Today 3/19 Sections 17.6-9
Tuesday: Problem Club with Ali
Wednesday: Finish ch 17
Thursday: Expt 9

$N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $\Delta H^\circ_{rxn} = 55\,kJ$ $\Delta S^\circ_{rxn} = 175.7\ \frac{J}{K}$

enthalpy not favored — entropy favored

ΔH°_f	$+11.4\ kJ/mol$	$+33.2\ kJ/mol$	
	$\times 1\ mol$	$\times 2\ mols$	
ΔH°_{rxn}	$-11.4\ kJ$	$+66.4\ kJ$ $= 55\,kJ$	
S°_f	$304.3\ \frac{J}{mol \cdot K}$	$240.0\ \frac{J}{mol \cdot K}$	
	$\times 1\ mol$	$\times 2\ mol$	
ΔS°_{rxn}	$-304.3\ \frac{J}{K}$	$+ 480\ \frac{J}{K} = 175.7\ \frac{J}{K}$	

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

$\Delta G^\circ = (55.00kJ) - (298K)(.1757\ \frac{kJ}{K})$

$\Delta G^\circ = 55kJ - 52.36$

$\boxed{\Delta G^\circ = +2.64\ kJ}$

non-spontaneous at 298K

estimate ΔG at 310K.

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

$\Delta G = 55kJ - (310K)(.1757\ \frac{kJ}{K})$

$\Delta G = 55kJ - 54.47kJ$

$\boxed{\Delta G = 0.533\,kJ}$ non-spontaneous at 310K

estimate ΔG at 320K

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

$\Delta G = 55kJ - (320K)(.1757\ \frac{kJ}{K})$

$\Delta G = 55kJ - 56.22\,kJ$

$\boxed{\Delta G = -1.224\,kJ}$ spontaneous at 320K

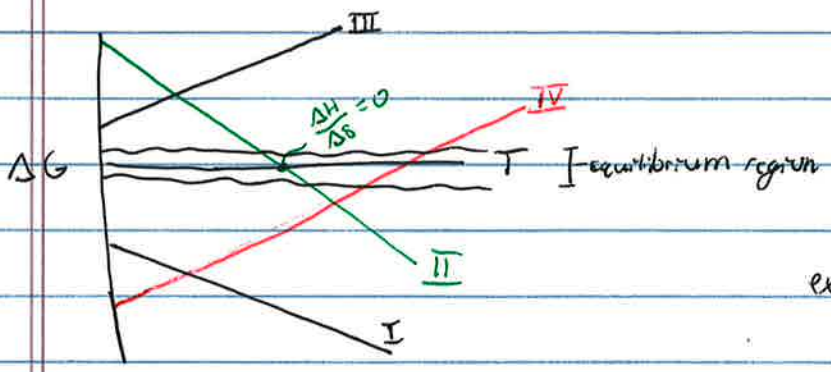
$\Delta G = \Delta H - T\Delta S$

I. $(-) = (-) - (+)(+)$ always spontaneous

II $= (+) - (+)(+)$ spontaneous at high T, nonspontaneous at low T

III $(+) = (+) - (+)(-)$ always non-spontaneous (never spontaneous)

IV $= (-) - (+)(-)$ spontaneous at low T, nonspontaneous at high T



$\Delta G = 0$ at equilibrium

$0 = \Delta H - T\Delta S$

$T = \frac{\Delta H}{\Delta S}$

example:

$T = \frac{55.0\text{kJ}}{.1757\text{kJ}} \cdot \frac{1\text{k}}{}$

$T = 313\text{k}$

example of type I:

$2O_3(g) \rightarrow 3O_2(g)$ $\Delta H^\circ = -286\text{kJ}$ $\Delta S^\circ = +137\text{J/K}$

$3O_2(g) \rightarrow 2O_3(g)$ $\Delta H^\circ = +286\text{kJ}$ $\Delta S^\circ = -137\text{J/K}$

$CH_4(g) \rightarrow CH_4(l)$ $\Delta H < 0$ $\Delta S < 0$

$CH_4(l) \rightarrow CH_4(g)$ $\Delta H > 0$ $\Delta S > 0$

↑ phase changes always have the same signs for $\Delta H/\Delta S$

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ $\Delta H = -803\text{ kJ}$, $\Delta S = -4.0\text{ J/K}$ } Rxn type IV

$R = 8.314 \frac{J}{mol \cdot K}$

in equilibrium region:

$\Delta G = \Delta G^\circ + RT \ln Q$ (Q_p for gases, Q_c for solns)

	$N_2O_4 \rightleftharpoons 2NO_2$		$\Delta G^\circ = +2.944 kJ$
I	1.0 atm	0.10 atm	$Q_p = \frac{(P_{NO_2})^2}{P_{N_2O_4}} = \frac{(0.10)^2}{(1)} = 1.0 \times 10^{-2}$
C	$-x$	$+2x$	
E	$1.0 - x$	$0.10 + 2x$	(x is will NOT be small)

$\Delta G = 2.944 kJ + (0.008314 \frac{kJ}{mol \cdot K})(298 K) \ln(1.0 \times 10^{-2})$

$\Delta G = -8.47 kJ$

at equilibrium, $\Delta G = 0$

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

$0 = \Delta G^\circ + RT \ln K$ (K_p for gases, K_c for solns)

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

$2.944 kJ = -(.008314 \frac{kJ}{mol \cdot K})(298 K) \ln K_p$

$K_p = 0.305$

* only works for reactions that go through 0 on ΔG vs T graph (so not I or III)