1. Consider the following reaction:

$$Zn(s) + I_2(s) \longrightarrow ZnI_2(s) \quad \Delta S^\circ_{rxn} = -3.02\,\frac{J}{K}$$

(a) When solid zinc and solid iodine are mixed at room temperature, $I_2(g)$ is observed. What can you infer about the sign of ΔH°_f of $ZnI_2(s)$?

Something happened, so spontaneous, $\Delta G < 0$. But since $\Delta S < 0$, so $\Delta H < 0$

(b) This reaction is spontaneous under what conditions?

☐ All temps ☐ High temps ■ Low temps ☐ No temps

2. Find the temperature at which $NH_3(l)$ and $NH_3(g)$ are equally favored.

Quantity	Value
$\Delta H^\circ_{\text{condensation of } NH_3}$	$-23.35\,\frac{kJ}{mol}$
S°_m of $NH_3(l)$	$95.36\,\frac{J}{mol{\cdot}K}$
S°_m of $NH_3(g)$	$192.77\,\frac{J}{mol{\cdot}K}$

$$\Delta G = \Delta H - T\Delta S = 0$$
$$\Delta S = 95.36\,\frac{J}{mol{\cdot}K} - 192.77\,\frac{J}{mol{\cdot}K}$$
$$= -97.41\,\frac{J}{mol{\cdot}K}$$
$$\Delta H = -23.35\,\frac{kJ}{mol} = -23{,}350\,\frac{J}{mol}$$
$$T = \frac{\Delta H}{\Delta S} = \frac{-23{,}350\,\frac{J}{mol}}{-97.41\,\frac{J}{mol{\cdot}K}}$$
$$= \boxed{240\,K}$$

3. Consider the following reaction:

$$PCl_3(g) + Cl_2(g) \longrightarrow PCl_5(g) \quad \Delta H = -87.9\text{ kJ}$$

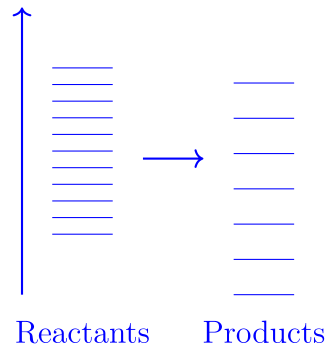
(a) Rank the three reactants and products in order of increasing standard molar entropy.

$$Cl_2 < PCl_3 < PCl_5$$

(b) This reaction is spontaneous under what conditions?

☐ All temps ☐ High temps ■ Low temps ☐ No temps

(c) On the diagram, sketch relative energy levels for the reactants and products, clearly illustrating the ΔH and ΔS for the reaction.



4. If a reaction is exothermic and non-spontaneous at a given temperature, then it

   - ☐ can never be spontaneous
   - ☐ may be spontaneous at a higher temperature
   - ■ may be spontaneous at a lower temperature
   - ☐ none of the above

5. Use the following data tables to answer the questions

Molecule	S°_m ($\frac{\text{J}}{\text{mol·K}}$)	ΔH°_f ($\frac{\text{kJ}}{\text{mol}}$)
N_2(g)	191.6	–
H_2(g)	130.7	–
NH_3(g)	192.5	–
H(g)	–	218.0
N(g)	–	472.7

Bond	Bond Enthalpy
N–N	945 $\frac{\text{kJ}}{\text{mol}}$
N–H	435 $\frac{\text{kJ}}{\text{mol}}$

(a) Determine the bond enthalpy of H–H bond.

$$\frac{1}{2}\,\mathrm{H_2(g)} \longrightarrow \mathrm{H(g)}$$
$$(\mathrm{H{-}H}) = 2 \times \Delta H^\circ_f(\mathrm{H(g)})$$
$$(\mathrm{H{-}H}) = \boxed{436\ \mathrm{kJ/mol}}$$

(b) Calculate ΔG°_f of NH_3(g), and determine the conditions where it is thermodynamically favored (Low T, High T, All T, or no T).

$$\frac{1}{2}\,\mathrm{N_2} + \frac{3}{2}\,\mathrm{H_2} \longrightarrow \mathrm{NH_3}$$
$$\begin{aligned}\Delta H^\circ_{\text{rxn}} &= \sum \text{bonds broken} - \sum \text{bonds formed}\\ &= \left(\tfrac{1}{2}\cdot 945\,\tfrac{\text{kJ}}{\text{mol}} + \tfrac{3}{2}\cdot 436\,\tfrac{\text{kJ}}{\text{mol}}\right) - 3\cdot 435\,\tfrac{\text{kJ}}{\text{mol}}\\ &= -178.5\,\tfrac{\text{kJ}}{\text{mol}}\end{aligned}$$

$$\begin{aligned}\Delta S^\circ_{\text{rxn}} &= 192.5\,\tfrac{\text{J}}{\text{mol·K}} - \left(\tfrac{1}{2}\cdot 191.6\,\tfrac{\text{J}}{\text{mol·K}} + \tfrac{3}{2}\cdot 130.7\,\tfrac{\text{J}}{\text{mol·K}}\right)\\ &= -99.35\,\tfrac{\text{J}}{\text{mol·K}}\end{aligned}$$

$$\begin{aligned}\Delta G^\circ &= \Delta H - T\Delta S\\ &= -178.5\,\tfrac{\text{kJ}}{\text{mol}} - 298\,\text{K}\left(-0.099\,\tfrac{\text{kJ}}{\text{mol·K}}\right)\\ &= \boxed{-148.9\,\tfrac{\text{kJ}}{\text{mol}}}\end{aligned}$$

$$\Delta H < 0 \quad \Delta S < 0 \implies \boxed{\text{Low Temperature}}$$

(c) Is ammonia a thermodynamically stable compound?

Yes, $\Delta G_f < 0$

Homework Problem 19

1. Given the entropy change of the reaction is $\Delta S^\circ = -99\,\frac{\mathrm{J}}{\mathrm{K}}$, determine the temperature range where this reaction is spontaneous.

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)}$$

Bond	Bond Enthalpy
H–H	$436\,\frac{\mathrm{kJ}}{\mathrm{mol}}$
Cl–Cl	$243\,\frac{\mathrm{kJ}}{\mathrm{mol}}$
H–Cl	$431\,\frac{\mathrm{kJ}}{\mathrm{mol}}$