***Chemistry***

**10: Liquids and Solids**

**10.3: Phase Transitions**

31. Heat is added to ice at 0 °C. Explain why the temperature of the ice does not change. What does change?

Solution

The heat is absorbed by the ice, providing the energy required to partially overcome intermolecular attractive forces in the solid and causing a phase transition to liquid water. The solution remains at 0 °C until all the ice is melted. Only the amount of water existing as ice changes until the ice disappears. Then the temperature of the water can rise.

33. Identify two common observations indicating some liquids have sufficient vapor pressures to noticeably evaporate?

Solution

We can see the amount of liquid in an open container decrease and we can smell the vapor of some liquids.

35. What is the relationship between the intermolecular forces in a liquid and its vapor pressure?

Solution

The vapor pressure of a liquid decreases as the strength of its intermolecular forces increases.

37. Why does spilled gasoline evaporate more rapidly on a hot day than on a cold day?

Solution

As the temperature increases, the average kinetic energy of the molecules of gasoline increases and so a greater fraction of molecules have sufficient energy to escape from the liquid than at lower temperatures.

39. When is the boiling point of a liquid equal to its normal boiling point?

Solution

When the pressure of gas above the liquid is exactly 1 atm.

41. Use the information in Figure 10.23 to estimate the boiling point of water in Denver when the atmospheric pressure is 83.3 kPa.

Solution

Follow an imaginary horizontal line at 83.3 kPa to the curve representing the vapor pressure of water. Then drop a vertical line to the temperature axis. The intersection is at approximately 95 °C.

43. Explain the following observations:

(a) It takes longer to cook an egg in Ft. Davis, Texas (altitude, 5000 feet above sea level) than it does in Boston (at sea level).

(b) Perspiring is a mechanism for cooling the body.

Solution

(a) At 5000 feet, the atmospheric pressure is lower than at sea level, and water will therefore boil at a lower temperature. This lower temperature will cause the physical and chemical changes involved in cooking the egg to proceed more slowly, and a longer time is required to fully cook the egg. (b) As long as the air surrounding the body contains less water vapor than the maximum that air can hold at that temperature, perspiration will evaporate, thereby cooling the body by removing the heat of vaporization required to vaporize the water.

45. Explain why the molar enthalpies of vaporization of the following substances increase in the order CH4 < C2H6 < C3H8, even though all three substances experience the same dispersion forces when in the liquid state.

Solution

Dispersion forces increase with molecular mass or size. As the number of atoms composing the molecules in this homologous series increases, so does the extent of intermolecular attraction via dispersion forces and, consequently, the energy required to overcome these forces and vaporize the liquids.

47. The enthalpy of vaporization of CO2(*l*) is 9.8 kJ/mol. Would you expect the enthalpy of vaporization of CS2(*l*) to be 28 kJ/mol, 9.8 kJ/mol, or –8.4 kJ/mol? Discuss the plausibility of each of these answers.

Solution

The boiling point of CS2 is higher than that of CO2 partially because of the higher molecular weight of CS2; consequently, the attractive forces are stronger in CS2. It would be expected, therefore, that the heat of vaporization would be greater than that of 9.8 kJ/mol for CO2. A value of 28 kJ/mol would seem reasonable. A value of –8.4 kJ/mol would indicate a release of energy upon vaporization, which is clearly implausible.

49. Ethyl chloride (boiling point, 13 °C) is used as a local anesthetic. When the liquid is sprayed on the skin, it cools the skin enough to freeze and numb it. Explain the cooling effect of liquid ethyl chloride.

Solution

The thermal energy (heat) needed to evaporate the liquid is removed from the skin.

51. How much heat is required to convert 422 g of liquid H2O at 23.5 °C into steam at 150 °C?

Solution



Heat needed to bring this amount of water to the normal boiling point: Δ*H1* = mCsΔT = (422 g)(4.184 J/g °C)(100.0 – 23.5) = 135,000 J

Heat needed to vaporize this amount of water: Δ*H*2 = *n*Δ*H*vap = (23.4 mol)(40,650 J/mol) = 951,000 J

Heat to needed to increase the temperature of the steam: Δ*H*3 = mCsΔT = (422 g)(2.09 J/g °C)(150 – 100) = 44,100 J.

Adding Δ*H*1, Δ*H*2, and Δ*H*3: 135,000 J + 951,000 J + 44,100 J = 1,130,000 J = 1130 kJ.

53. Titanium tetrachloride, TiCl4, has a melting point of –23.2 °C and has a Δ*H* fusion = 9.37 kJ/mol.

(a) How much energy is required to melt 263.1 g TiCl4?

(b) For TiCl4, which will likely have the larger magnitude: Δ*H*fusion or Δ*H*vaporization? Explain your reasoning.

Solution

(a) . Heat required to melt this amount of TiCl4 is *n*Δ*H*fusion = 1.385 mol  9.37 kJ/mol = 13.0kJ. (b) It is likely that the heat of vaporization will have a larger magnitude since in the case of vaporization the intermolecular interactions have to be completely overcome, while melting weakens or destroys only some of them.

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