if no subscript written w/ ΔS
assume ΔS_{sys}

$\Delta S = S_{final} - S_{initial}$

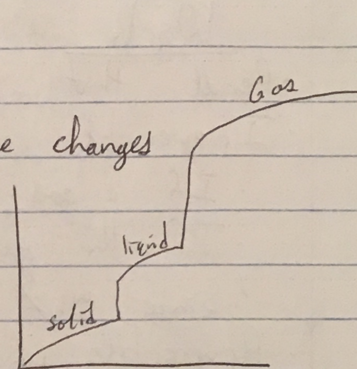
ΔS is a state function

$\boxed{\Delta S = \frac{q_{rev}}{T} = \frac{\Delta H_{sys}}{T}}$ $\quad q_p = \Delta H$

for a phase change liquid $\rightarrow$ gas ΔH_{vap}

$\Delta S_{vap} = \frac{\Delta H_{vap}}{T}$

Entropy Change and Temperature



- Entropy increases with temperature and at phase changes

Ex: H_2O at 298°K

$S(H_2O(l)) = 69.9\ J/(mol \cdot K)$

$S(H_2O(g)) = 188.8\ J/(mol \cdot K)$

$\Delta S_{vap} = 119\ J/(mol \cdot K)$

Entropy of the Surroundings

cont. from previous problem

- 6.01×10^3 J heat melts 1.00 mol ice at 0°C
- Suppose 1.00 mol ice melts at room temp at 23°C. What is ΔS for the surroundings?

$q_{surr} = -q_{sys}$

$\Delta S_{sys} = \frac{q_{sys}}{T}$

$\Delta S_{surr} = \frac{q_{surr}}{T} = \frac{-q_{sys}}{T}$

$= \frac{(6.01 \times 10^3\ J/mol)(1.00\ mol)}{295\ K} = -20.4\ J/K$

$\boxed{\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}}$

$\Delta S_{sys} = 22.0\ J/K$

$\Delta S_{surr} = -20.4\ J/K$

$\Delta S_{univ} = 1.6\ J/K$

$\Delta S_{univ} > 0 \Rightarrow$ spontaneous

Entropy and the 2nd Law

- Entropy (S) - measure of how dispersed energy in a sys. is at T
- 2nd Law - entropy of universe increases in any spontaneous process

$\hookrightarrow \Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr} > 0$