Exercise 11.2
Solution Formation - Answers

Name: ______________________

Date:________________ Per:________

DIRECTIONS: Answer the following in the space provided:

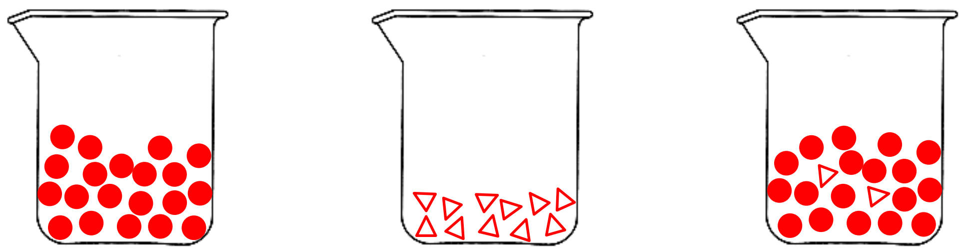
1. Define:
   a. solution: ____
   b. solvent: ____
   c. solute: ____
2. Give an example of a common solution found in everyday life in each of the following phases.
   a. gas: ____*air – oxygen dissolved in nitrogen*______________________________
   b. liquid: ____*sea water – sodium chloride dissolved in water*______________________
   c. solid: ____*14kt gold (an alloy) – silver and copper dissolved in gold*______________

Although solutions can occur for any solvent, and in any phase of matter, the most commonly encountered solutions are aqueous. Therefore, we will consider the solution process more closely for aqueous solutions.

3. In the beakers shown, use symbols to represent the appropriate compounds.

● = H_2O
△ = Ethanol C_2H_5OH

pure water　　　pure ethanol　　　10 % ethanol, 90% water



4. During this process, significant changes in the intermolecular forces of attraction occur. List the IMF's present in each of the pure compounds.
   a. pure water ____*dispersion / hydrogen bonds*____　　b. ethanol ____*dispersion / hydrogen bonds*____
5. In the solution, what IMF's are present? List what compounds are interacting and how. ____*dispersion / hydrogen bonds*
*Pure water experiences hydrogen bonding. Ethanol experiences hydrogen bonding. The mixture will also have hydrogen bonding. Hydrogen bonds in the pure liquids are being replaced by hydrogen bonds between water and ethanol.*______
6. One way to view the changes in energy associated with formation of solutions is to break the solution process down into a series of steps. The various steps are listed below. For our example of ethanol in water, circle the appropriate enthalpy change. Base your answer on the IMF's present and how these will be affected.

Step	Process	ΔH	Why?
Step 1 (ΔH_1)	Separating solvent molecules to make space available for the solute	exothermic $\boxed{\text{endothermic}}$	
Step 2 (ΔH_2)	Separating all solute molecules from each other.	exothermic $\boxed{\text{endothermic}}$	
Step 3 (ΔH_3)	Placing solute molecules in the available spaces in the solvent after step 1.	$\boxed{\text{exothermic}}$ endothermic	

The overall enthalpy change is the sum of the enthalpy changes for these three steps,

$$\Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

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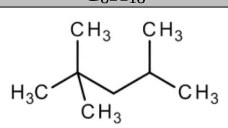
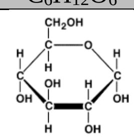
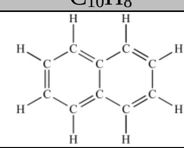
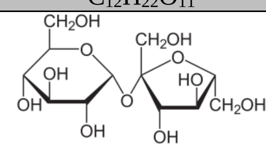
While each combination of solute and solvent will have a different value for ΔH_{soln}, as a general trend, a solution will form if ΔH_{soln} is exothermic and will not form is ΔH_{soln} is highly endothermic. Since ΔH_1 and ΔH_2 are ALWAYS endothermic, the magnitude of ΔH_3 is critical in determining solubility. If the IMF's formed between solute and solvent are of comparable magnitude to those broken, the solute will usually be soluble in that solvent. Exactly how much solute will dissolve before the solution becomes saturated is also determined by these relative strengths of IMF's. This type of analysis is often summarized in the common statement: "Like dissolves like."

7. Based upon the ideas illustrated above, complete the following table of solubilities.

Solvent		Solute		Soluble / Insoluble
Compound	polar/nonpolar?	Compound	Ionic? Molecular? (polar /nonpolar?)	
water	*polar*	KCl	*ionic*	*soluble*
water	*polar*	NH_4NO_3	*ionic*	*soluble*
water	*polar*	glucose	*polar*	*soluble*
water	*polar*	$C_{10}H_8$	*non-polar*	*insoluble*
hexane, C_6H_{14}	*non-polar*	water	*polar*	*insoluble*
hexane	*non-polar*	$C_{10}H_8$	*non-polar*	*soluble*
gasoline, C_8H_{18}	*non-polar*	sucrose (a sugar)	*polar*	*insoluble*

Gasoline (Isooctane) C_8H_{18}	Glucose $C_6H_{12}O_6$	Naphthalene $C_{10}H_8$	Sucrose $C_{12}H_{22}O_{11}$
			

8. Is the dissolving of $CaCl_2(s)$ into water endothermic or exothermic? (i.e., is ΔH_{soln} positive or negative?)

$E^o_{Lattice}$ $CaCl_2$	2258 kJ/mol
$H^o_{hydration}$ Ca^{2+}	1577 kJ/mol
$H^o_{hydration}$ Cl^-	381 kJ/mol

ΔH_{soln} = 2258 kJ/mol – (1577kJ/mol + 2(381kJ/mol)) = $\boxed{\text{-81 kJ/mol}}$

9. Is the dissolving of $NH_4Cl(s)$ into water endothermic or exothermic? (i.e., is ΔH_{soln} positive or negative?)

$E^o_{Lattice}$ NH_4Cl	708 kJ/mol
$H^o_{hydration}$ NH_4^+	307 kJ/mol
$H^o_{hydration}$ Cl^-	381 kJ/mol

ΔH_{soln} = 708 kJ/mol – (307kJ/mol + 381kJ/mol) = $\boxed{\text{20 kJ/mol}}$

Having a negative ΔH_{soln} is one factor that favors formation of solutions. However, many solutions are known to form spontaneously when ΔH_{soln} is zero or even positive. Another factor must also be involved in determining whether a solution will form. That factor is disorder or *entropy*.

10. Examine your three pictures on page 1 of this handout. Is disorder INCREASING OR DECREASING as the ethanol dissolves in the water? Explain your answer.
Because the attractions between the particles are so similar, the freedom of movement of the ethanol molecules in the water solution is about the same as their freedom of movement in the pure ethanol. The same can be said for the water. Because of this freedom of movement, both liquids will spread out to fill the total volume of the combined liquids. In this way, they will shift to the most probable, most dispersed state available, the state of being completely mixed. There are many more possible arrangements for this system when the ethanol and water molecules are dispersed throughout a solution than when they are restricted to separate layers.

11. Does entropy increase or decrease when a solid lattice breaks apart into ions? _____*increase*_______________

12. Does entropy increase or decrease when solutes are mixed into solvents? _____*increase*_______________

13. Ions with high charges (+2, +3) strongly attract a shell of hydrated water molecules. Would locking a large number of water molecules in place this way increase or decrease the entropy of the system? ____*decrease*_______________