**Ch 14 Solutions Name:**

*14.1 Thirsty Solutions: Why You Shouldn’t Drink Seawater*

Use the terms solution, solvent, and solute to describe seawater.

What is shown by *Figure 14.2*?

Why is drinking seawater dangerous?

*14.2 Types of Solutions and Solubility*

Note all the common types of solutions in *Table 14.1* – it is important to remember that solutions are \_\_\_\_\_\_geneous mixtures and may occur in any phase or combination of phases. Identify all the “phases” used to make Coca-cola.

When Ne and Ar gas are mixed *(Figure 14.3)* there are no *significant* \_\_\_\_\_\_s between particles. When the 2 gases

mix their PE remains un\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. They mix due to \_\_\_\_\_\_\_\_\_\_\_\_\_. What is entropy?

When the barrier is removed, each gas along with its \_\_\_\_\_\_\_\_\_\_\_\_\_\_ energy becomes \_\_\_\_\_\_\_\_\_\_\_\_ out or

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ed. The mixture has a greater energy \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ or greater \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ than the

separated gases. Another example of this tendency for dispersal is the \_\_\_\_\_\_\_\_\_\_\_\_\_\_ of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ energy

from hot to cold. *(Is it possible for a metal rod to spontaneously become hot at one end and cold at the other?) ­­\_\_\_\_*

*CC 14.1 \_\_\_\_*

*Figure 14.4* shows how IMFs can “promote” solution formation through attractions (remember ALL substances can be

attracted through LDFs). *Table 14.2* show that attractions between the \_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_ must be

greater than or equal to the attractions within the solute and solvent for a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ to form.

Pentane(C5H12) and heptane(C7H14) are \_\_\_\_\_\_\_\_carbons. The IMFs present between particles of the pure substance

and between pentane and heptane are \_\_\_\_\_\_s. The magnitude of the forces are similar so they are \_\_\_\_\_\_ble. The

formation of the solution is driven by a tendency to greater \_\_\_\_\_\_\_\_\_\_\_\_\_\_. When hexane (C6H14) is mixed with

water a solution does NOT form. The attractions between water molecules (especially strong \_\_ bonding) is greater

than the attraction of water molecules for hexane molecules (they would only attract by \_\_\_\_\_s). There IS a tendency

toward mixing (increased \_\_\_\_\_\_\_\_\_\_\_\_\_\_) but the energy needed to separate water molecules is too \_\_\_\_\_\_\_\_\_.

The general rule is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (polar mix with \_\_\_\_\_\_\_\_\_ or \_\_\_\_\_\_\_\_\_s and NP with NP).

For AP, you must be able to identify the reasons based on attractive forces, energy, and entropy. *CC14.2 \_\_\_\_*

*Practice 14.1 (Y/N, if yes ID the IMFs) a) \_\_\_ \_\_\_\_\_\_\_\_ b) \_\_\_ \_\_\_\_\_\_\_\_ c) \_\_\_ \_\_\_\_\_\_\_\_ d) \_\_\_ \_\_\_\_\_­­\_\_\_*

*CC 14.3 \_\_\_\_ (Note butanol & pentanol are basically insoluble in water while methanol is insoluble in hexane)*

Explain your choice.

*14.3 Energetics of Solution Formation*

When sodium \_\_\_\_\_\_\_\_\_\_\_\_\_\_ (NaOH – main ingredient in Draino) is dissolved in \_\_\_\_\_\_\_ the process is \_\_\_\_thermic.

(This is one reason why using drain cleaning chemicals can be so dangerous – especially in the solid form.)

When ammonium \_\_\_\_\_\_\_\_\_\_\_(NH4NO3 – used in “cold” packs) is dissolved in water the process is \_\_\_\_thermic.

Separating particles (solute or solvent) is always \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Mixing particles tends to be \_\_\_\_thermic as

energy is released as particles attract through \_\_\_\_\_\_s.

The main difference between “a” and “b” in *Figure 14.6* is the large \_\_\_\_thermic Δ*Hmix* step in “a” makes the “net”

overall process *(*Δ*Hsoln)* \_\_\_thermic. When water is the solvent for an ionic compound in a solution the Δ*Hsolvent* and

Δ*Hmix* terms can be combined into a single term called the heat of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (Δ*Hhydration*) because \_\_\_\_-

\_\_\_\_\_\_\_\_ forces are stronger than \_\_-bonds in water this term is always largely \_\_\_thermic for ionic compounds (*soluble ones*).

*Note how the “dissolution” process equations are written. (We can now see that dissolving is more complicated than just being a “physical” change.) Write a dissolution equation for magnesium nitrate (*Δ*Hsoln =-90.0kJ/mol).*

*Remember to balance the charges in the compound and to balance the equation.*

When Δ*Hsolute <*Δ*Hhydration* energy required to separate the solute into ions is \_\_\_\_\_\_than the energy given \_\_\_\_\_ as

the ions are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Δ*Hsoln* is \_\_\_\_thermic (solution feels \_\_\_\_\_\_).

When Δ*Hsolute >*Δ*Hhydration* energy required to separate the solute into ions is \_\_\_\_\_\_\_\_than the energy given \_\_\_\_\_ as

the ions are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. Δ*Hsoln* is \_\_\_\_thermic (solution feels \_\_\_\_\_\_).

When Δ*Hsolute ≈* Δ*Hhydration* then Δ*Hsoln* ≈ 0. Which 2 halides are examples? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

*14.4 Solution Equilibrium and Factors Affecting Solubility*

Dispute the following statement. When NaCl is added to water it eventually stops dissolving. *(Examine Figure 14.9)*

How does the equation for an unsaturated vs a saturated solution of NaCl differ? Why?

The solution in the first frame of *Figure 14.10* is \_\_\_\_\_\_\_saturated. Can you tell by just looking at it?

This solution was made by heating the solvent (water) and then adding a large amount of sodium \_\_\_\_\_\_\_\_\_\_\_\_.

Then the hot solution is allowed to cool undisturbed so that \_\_\_\_\_\_\_\_\_\_\_\_tion does not occur.

*(The best way to “test” if a solution is unsat, sat, or supersat is by adding a small crystal to see what occurs.)*

What does *Figure 14.11* show about the relationship between solubility and temperature for most solids?

Which solid does not follow the trend? \_\_\_\_\_\_\_\_\_\_\_\_\_\_ Which solid does not show much change? ­­­­\_\_\_\_\_\_\_\_

How does temperature affect the solubility of gases? *(What happens to soda that is warm?)*

*CC14.5 \_\_\_\_*

Why does this make sense in terms of KMT and IMFs?

Henry’s \_\_\_\_\_\_ basically states that the solubility of a gas (Sgas) is directly \_\_\_\_\_\_\_\_\_\_\_\_\_al to the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ on the gas above the liquid. What should happen to the Sgas in a liquid if the pressure on the gas is doubled?

*CC14.6 \_\_\_\_*

*14.5 Expressing Solution Concentration*

Dilute and concentrated are “\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_” terms of concentration.

Read *Chemistry in the Environment: Lake Nyos*

*You should be familiar with the solution concentration terms in Table 14.5. Molality is not tested in AP.*

*The process in Figure 14.13 of preparing a certain molarity concentration solution is required knowledge by AP and is often part of FRQ questions.* Be FAMILIAR with the process. The flask in the image is called a volumetric flask – it has a line on the neck which precisely marks a volume of 1L. Why do you not want to fill the flask with the water (solvent) first and then add the NaCl (solute)?

Molality (m) is often used when solution are heated. The *molarity* of a solution will change slightly due to expansion and contraction with change in temperature where the *molality* will not since it is based on mass not volume.

*Practice 14.3*

*Practice 14.4 a, b, c, d (not e)*

Read *Chemistry in the Environment: The Dirty Dozen* The concentration of mg/L used in *Table 12.7* is basically the same as ppm (since mg/L is mg/mL which is 1/1000 out of 1/1000 or 1ppm). Dioxins and PCBs were prevalent pollutants in this area. *(I remember my father using the insecticide chlordane.)* *Practice 14.5*

*14.6 Colligative Properties: Vapor Pressure Lowering, Freezing Point Depression, Boiling Point Elevation, and Osmotic Pressure* *(NOTE: AP does not test these concepts but they are considered previous knowledge)*

The 4 colligative properties depend on the \_\_\_\_\_\_\_\_\_\_\_\_\_ of particles dissolved in \_\_\_\_\_\_\_\_\_\_ NOT the \_\_\_\_\_\_\_ of

particle. When 1mol of NaCl (an \_\_\_\_\_\_\_\_lyte) is dissolved in water the solution has \_\_\_moles of dissolved particles.

When a \_\_\_volatile solute is dissolved in a \_\_\_vent, the vapor pressure of the solution is \_\_\_\_\_\_ that that of the \_\_\_\_\_

solvent. An explanation based on entropy is on pg 602. You can also think of the solute particles in the solution

attracting the solvent particle (IMFs) preventing them from leaving the solution. \_\_\_\_\_\_\_\_\_ law shows that the VP of

solution can be calculated by multiplying the mole fraction of the \_\_\_\_\_\_\_\_\_\_\_\_ times the VP of the pure \_\_\_\_\_\_\_\_\_.

Notice in the example that a solution that is 90% solvent will have a VP that is \_\_\_% of the pure solvent *(or since the*

*solution is 10% solute the VP of the solvent is lowered by 10%)*. *Practice 14.6*

Pgs 605-7 show the effects of a volatile solute in a volatile solvent. *Figure 14.15 shows the different possibilities.*

*In an ideal solution “a” the IMFs for the solute and solvent are similar in magnitude. “b” shows how the VP of the solution is lower than ideal since the components have strong attractions and “c” shows how the VP is higher when the attractions are weak.*

The phase diagram on pg 608 shows that the mp of a solution is \_\_\_\_er than that of the pure solvent and the bp of the

solution is \_\_\_\_\_\_er than that of the pure solvent. *(Again particle attractions disrupt these phase changes.)*

Give 2 reasons why antifreeze is used in a car. Why does ice formed from saltwater not taste salty? (after rinsing)

*Practice 14.8 CC14.10 \_\_\_\_*

Review *Example 14.9* and complete *Practice 14.9*

Read *Chemistry in Your Day: Antifreeze in Frogs* Answer the question.

Why does the water level on the left side of the osmotic cell rise in *Figure 14.16*?

If external \_\_\_\_\_\_\_\_\_\_\_\_\_\_ is applied to the water in the left cell, this process can be opposed *(reverse osmosis)*.

*Practice 14.10*

*14.7 Colligative Properties of Strong Electrolytes*

Why does NaCl have and van’t Hoff factor that would ideally be 2?

Why would the value be closer to the ideal when the solution is very dilute?

What causes the measure values to be lower than ideal *(Figure 14.9)*

Notice how the factor (i) is added to the formulas.

*CC 14.7 \_\_\_\_* Explain your answer.

*Practice 42.11*

Why must IV solutions be isomotic (or isotonic)?

*14.8 Colloid*

What is a colloid? What determines whether a mixture will be a colloid, solution, or suspension?

Note the examples in *Table 14.10*. Colloids are often called dispersions or emulsions. Egg helps to emulsify the oil and vinegar (water) to make mayonnaise.

*Figure 14.22* shows how micelles form when soap molecules emulsify oily dirt. (Does the image look familiar?)

In *Figure 14.23* when particles \_\_\_\_\_\_\_\_\_er some of the light you see the \_\_\_\_\_\_\_\_\_\_\_\_\_\_ effect.

*Exercises* (pgs 623-626) #34, 35, 41, 43, 47, 51, 53, 58, 59, 65, 71, 77, 81, 83, 85, 87b, 91a, 93, 106, 112