C. Entropy Changes in Surroundings

- Heat that flows into or out of the system changes the entropy of the surroundings

↳ For an isothermal process:

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

* If P is constant, $q_{sys} = \Delta H^\circ$ for the reaction

- The universe is composed of the system and the surroundings.

$$\therefore \Delta S_{univ} = \Delta S_{system} + \Delta S_{surroundings}$$

* Remember - spontaneous processes $\Delta S_{univ} > 0$

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

$\Delta S_{universe} = \Delta S_{sys} + \frac{-\Delta H_{sys}}{T}$, rearrange

$-T\Delta S_{univ} = \Delta H_{sys} - T\Delta S_{sys}$

defined as Gibbs free energy (ΔG)

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

ΔG will always be negative if ΔS is positive, spontaneous

$\Delta G^\circ = \Sigma \Delta G^\circ_{products} - \Sigma \Delta G^\circ_{reactants}$

(1) How does ΔG^0 change with Temperature?

(a) If ΔG is negative, the forward reaction is spontaneous
(b) If $\Delta G = \emptyset$, the system is at equilibrium
(c) If ΔG is positive, the reverse reaction is spontaneous

Effect of Temperature

ΔH	ΔS	$-T\Delta S$	$\Delta G = \Delta H - T\Delta S$	Reaction characteristics	Example
−	+	−	−	Spontaneous at all T.	$2O_3(g) \rightarrow 3O_2(g)$
+	−	+	+	Nonspontaneous at all T.	$3O_2(g) \rightarrow 2O_3(g)$
−	−	−	+ or −	Spontaneous at low T, nonspontaneous at high T	$H_2O(l) \rightarrow H_2O(s)$
+	+	−	+ or −	spontaneous at high T, nonspontaneous at low T	$H_2O(s) \rightarrow H_2O(l)$

(2) Under any conditions (standard or nonstandard), the free energy can be found this way:

$$\Delta G = \Delta G^0 + RT\ln Q$$

R: 8.314 J/mol K

standard conditions: concentrations - 1M, so Q = 1

at equilibrium, $Q = K$ & $\Delta G = 0$

$$0 = \Delta G^0 + RT\ln K$$

rearranged,

$$\Delta G^0 = -RT\ln K$$

or

$$K = e^{\frac{-\Delta G^0}{RT}}$$

HW: due mon., 12/9
p 837 − 839
56 a & c, 58 c & d, 66, 76, 78, 82