**Thermodynamics Unit - Practice Thermodynamics problems**

*True/False*

- **T  F** For an isothermal process, $\Delta S_{sys}$ can never decrease.
  
  **False.** For example vaporization.

- **T  F** For all phase transitions, $\Delta H = 0$
  
  **False.** $\Delta H$ is never zero for a phase transition.

- **T  F** A process that doubles the number of microstates of system will double the entropy of the system.
  
  **False.** Entropy is proportional to the natural log of the number of microstates.

- **T  F** Dropping an eraser from a height of three feet to the floor leads to an increase in the entropy of the Universe.
  
  **True.** This is a spontaneous process.

- **T  F** The standard entropy of an element in its standard state at 298.15 K and 1 bar is zero.
  
  **False.** The standard entropy of a perfect crystal of a substance at absolute zero is zero.

- **T  F** Conservation of energy tells that $\Delta U = 0$ for all processes.
  
  **False.** Energy can be exchanged between system and surroundings. $\Delta U_{Universe} = 0$.

- **T  F** If adding 25 J of heat to a 5.6 g block of iron increases it temperature by 10° C, then adding 25 J of heat to a 2.8 g block of iron will increase its temperature by 20°C.
  
  **True.** The heat capacity of a 2.8 g block will be half of that of a 5.6g block. Therefore the temperature change will be double.

- **T  F** When the heat for a process is positive, there is always an increase in temperature of the system.
  
  **False.** Not for a phase change or chemistry.
For each of the following note what you would expect for the entropy of the system, surroundings, and total.

A container of liquid honey (the system) sitting in your kitchen (the surroundings) crystallizes

\[ \Delta S_{SYS} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

**Going from liquid to solid**

\[ \Delta S_{SURR} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

*It must be if the total is increasing*

\[ \Delta S_{TOTAL} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

Spontaneous so it is increasing

1 mole of an ideal gas initially at a pressure of 10 bar, expanding isothermally against a constant external pressure of 1 bar until mechanical equilibrium is reached.

\[ \Delta S_{SYS} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

**Volume is going up**

\[ \Delta S_{SURR} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

This must be endothermic. Work out and heat in. No change internal energy

\[ \Delta S_{TOTAL} \quad \text{Increase} \quad \text{Decrease} \quad \text{Stay the Same} \quad \text{No Way to Know} \]

Spontaneous.

A 25 g block of solid iron at a temperature 50 °C is dropped into a glass of ice water that contains 50 g of solid water and 50 g of liquid water at 0°C? Does all the ice melt?

\[
\begin{align*}
C_P, \text{solid water} &= 36 \text{ J K}^{-1} \text{ mol}^{-1} \\
C_P, \text{liquid water} &= 75.3 \text{ J K}^{-1} \text{ mol}^{-1} \\
C_P, \text{solid iron} &= 25.1 \text{ J K}^{-1} \text{ mol}^{-1} \\
\Delta_{\text{FUS}} H^o &= 6.02 \text{ kJ mol}^{-1}
\end{align*}
\]
How much heat is required to cool the block from 50°C to 0°C (a change of +50K)?

\[ q = nC\Delta T = \left(\frac{25g}{55.8g\text{ mol}^{-1}}\right)(25.1 J K^{-1}mol^{-1})(50 K) = 562.3 J \]

will that melt all the ice?

How much heat is required to melt 50g of ice?

\[ q = n\Delta H_{fus} = \left(\frac{50g}{18g\text{ mol}^{-1}}\right)(6020 J\text{ mol}^{-1}) = 16722 J \]

So the iron block will cool to 0°C and there will be lots of ice left over. The heat from the iron is only enough to melt 0.11 moles of water (or 1.8g)

\[ n = \frac{q}{\Delta H_{fus}} = \left(\frac{562.3 J}{6020 J\text{ mol}^{-1}}\right) = 0.11 \text{ mol} \]

(note: for this problem there are several “extra” piece of data that are not needed)