Practice Essay Questions

Answer 1-4 and 6 in the form of a paragraph. Include an introductory sentence a topic (claim) sentence, good grammar, correct use of terminology, and a concluding sentence. Explain clearly and give clear examples to back up your claims.

1. Explain how vapor pressure, intermolecular forces, surface area, and temperature influence evaporation. Be sure to define each term as it relates to the question. Pg. 444-446.

2. An open container of alcohol and a separate open container of the same volume of water are at 30 $_{o}$C and 1.1 atmospheres. Which liquid will evaporate first? Use the terms vapor pressure, temperature, intermolecular force, kinetic energy correctly in your answer.

Not all liquids evaporate the same way. If there is one container each of alcohol and another of water at the same temperature, pressure, volume and surface area, the alcohol will evaporate first. Since both have the same temperature then they both have the same average kinetic energy. But the intermolecular forces of the two liquids is not the same. The forces of attraction between water molecules are stronger than the forces of attraction between different alcohol molecules. Therefore it is takes less kinetic energy to overcome the attractive forces between alcohol molecules. In this case both liquids are at the same temperature, so alcohol will evaporate more quickly. At the temperature of 30 degrees it will be easier for alcohol molecules to overcome the attractive forces between them and escape into the atmosphere. Alcohol has a higher vapor pressure than water at the same temperature.

3. Describe 2 ways to make a liquid boil. Explain why each method works. Define your terms. Pg.- 450-451

Boiling occurs when the vapor pressure of a liquid equals the atmospheric pressure over the liquid. Therefore there are two ways to make a liquid boil: either increase the vapor pressure or decrease the atmospheric pressure.

The first way to make a liquid evaporate is to increase its vapor pressure. Molecules of a liquid are attracted to each other with inter molecular forces. These forces prevent the molecules at the surface from vaporizing and becoming a liquid. If a molecule achieves sufficient kinetic energy at the surface it will overcome the intermolecular forces and escape into the vapor phase. If the molecules of the liquid are heated and reach enough kinetic energy, they will form bubbles of vapor inside the liquid and reach boiling. At this point the vapor pressure equals the atmospheric pressure.

The other way to achieve boiling is to lower the atmospheric pressure. If the outside pressure on the liquid is lowered then the liquid's current kinetic energy will be sufficient to go from liquid phase to vapor phase. If the atmospheric pressure is lowered enough it will eventually reach equality with the vapor pressure of the liquid and boiling will occur. To make boiling happen either the atmospheric pressure must be decreased or the vapor pressure must be increased until the two are equal.

4. Explain what happens to the temperature of a liquid as it evaporates. Be sure to correctly use the terms kinetic energy, temperature, molecular motion. Pg. 445

As a liquid evaporates the temperature of that liquid decreases. All molecules of a liquid have kinetic energy which is directly related to the motion of the molecules. Not all molecules have the same kinetic energy. The temperature of a liquid is the average kinetic energy of all of the molecules. A liquid becomes a vapor when a molecule achieves sufficient kinetic energy to overcome the forces of attraction between it and surrounding molecules. When a liquid evaporates it is the molecules of the highest kinetic energy that escape. When those molecules of the highest kinetic energy have escaped they leave behind the molecules with lower kinetic energy thus lowering the average kinetic energy of those molecules still in the liquid phase. The temperature of a liquid is directly related to the average kinetic energy of the molecules. Since the average kinetic energy of those molecules still in the liquid is lower the temperature is lower.

5. Draw the phase diagram of a substance with the following properties. Label theaxes. Show that the normal boiling point is 85$_o$ C and the normal melting point is -25$_o$C. Show that the triple point is at – 45 $_o$C and .4 atm and that the critical temperature is at 125 $_o$C and 1.5 atm. Label the regions of solid, liquid and gas. Label a point at which the rate of sublimation equals the rate of crystallization. Label a point at which the liquid and gas phases are in equilibrium. Which phase is more dense, the liquid phase or the solid phase. How can you tell.

6. What is the difference between an ideal and a real gas. Define each. Give examples, evidence, and clear explanations to back up any claims.

Models are used in chemistry to help us understand real life situations. Ideal gases are a model that helps us understand how real gases work. Although ideal gases are useful, they deviate from real gases under two conditions: high pressure and low temperature.

First, the model for ideal gases assumes that there is no attraction between the molecules. This approximation works for conditions of high temperature. In reality there is some attraction between the molecules of a gas. The attraction is negligible at high temperatures. Under the ideal gas model, when two particles get close to each other, there is no attraction and no change of direction. This means that the particles of an ideal gas always travel in straight lines. In reality, the particles do attract each other a little bit. As real gas particles come close to each other, the small attractive force causes a slight curvature of their path. This means that the distance they actually travel is slightly longer than the distance that they would travel without the attraction. Consequently the number of collisions with the walls of the container is decreased, and therefore the pressure predicted under ideal conditions is greater than under real conditions. This difference is unnoticeable under high temperatures such as room temperature. But at low temperatures, this deviation becomes larger and the decrease in pressure is significant. Also, most gases at low temperatures have such a high attraction that they become liquids and no longer behave as gases at all.

Secondly, the model of ideal gases assumes that the particles have no volume themselves. This means that the volume predicted by the ideal gas equation is equal to only the empty space between the molecules. But, the real gas volume is the sum of the molecules and the space between them. At low pressures the distance between the particles of a real gas are so large that the actual molecular volume is insignificant. But as the pressure on the gas increases, causing the space between the molecules to become less and less, the proportion of the space taken up by the molecules becomes greater. At very high pressures the space occupied by the molecules is no longer negligible relative to the empty space. Thus the volume predicted by the ideal gas equation is less than the volume of a real gas.

The ideal gas equation is useful for predicting the behavior of real gases. But at low temperature and high pressure the model no longer approximates reality. This is because the ideal gas equation assumes the molecules have no volume and that they have no attraction to each other. At extreme conditions, ideal gases deviate from reality.