

**Problem # 1** Determine the kinds of intermolecular forces present in each element or compound:

- A. Kr
- B.  $\text{NCl}_3$
- C.  $\text{SiH}_4$
- D. HF

SOLUTION: Kr is a single atom, hence it can have no permanent dipole. The only possible force is a (weak) van der Waal's force.

$\text{NCl}_3$  will have van der Waal's force (everything does). The other forces depend on the existence of a dipole, so I look at the electronegativities. N is 3.0 and so is Cl. So, no electronegativity difference, no permanent dipole.

Same basic analysis for  $\text{SiH}_4$ . It has van der Waal's forces. The electronegativity of Si 1.8 and H is 2.1. That is below 0.5 difference ( $2.1-1.8$ ) so the bond itself is considered to be non-polar so no dipole forces.

HF has van der Waal's forces. The electronegativity of F is 4.0 compared to H 2.1. This is a big difference, almost ionic. H bonded to F, O, N causes "hydrogen bonding" which is just really strong dipole/dipole interaction.

**Problem # 2** A 45.0 mL sample of water is heated to its boiling point. How much heat is required (in kJ) to vaporize it? (Assume a density of 1.0 g/mL.)

SOLUTION: The energy required to make the phase transition is the  $\Delta H_{\text{vap}}$ . For water at  $100^\circ\text{C}$ ,  $\Delta H_{\text{vap}}=40.7 \text{ kJ/mol}$ . Units! Units! Units! I have  $\frac{\text{kJ}}{\text{mol}}$  and I want kJ. All I need to know is how many moles of water I have and I'm done!  $\frac{\text{kJ}}{\text{mol}} \text{ mol} = \text{kJ}$

I don't know moles, I do know mL...that's okay, I just convert the units!

$$45.0 \text{ mL} \frac{1.0 \text{ g}}{\text{mL}} \frac{1 \text{ mol}}{18.02 \text{ g}} = 2.50 \text{ mol } \text{H}_2\text{O}$$

$$\frac{40.7 \text{ kJ}}{\text{mol}} 2.50 \text{ mol} = 102 \text{ kJ}$$

**Problem #3** Suppose that 1.02 g of rubbing alcohol ( $C_3H_8O$ ) evaporates from a 55.0 g aluminum block. If the aluminum block is initially at  $25^\circ C$ , what is the final temperature of the block after the evaporation of the alcohol? Assume that the heat required for the evaporation of the alcohol comes only from the aluminum block and that the alcohol vaporizes at  $25^\circ C$ .

SOLUTION: This is really just a conservation of energy problem. The energy moves around but the total amount stays the same. The total change in energy for the system is 0. Or, in algebraic form:

$$Q_{\text{block}} + Q_{\text{alcohol}} = 0$$

The last sentence gives you two simplifying assumptions. "the heat...comes only from the aluminum block" is telling you that the air isn't part of the problem so you don't need to worry about other things in the room. The "alcohol vaporizes at  $25^\circ C$ " is informing you that the alcohol doesn't change temperature during the process.

If I consider my two heat terms:

$Q_{\text{alcohol}}$  is simply a change of state for the alcohol, so it is a  $n\Delta H_{\text{vap}}$ .

$Q_{\text{block}}$  is a temperature change of the block (no phase change) so it is an  $mc\Delta T$ .

$$m_{\text{block}}c_{\text{Al}}\Delta T_{\text{block}} + n_{\text{alcohol}} \Delta H_{\text{vap,alcohol}} = 0$$

or

$$m_{\text{block}}c_{\text{Al}}\Delta T_{\text{block}} = - n_{\text{alcohol}} \Delta H_{\text{vap,alcohol}}$$

Recall that  $\Delta T = T_f - T_i$

$$m_{\text{block}}c_{\text{Al}}(T_{f,\text{block}} - T_{i,\text{block}}) = - n_{\text{alcohol}} \Delta H_{\text{vap,alcohol}}$$

Let's look at what we know:

$$m_{\text{block}} = 55.0 \text{ g}$$

$$m_{\text{alcohol}} = 1.02 \text{ g}$$

$$T_{i,\text{block}} = 25^\circ C$$

I can also look up the two physical constants:

$$c_{\text{Al}} = 0.903 \text{ J}/(\text{g}^\circ C)$$

$$\Delta H_{\text{vap, rubbing alcohol}} = 45.4 \text{ kJ/mol (at } 25^\circ C)$$

I actually need moles of alcohol, not grams but if I look at my conservation of energy equation, I know everything but the  $T_f$  of the aluminum block.

$$1.02 \text{ g alcohol} \frac{1 \text{ mol}}{60.097 \text{ g}} = 0.01697 \text{ mol alcohol}$$

$$55.0 \text{ g Al} \frac{0.903 \text{ J}}{\text{g}^\circ\text{C}} (T_f - 25^\circ\text{C}) = -0.01697 \text{ mol alcohol} \frac{45.4 \times 10^3 \text{ J}}{\text{mol}}$$

(notice that I converted the kJ to J so that I have J on both sides of the equation.

$$49.665 T_f - 1241.7 = -770.5543$$

$$49.665 T_f = 471.07$$

$$T_f = 9.5^\circ\text{C}$$

**Problem #4** Benzene has a heat of vaporization of 30.72 kJ/mol and a normal boiling point of 80.1 °C. At what temperature does benzene boil when the external pressure is 445 torr.

SOLUTION: Just a Clausius-Clapeyron equation

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

It is important to remember what “normal” means. The “normal boiling point” is the boiling point at 1 atm (760 torr) of pressure. And recall that “boiling” occurs when the vapor pressure = atmospheric pressure.

So, I actually know more than it looks like at first glance.

$$P_1 = 760 \text{ torr}$$

$$T_1 = 80.1^\circ\text{C}$$

$$P_2 = 445 \text{ torr}$$

$$T_2 = ?$$

$$\Delta H_{vap} = 30.72 \text{ kJ/mol}$$

$$R = 8.314 \text{ J/(mol K)}$$

It is very important to have the units correct in a problem such as this.

The pressure units don't matter. They cancel in the ratio and all pressure units are proportional.

Since R has “Joules” in it, I need  $\Delta H_{vap}$  to be in J so it cancels.

The other big issue is temperature. EVERY time you are talking energy, temperature must be on an absolute scale. You need zero temperature to correspond to zero energy. So, kelvins is it.

$$P_1 = 760 \text{ torr}$$

$$T_1 = 80.1^\circ\text{C} + 273.15 = 353.25 \text{ K}$$

$$P_2 = 445 \text{ torr}$$

$$T_2 = ?$$

$$\Delta H_{\text{vap}} = 30.72 \text{ kJ/mol} = 30.72 \times 10^3 \text{ J/mol}$$

$$R = 8.314 \text{ J/(mol K)}$$

Now it's just a plug and chug!

$$\ln\left(\frac{445 \text{ torr}}{760 \text{ torr}}\right) = \frac{-30.72 \times 10^3 \frac{\text{J}}{\text{mol}}}{8.314 \frac{\text{J}}{\text{mol K}}} \left(\frac{1}{T_2} - \frac{1}{353.25 \text{ K}}\right)$$

$$\ln(0.5855) = -3695 \text{ K} \left(\frac{1}{T_2} - \frac{1}{353.25 \text{ K}}\right)$$

$$-0.5353 = -3695 \text{ K} \left(\frac{1}{T_2} - \frac{1}{353.25 \text{ K}}\right)$$

$$1.449 \times 10^{-4} = \left(\frac{1}{T_2} - \frac{1}{353.25 \text{ K}}\right)$$

$$1.449 \times 10^{-4} = \frac{1}{T_2} - 1.4487 \times 10^{-4}$$

$$\frac{1}{T_2} = 2.976 \times 10^{-3}$$

$$T_2 = 336 \text{ K}$$

**Problem #5** How much heat (in kJ) is evolved in converting 1.00 mol of steam at 145.0 °C to ice at -50 °C? The heat capacity of steam is 2.09 J/g°C and of ice is 2.09 J/g°C.

SOLUTION: There's actually a typo here. The values for the heat capacity of steam and ice should NOT be the same. Steam is actually less. But I'll just solve it using the numbers they give you.

For ALL heating/cooling problems, you need to break up the energy journey of the substance into two different types of legs: keeping the phase and changing temperature or changing phase and keeping the temperature constant.

We start with 1.00 mol of steam at 145 C. This is above the normal boiling point of 100 C. So the first thing that happens if I start taking out energy is that the steam cools down to its boiling (condensing) point. This is an  $mC\Delta T$  problem.

$$1.00 \text{ mol } H_2O \frac{18.02 \text{ g}}{\text{mol } H_2O} = 18.02 \text{ g}$$

$$Q = mC_{\text{steam}}\Delta T$$

$$Q = (18.02 \text{ g}) \left( 2.09 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (100^\circ\text{C} - 145^\circ\text{C}) = -1694 \text{ J}$$

The (-) sign just means that the heat is given off rather than absorbed.

Once I get to the boiling point, I can condense into liquid form. This is an  $n\Delta H$  problem. The heat of vaporization at 100 C is 40.7 kJ/mol. Since we are condensing not boiling, the enthalpy of condensation is -40.7 kJ/mol

$$Q = n\Delta H = 1.00 \text{ mol} \left( \frac{-40,700 \text{ J}}{\text{mol}} \right) = -40,700 \text{ J}$$

Now I'm liquid water. I need to cool down to my freezing point before I'll freeze. This is an  $mC\Delta T$  problem again.

$$Q = mC_{\text{steam}}\Delta T$$

$$Q = (18.02 \text{ g}) \left( 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (0^\circ\text{C} - 100^\circ\text{C}) = -7532 \text{ J}$$

Now I'm at my freezing point, so I freeze!

$$Q = n\Delta H = 1.00 \text{ mol} \left( \frac{-6020 \text{ J}}{\text{mol}} \right) = -6020 \text{ J}$$

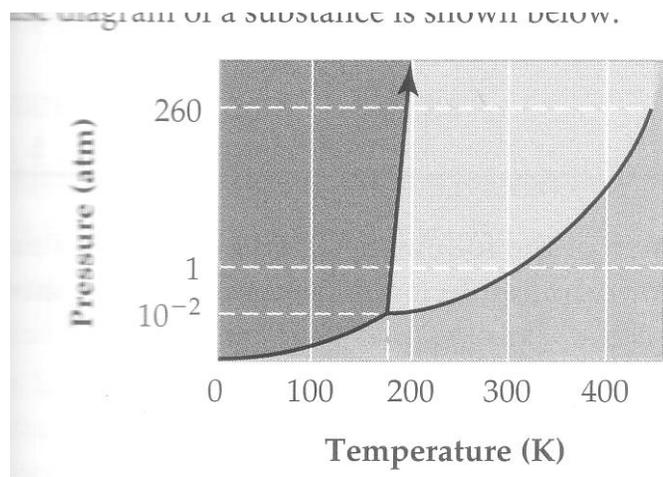
Now I'm ice! I just cool down to my final temperature.

$$Q = (18.02 \text{ g}) \left( 2.09 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (-50^\circ\text{C} - 0^\circ\text{C}) = -1883 \text{ J}$$

The total heat given off is just the sum of all the legs of the journey:

$$Q_{\text{total}} = -1694 \text{ J} + (-40,700 \text{ J}) + (-7532 \text{ J}) + (-6020 \text{ J}) + (-1883 \text{ J}) = -57,829 \text{ J}$$

**Problem #6** Consider the phase diagram below and answer each question.



- What is the normal boiling point?
- What is the normal melting point?
- What is the physical state of the substance when  $T=150\text{K}$  and  $P=0.5 \text{ atm}$ ?
- What is the physical state of the substance when  $T=325\text{K}$  and  $P=0.9 \text{ atm}$ ?
- What is the physical state of the substance when  $T=450\text{K}$  and  $P=265 \text{ atm}$ ?

Solution: Normal boiling/melting point is just the phase change temperature at 1.0 atm. So, reading across the dotted 1 atm line, the melting point is at about 180K (where it crosses the first dark black line) and the boiling point is about 305 K (where it crosses the second dark black line).

$T=150 \text{ K}$  and  $P=0.5 \text{ atm}$  puts it in the darkest gray area which is a solid.  $T=325 \text{ K}$  and  $P=0.9 \text{ atm}$  puts it in the middle gray area which is a gas.  $T=450 \text{ K}$  and  $P = 265$  appears to be in the supercritical fluid region [essentially a very dense gas].

**Problem #7** An aqueous  $\text{KNO}_3$  solution is made using 72.3 g of  $\text{KNO}_3$  diluted to a total solution volume of 1.50 L. Calculate the molarity, molality, and mass percent of the solution. (Assume a density of 1.05 g/mL for the solution.)

SOLUTION: Just a series of unit conversions.

$$\frac{72.3 \text{ g } KNO_3}{1.50 \text{ L solution}} \frac{1 \text{ mol } KNO_3}{101.08 \text{ g } KNO_3} = \frac{0.477 \text{ mol } KNO_3}{L} = 0.477M$$

$$\begin{aligned} \frac{0.477 \text{ mol } KNO_3}{L \text{ solution}} \frac{1 L}{1000 \text{ mL}} \frac{1 \text{ mL solution}}{1.05 \text{ g}} &= \frac{4.54 \times 10^{-4} \text{ mol } KNO_3}{\text{g solution}} \frac{101.08 \text{ g } KNO_3}{\text{mol}} \\ &= \frac{0.0459 \text{ g } KNO_3}{\text{g solution}} \times 100 = 4.59\% \text{ by mass} \end{aligned}$$

Remember, molality is in terms of kg of solvent not kg of solution. So to get to molality, I need to subtract the mass of the solute from the total mass of solution. I can start with the Molarity which tells me that there are 0.477 moles in 1 L of solution.

$$1 \text{ L solution} \frac{1000 \text{ mL}}{1 L} \frac{1.05 \text{ g}}{1 \text{ mL solution}} = 1050 \text{ g solution}$$

So, 1 L of solution weighs 1050 g. That is water and  $KNO_3$ . 1 L of solution contains 0.477 moles of  $KNO_3$  which has a mass of:

$$0.477 \text{ mol } KNO_3 \frac{101.08 \text{ g } KNO_3}{\text{mol } KNO_3} = 48.2 \text{ g } KNO_3$$

So, the actual mass of the solvent is:

$$1050 \text{ g solution} - 48.2 \text{ g } KNO_3 = 1002 \text{ g water (solvent)}$$

$$\frac{0.477 \text{ mol } KNO_3}{1.002 \text{ Kg solvent}} = 0.476 \text{ molal}$$

**Problem #8** To what volume should you dilute 125 mL of an 8.00 M  $\text{CuCl}_2$  solution so that 50.0 mL of the diluted solution contains 5.9 g  $\text{CuCl}_2$ ?

SOLUTION: There's a number of different ways to do this. Whatever method you use requires the units of the undiluted solution to be the same as the diluted solution.

$$\frac{5.9 \text{ g CuCl}_2}{50.0 \text{ mL}} \frac{1000 \text{ mL}}{1 \text{ L}} \frac{1 \text{ mol CuCl}_2}{134.36 \text{ g}} = 0.878 \text{ M}$$

So, I want the diluted solution to have a concentration of 0.878 M

$$M_1V_1 = M_2V_2$$

$$8.00 \text{ M} (125 \text{ mL}) = (0.878 \text{ M})(x \text{ mL})$$

$$x = 1139 \text{ mL}$$

**Problem #9** Lead is a toxic metal that affects the central nervous system. A Pb-contaminated water sample contains 0.0011% Pb by mass. How much of the water (in mL) contains 150 mg of Pb? (Assume a density of 1.0 g/mL.)

SOLUTION:

$$150 \text{ mg Pb} \frac{100 \text{ mg solution}}{0.0011 \text{ mg Pb}} \frac{1 \text{ g}}{1000 \text{ mg}} \frac{1 \text{ mL}}{1.0 \text{ g}} = 13,636 \text{ mL}$$

**Problem #10** A solution contains naphthalene ( $\text{C}_{10}\text{H}_8$ ) dissolved in hexane ( $\text{C}_6\text{H}_{14}$ ) at a concentration of 10.85% naphthalene by mass. Calculate the vapor pressure at 25 °C of hexane above the solution. The vapor pressure of pure hexane at 25 °C is 151 torr.

SOLUTION: Vapor pressure of a solution means, say "Hola" to Raoult!

$$P_{vap} = \chi_{solvent} P_{vap}^0$$

We have the concentration in %by mass, we need to convert this to mole fraction

$$\frac{10.85 \text{ g naphthalene}}{100 \text{ g solution}}$$

This means that every 100 g of solution contains 10.85 g naphthalene and 89.15 (100 g – 10.85 g) of hexane.

$$10.85 \text{ g naphthalene} \frac{1 \text{ mol}}{128.2 \text{ g}} = 0.08465 \text{ mol naphthalene}$$

$$89.15 \text{ g hexane} \frac{1 \text{ mol}}{86.18 \text{ g}} = 1.034 \text{ mol hexane}$$

$$P_{vap} = \chi_{solvent} P_{vap}^0$$

$$P_{vap} = \frac{1.034 \text{ mol hexane}}{(1.034 \text{ mol hex} + 0.08465 \text{ mol naph})} (151 \text{ torr})$$

$$P_{vap} = (0.9243)(151 \text{ torr}) = 140 \text{ torr}$$

**Problem #11** An ethylene glycol solution contains 21.2 g of ethylene glycol (C<sub>2</sub>H<sub>6</sub>O<sub>2</sub>) in 85.4 mL of water. Calculate the freezing point and boiling point of the solution. (Assume a density of 1.00 g/mL for water.)

SOLUTION: It's all about molality here

$$\frac{21.2 \text{ g ethylene glycol}}{85.4 \text{ mL water}} \frac{1 \text{ mL water}}{1.00 \text{ g water}} \frac{1000 \text{ g}}{1 \text{ kg water}} \frac{\text{mol eth glycol}}{62.07 \text{ g}} = \frac{4.00 \text{ mol}}{\text{kg solvent}}$$

$$\Delta T_b = K_b \text{ molality} = 0.512 \frac{^\circ\text{C}}{\text{molal}} (4.00 \text{ molal}) = 2.04^\circ\text{C}$$

$$\Delta T_b = (T_{b,\text{solution}} - T_{b,\text{water}}) = (T_{b,\text{solution}} - 100^\circ\text{C}) = 2.04^\circ\text{C}$$

$$T_{b,\text{solution}} = 2.04^\circ\text{C} + 100^\circ\text{C} = 102.4^\circ\text{C}$$

$$\Delta T_f = K_f \text{molality} = 1.86 \frac{^\circ\text{C}}{\text{molal}} (4.00 \text{ molal}) = 7.44^\circ\text{C}$$

$$\Delta T_f = (T_{f,\text{water}} - T_{f,\text{solution}}) = (0^\circ\text{C} - T_{f,\text{solution}}) = 7.44^\circ\text{C}$$

$$T_{f,\text{solution}} = 0^\circ\text{C} - 7.44^\circ\text{C} = -7.44^\circ\text{C}$$

**Problem #12** An aqueous solution containing 35.9 g of an unknown molecular (nonelectrolyte) compound in 150.0 g of water was found to have a freezing point of  $-1.3^\circ\text{C}$ . Calculate the molar mass of the unknown compound.

SOLUTION:

$$\Delta T_f = 1.3^\circ\text{C} = K_f \text{molality} = 1.86 \frac{^\circ\text{C}}{\text{molal}} (X \text{ molal})$$

$$X = \frac{1.3^\circ\text{C}}{1.86 \frac{^\circ\text{C}}{\text{molal}}} = 0.699 \text{ molal} = \frac{0.699 \text{ mol}}{\text{kg water}}$$

$$\frac{0.699 \text{ mol}}{\text{kg water}} 0.150 \text{ kg water} = .105 \text{ mol}$$

$$\frac{35.9 \text{ g}}{.105 \text{ mol}} = 342 \text{ g/mol}$$