is the dissolving salt. The surroundings are the water, the beaker, and everything touching them. (c) $NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$. (d) ΔHsol = +3.9 kJ/mol: the system absorbs a little heat, so the solution cools very slightly.

A2. (a) Negative. The system releases heat and loses enthalpy, so H(products) is less than H(reactants). (b) Positive. The system absorbs heat and gains enthalpy, so H(products) is greater than H(reactants). (c) Warms: sodium hydroxide, ΔHsol = −44.5 kJ/mol. Cools: potassium nitrate, ΔHsol = +34.9 kJ/mol. (d) Endothermic dissolving powers a cold pack: it absorbs heat from the surroundings, including the skin.

A3. (a) Positive. Separating oppositely charged ions against their attractions costs energy. (b) Negative. Water molecules form attractions with the separated ions, and forming attractions releases energy. (c) ΔHsol = ΔHlattice + ΔHhydration. (d) ΔHsol = +787 + (−783) = +4 kJ/mol, which matches the measured +3.9 kJ/mol within rounding.

A4. (a) If dissolving is endothermic, solubility increases with temperature. If dissolving is exothermic, solubility decreases with temperature. (b) It increases: the dissolving is endothermic, so heat acts like a reactant and raising the temperature pushes it forward. (c) It decreases: 0.173 g dissolve per 100 g of water at 20 °C but only 0.077 g at 60 °C. (d) Almost no change: 35.7 g at 0 °C and 39.1 g at 100 °C, because ΔHsol is only +3.9 kJ/mol.

A5. (a) The heat exchanged doubles, but ΔHsol in kJ/mol stays the same, because it is defined per mole. (b) Magnesium sulfate, calcium chloride, sodium hydroxide: −91.2, −81.3, −44.5 kJ/mol. (c) Potassium nitrate, ammonium nitrate, ammonium chloride: +34.9, +25.7, +14.8 kJ/mol. (d) No. Cooling depends on ΔHsol, not solubility: +34.9 against +14.8 kJ/mol.

B1. (a) ΔHsol = 8.90 ÷ 0.200 = 44.5 kJ/mol, and the heat is released by the system, so ΔHsol = −44.5 kJ/mol. (b) The system is the dissolving sodium hydroxide; the surroundings are the water, beaker, air and hand. (c) System: −8.90 kJ. Surroundings: +8.90 kJ. The two cancel, as conservation requires. (d) It increases: the energy becomes kinetic energy of the water particles. (e) Heat flows from the system into the surroundings, so the beaker warms and your hand detects it as hot.

B2. (a) ΔHsol = 6.98 ÷ 0.200 = 34.9 kJ/mol, and the heat is absorbed by the system, so ΔHsol = +34.9 kJ/mol. (b) The system is the dissolving potassium nitrate; the surroundings are the water, beaker and air. (c) Heat absorbed = 6,980 J. Temperature fall = 6,980 ÷ (200 × 4.18) = 6,980 ÷ 836 = 8.35 °C. (d) It decreases: the water particles gave up energy to the system. (e) Heat absorbed = 0.400 × 34.9 = 13.96 kJ. Temperature fall = 13,960 ÷ 836 = 16.70 °C, double the first fall, because the heat doubled while the water mass stayed the same.

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B3. (a) Heat released = 0.0500 × 81.3 = 4.065 kJ = 4,065 J. (b) Temperature rise = 4,065 ÷ (200 × 4.18) = 4,065 ÷ 836 = 4.86 °C. (c) The system is the dissolving calcium chloride; the surroundings are the water, beaker and air. (d) Its ΔHsol is strongly negative, −81.3 kJ/mol, so a few moles release a large amount of heat. (e) Temperature rise = 4,065 ÷ (400 × 4.18) = 4,065 ÷ 1,672 = 2.43 °C, half the first rise, because the same heat is shared among twice as much water.

B4. (a) ΔHsol = +632 + (−612) = +20 kJ/mol, which matches the measured +20.3 kJ/mol within rounding. (b) ΔHsol = +853 + (−890) = −37 kJ/mol, which matches the measured −37.0 kJ/mol. (c) Lithium chloride warms its solution (ΔHsol negative); potassium iodide cools its solution (ΔHsol positive). (d) For lithium chloride, hydration (890) beats the lattice (853), so the net is negative. For potassium iodide, the lattice (632) beats hydration (612), so the net is positive. (e) The lattice enthalpy is always positive: breaking a lattice apart costs energy. The negative ΔHsol comes from hydration releasing more energy than the lattice step costs.

B5. (a) Heat absorbed = 0.200 × 14.8 = 2.96 kJ = 2,960 J. (b) Temperature fall = 2,960 ÷ (100 × 4.18) = 2,960 ÷ 418 = 7.08 °C. (c) The system is the dissolving ammonium chloride; the surroundings are the water, beaker and air. (d) Two reasons: ammonium chloride absorbs less heat per mole (2.96 kJ against 6.98 kJ), and the heat is shared among half as much water (100 g against 200 g). The two effects pull in opposite directions, so the falls are close: 7.08 °C against 8.35 °C. (e) Potassium nitrate: +34.9 against +14.8 kJ/mol, so it absorbs more heat for the same number of moles.

B6. (a) 63.9 ÷ 31.6 = 2.02, so potassium nitrate is about 2.0 times more soluble at 40 °C than at 20 °C. (b) 63.9 × 2.5 = 159.75 g, about 160 g, dissolve in 250 g of water at 40 °C. (c) 31.6 × 2.5 = 79.0 g dissolve at 20 °C. The difference is 159.75 − 79.0 = 80.75 g, about 81 g more at 40 °C. (d) The dissolving is endothermic, ΔHsol = +34.9 kJ/mol, so heat acts like a reactant and raising the temperature pushes it forward strongly. (e) Solid potassium nitrate crystallises out: cooling removes the heat the endothermic dissolving needs, so the solution holds less salt.

C1. (a) False. ΔHsol is defined from the system's point of view: +25.7 kJ/mol means the system gained 25.7 kJ per mole. The surroundings lost 25.7 kJ per mole. (b) False. Breaking a lattice apart costs energy, so the lattice enthalpy is always positive. (c) False. Calcium hydroxide dissolves exothermically, so its solubility falls as the temperature rises: 0.173 g per 100 g of water at 20 °C but only 0.077 g at 60 °C. (d) True. Heat exchanged scales with the quantity dissolved; ΔHsol in kJ/mol is defined per mole, so it is unchanged. (e) True. A more negative ΔHsol means more heat released per mole.

C2. (a) ΔHsol = 5.14 ÷ 0.200 = 25.7 kJ/mol, and the heat is absorbed by the system, so ΔHsol = +25.7 kJ/mol. (b) Heat absorbed = 5,140 J. Temperature fall = 5,140 ÷ (100 × 4.18) = 5,140 ÷ 418 = 12.30 °C. (c) The system is the dissolving ammonium nitrate inside the pack. The surroundings are the pack material, the ankle and skin, and the air. (d) The dissolving absorbs heat, so heat flows from the skin into the pack, the skin's particles lose kinetic energy, and that loss is felt as cold. (e) System +5.14 kJ, surroundings −5.14 kJ. Total: +5.14 + (−5.14) = 0.

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C3. (a) Heat released = 0.0500 × 91.2 = 4.56 kJ = 4,560 J. (b) Temperature rise = 4,560 ÷ (50 × 4.18) = 4,560 ÷ 209 = 21.82 °C. (c) Heat from calcium chloride = 0.0500 × 81.3 = 4.065 kJ. Magnesium sulfate warms the water more, because 4.56 kJ is greater than 4.065 kJ for the same number of moles. (d) Magnesium sulfate has doubly charged Mg^{2+} and SO_4^{2-} ions, so its hydration enthalpy is very large and negative, outweighing the lattice enthalpy by more. (e) No. Warming depends on ΔHsol; solubility depends on lattice and hydration enthalpies and on entropy. A strongly exothermic dissolver can still be only moderately soluble.

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