Students' Reasoning About Entropy in Chemical and Physical Contexts

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Turkish Preservice Chemistry Teachers

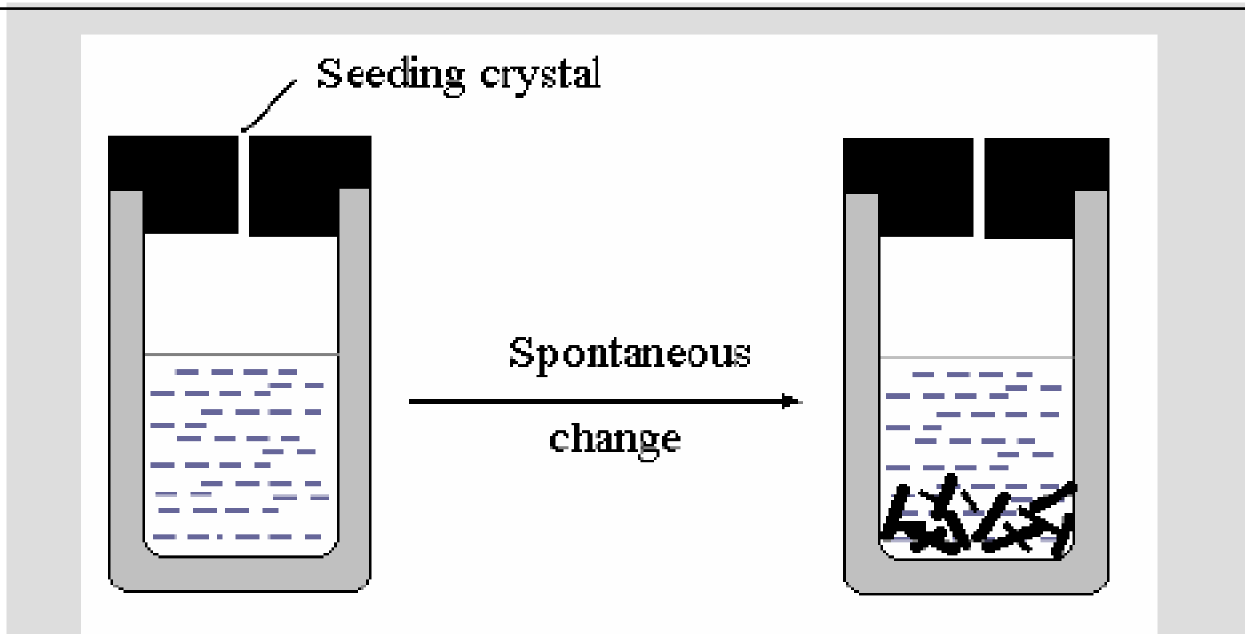
- Third-year undergraduates enrolled in one-year physical chemistry course; $N = 46$
- Lecture course covering both classical and statistical thermodynamics
- Free-response questions given pre- and post-instruction
- Follow-up interviews with $\approx$ 25% of students after pre-test, and with $\approx$ 10% after post-test

Reference: M. Sözbilir and J. Bennett, J. Chem. Educ. **84**,1204 (2007)

Authors' Summary of Students' "Misunderstandings"

- Defining entropy as "disorder" and considering visual disorder and entropy as synonymous
- Inaccurately connecting entropy with the number of collisions and intermolecular interactions
- Inaccurately connecting entropy of the system and accompanying entropy changes in the surroundings
- Believing that the entropy of the whole system decreases or does not change when a spontaneous change occurs in an isolated system

A tiny seeding crystal is dropped into a sealed, thermally insulated flask of supercooled, saturated solution; crystallization occurs spontaneously with an apparent increase of organization. What happens to the entropy of the system [inside the flask] when the crystals form?



Expected Answer: Entropy of the system increases since entropy of universe must increase. Increasing organization (fewer available states) is compensated by increased temperature (more available states) due to bond formation.

Responses to Crystallization Question

- Completely correct:
  - Pretest: 0%; Post-test: 11%
- Entropy of the system increases:
  - Pretest: 13%; Post-test: 6%
- Entropy of the system decreases:
  - Pretest: 26%; Post-test: 40%
- Entropy of the system doesn't change:
  - Pretest: 6%; Post-test: 4%

Summary: Identification of entropy with "disorder" overshadowed implications of second law of thermodynamics

English University Chemistry Students

- First-year undergraduates enrolled in seven-week course in chemical thermodynamics; $N = 16$
- Lecture course on classical and statistical thermodynamics, supplemented with weekly problem sessions; (previous course on QM)
- Students interviewed pre- and post-instruction

Reference: E. Carson and J. Watson, University Chem. Educ. **6**, 4 (2002)

Interview Protocol

- Students asked to consider three different chemical reactions (e.g., magnesium ribbon dissolving in hydrochloric acid);
- Students asked to respond in thermodynamic terms to questions about the reactions, e.g. “What happened in this reaction to cause the temperature change?” and “Why did the reaction happen?”
- Students then asked directly what they understood by the terms “entropy” and “Gibbs free energy.”

Authors' Summary of Students' Thinking

- Did not clearly differentiate between enthalpy and entropy or between Gibbs free energy and entropy, seeing them all as simply "forms of energy";
- Showed frequent confusion about *system* and *surroundings*, often with surroundings being ignored; neglected effects of energy transfer to and from the surroundings;
- Students talked about randomness or disorder but failed to explain what these terms meant; made no mention of microstates nor of energy levels and simply explained entropy as randomness or disorder.

Instructor’s target concepts and related student ideas (I)

- For a chemical reaction to be possible, the total entropy change ($\Delta S_{system} + \Delta S_{surroundings}$) must be positive or not negative.
  - Pre-instruction ≈ 25% of students “with this idea”
  - Post-instruction: ≈ 60% of students “with this idea,” but ≈ 25% had acquired previously absent incorrect ideas

Instructor's target concepts and related student ideas (II)

- A spontaneous reaction is one that is thermodynamically feasible; that is, one for which the total entropy change (ΔS_{system} + $\Delta S_{surroundings}$) is positive. The reaction may not occur because of kinetic barriers.
  - Pre-instruction ≈ 15% of students "with this idea"
  - Post-instruction: ≈ 20% of students "with this idea," but ≈ 50% had acquired previously absent incorrect ideas

Instructor's target concepts and related student ideas (III)

- If energy transferred from an exothermic reaction increases the temperature of the surroundings, the entropy of the surroundings is increased. [$\Delta S_{surroundings}$ can be calculated from the expression: $\Delta S_{surroundings} = -\Delta H/T$].
  - Pre-instruction ≈ 20% of students "with this idea"
  - Post-instruction: ≈ 40% of students "with this idea," but ≈ 30% had acquired previously absent incorrect ideas

Comparison:Introductory Physics Students' Thinking on Spontaneous Processes

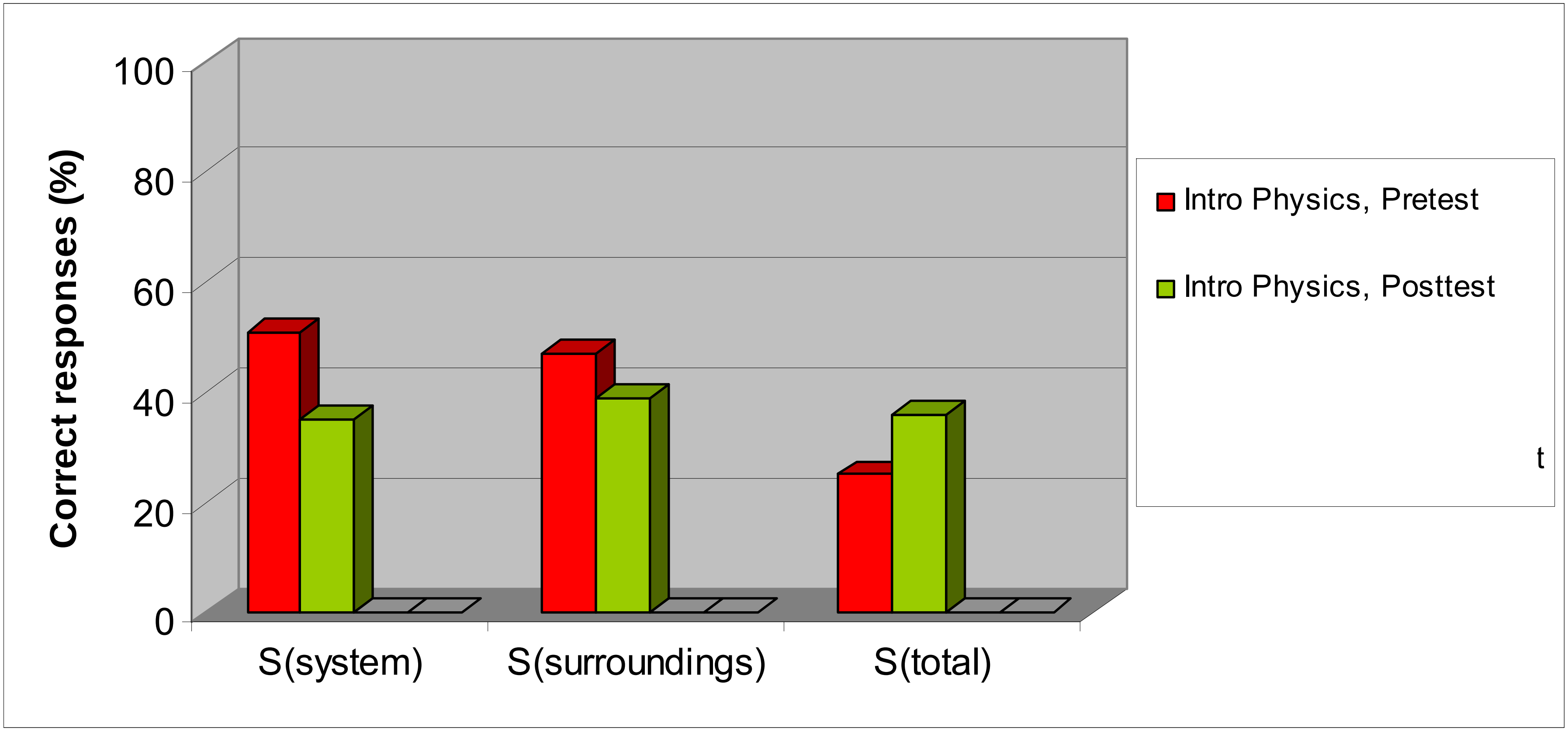
- Tendency to assume that "system entropy" must *always* increase

- Do not accept the idea that entropy of system plus surroundings ***increases***

  - *Strong implied belief in "conservation" of total entropy*
  - *Little change after standard instruction*

Physics Students' Responses on Spontaneous Process

before … and *after* instruction…



Little change on post-test

Summary

- Studies of student thinking in chemical context are consistent with those made in physics context
- Confusion between “system” and “surroundings” is a common problem
- Acceptance of total entropy increase is a common difficulty