Department of Chemistry
University of Texas at Austin

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Physical Equilibria Unit Activity - "Thinking About Solutions" KEY

A major goal for this class is for you to learn the concept of macro/micro thinking or "Thinking Like a Chemist". Thinking like a chemist is the ability to look at the macroscopic properties of a static substance or a substance undergoing a change and be able to simultaneously account for those properties on a microscopic (molecular) level. Today we will practice this skill while considering the process of dissolution. Platinum stars will be on the line.

Write your answers on a separate sheet of paper.

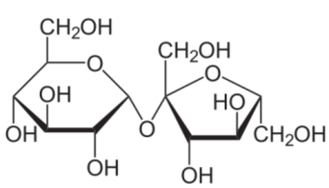
Consider the following two crystalline substances, sucrose (table sugar) and sodium chloride (table salt). You should have prior experience at placing each of these substances in water at room temperature.

1. Please draw/explain to the best of your ability a **macroscopic** description of the sugar dissolving in water and the salt dissolving in water:

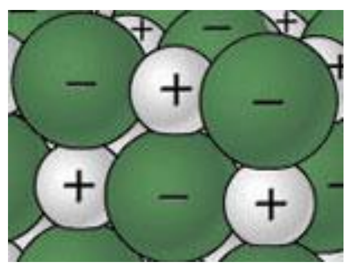
Both sugar and salt are crystalline white solids that dissolve easily in water such that the water looks about the same as it did before anything was dissolved in it. Also, the temperature did not change.

Given the chemical formulae/structures, consider the microscopic properties.

$C_{12}H_{22}O_{11}$



NaCl



(Where Na^+ is represented by the white cations and Cl^- is represented by the green anions).

2. What type of solid is each of these substances?

Sucrose is a molecular solid

NaCl is an ionic solid

3. What types of intermolecular forces hold these substances together as solids?

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Molecular solids are held together by intermolecular forces (IMFs). Sucrose has hydrogen-bonding, dipole-dipole and London dispersion forces.

Ionic solids are held together by ionic charges—that is, FULL charges—of the atoms or polyatomic ions. Their interaction is governed by the coulombic attraction force (aka ionic forces).

4. Draw a microscopic view of the dissolution process for each solid.

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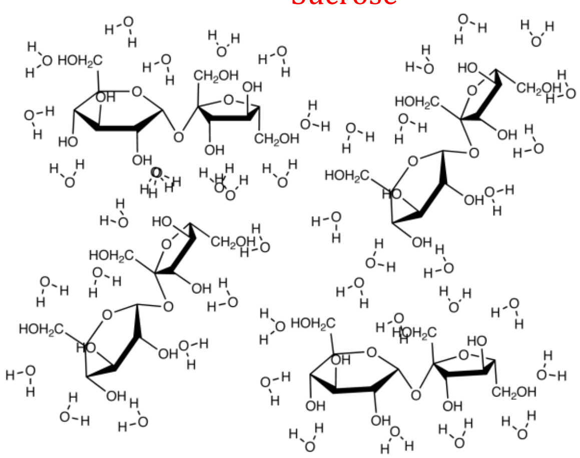
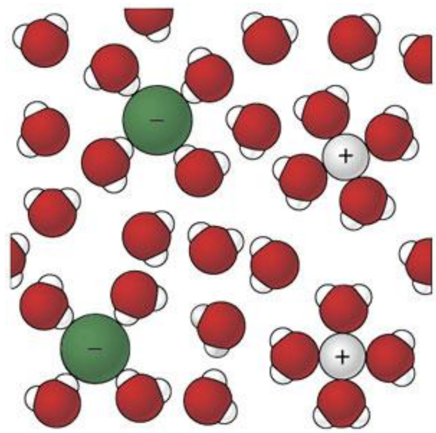
Sucrose



NaCl



(Where Na^+ is represented by the white cations and Cl^- is represented by the green anions).

Now we are ready to look at the energy changes associated with the dissolution process.

5. Is the dissolution process an endothermic or exothermic process?

Dissolution can be either endothermic or exothermic.

Watch a class demonstration on dissolving different compounds (sodium chloride, sucrose, and ammonium nitrate in water).

6. In general, is the dissolution process endothermic or exothermic process? Explain.

In general, dissolution is an endothermic process. In general, it takes more energy to break the forces (whether IMF or ionic) that hold the solid together than are gained from the solvent (water) interacting with the solute (either sucrose or NaCl).

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7. Fully describe the change in enthalpy for the dissolution process, that is what parts of the process should require energy and in what parts of the process should energy be given off.

$\Delta H_{lattice\ energy} + \Delta H_{solvation} = \Delta H_{solution}$

Typically, it will take energy (endothermic process) to break the forces holding the solute (solid) together in its lattice structure ($\Delta H_{lattice\ energy}$), and energy will be given off (exothermic process) as the solvent interacts with the solute ($\Delta H_{solvation}$). Typically, it will require more energy to break the bonds in the solid. So, $\Delta H_{lattice\ energy}$ is usually positive (endothermic) and larger in magnitude than $\Delta H_{solvation}$, which is negative (exothermic). Therefore, usually entire dissolution process is endothermic (positve $\Delta H_{solution}$).

If the IMFs of the solid (lattice energy) are greater than that of any new IMFs formed by the solution, dissolution will be endothermic (typical case).

If the IMFs of the solid are less than those of the any new IMFs formed by the solution, dissolution will be exothermic (not typical, but possible).

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8. Is the entropy of the system (pure solvent + pure solute) higher or lower after the solute has been added to the solvent and a solution has formed?

The entropy of the system is greater after a solution is formed – greater dispersal of energy.

9. Is there ever a situation in which the entropy of the solvent and solute would be greater than the entropy of the solution? Explain

In cases where there are high charge density ions which "lock in" water molecules around them, the entropy of the solution can actually decrease. Another example is dissolved gasses in liquids.

10. Is the dissolution process always spontaneous?

Dissolution is not always spontaneous, especially not at all temperatures.

11. What thermodynamic property can predict the spontaneity of the dissolution process?

Gibb's free energy. When ΔG is negative, the process is spontaneous.

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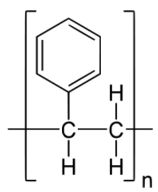
12. Thinking about the Gibb's free energy, what must be true about the enthalpy, entropy, and temperature for dissolution to be spontaneous?

$\Delta G = \Delta H - T\Delta S$
Given this equation, for the typical dissolution process, one that is endothermic and has an increase in entropy, for ΔG to be negative: ΔH must be SMALLER than TΔS

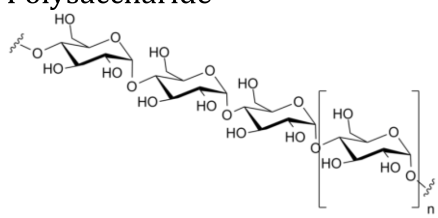
Also, temperature can be a significant factor in dissolution. An increase in temperature will increase the TΔS factor such that it can overcome the positive ΔH and dissolution can be spontaneous

Watch a class demonstration on dissolving different packing peanuts in acetone vs water.

Water (H_2O)
Standard Packing Peanuts
Polystyrene



Acetone (CH_3COCH_3)
Bio-degradable Packing Peanuts
Polysaccharide



Consider the four possible combinations
A. polystyrene packing peanuts in water – Does not dissolve
B. polystyrene packing peanuts in acetone – Does Dissolve
C. bio-degradable peanuts in water – Does Dissolve
D. bio-degradable peanuts in acetone - Does not Dissolve

13. What do you think the sign is for change in free energy of solution for each of these: "+", "-" or about zero.

A. +
B. –
C. –
D. +

What do you think the sign is for change in entropy of solution for each of these: "+", "-" or about zero.

A. +
B. +
C. +
D. +

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What do you think the sign is for change in enthalpy of solution for each of these: “+”, “-“ or about zero.

A. +
B. Around 0
C, Around 0
D. +

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