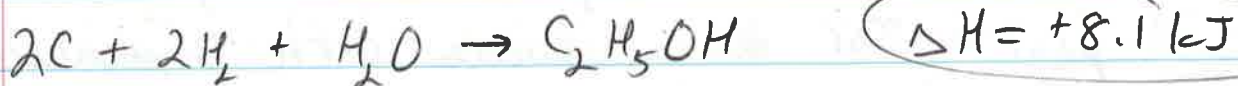


Sample #1



Hess' Law!



(c) $\Delta S = \sum S^\circ(\text{prod}) - \sum S^\circ(\text{react})$
 $= (160.7 \text{ J/K}) - [2(5.740) + 2(130.6) + 69.91]$
 $= -181.9 \text{ J/K}$

(d) $\Delta G = \Delta H - T\Delta S$
 $= (+8.1 \text{ kJ}) - (298K)(-0.1819 \text{ J/K})$
 $= +62.3 \text{ kJ}$

NONSPONTANEOUS

Sample #2

(a) The reaction is predicted to have **NEGATIVE** ΔS .
The reaction goes from 3 moles of gas (higher entropy) to 2 moles of gas + solid (lower entropy).
Entropy **DECREASES**, so ΔS is negative.

(b) As temperature increases, the forward reaction would become less spontaneous (eventually nonspontaneous).

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

For reaction with the neg ΔH and neg ΔS , the positive $T\Delta S$ term increases with temperature, and eventually $T\Delta S > \Delta H$, resulting in a pos ΔG .

(d)
$$\Delta G = \Delta H - T\Delta S$$

Set $\Delta G = 0$ and solve for T

(Why is it approximate? Because ΔH and ΔS values are not constant - they also change slightly as T changes. If we always plug in the same ΔH + ΔS at every T , this is not exact - only approximate.)

Sample #3

(a) $73.1 \text{ g NO} \times \frac{1 \text{ mol}}{28.0 \text{ g}} = 2.61 \text{ mol NO} \times \frac{-114.1 \text{ kJ}}{2 \text{ mol NO}} \Rightarrow$

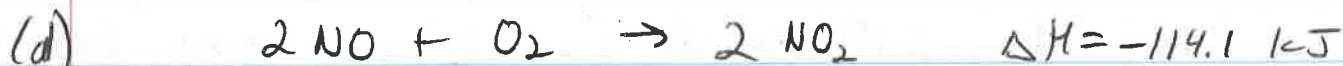
-148.9 kJ released

(b) (i) OMIT

(ii) As temp increased, ΔG becomes less negative (more positive.) Since $\Delta G = \Delta H - T\Delta S$ and both ΔH and ΔS are neg, the answer for ΔG depends on ΔH versus $T\Delta S$. At higher temperatures, the positive $T\Delta S$ term is $> \Delta H$, resulting in less negative (or even pos) ΔG .

(c) $\Delta S = \sum S^\circ(\text{prod}) - \sum S^\circ(\text{react})$
 $-146.5 = [2(240.1)] - [2(210.8) + S^\circ \text{ of } \text{O}_2]$

$S^\circ \text{ of } \text{O}_2 = 146.5 + 58.6 = \text{+205.1 J/mol K}$



Bond energies	2 N-O	$2(607)$		2 O-O	
	O-O	495		$4 \times 4 \text{ N-O}$	
break		1709 kJ		form	$-4x$

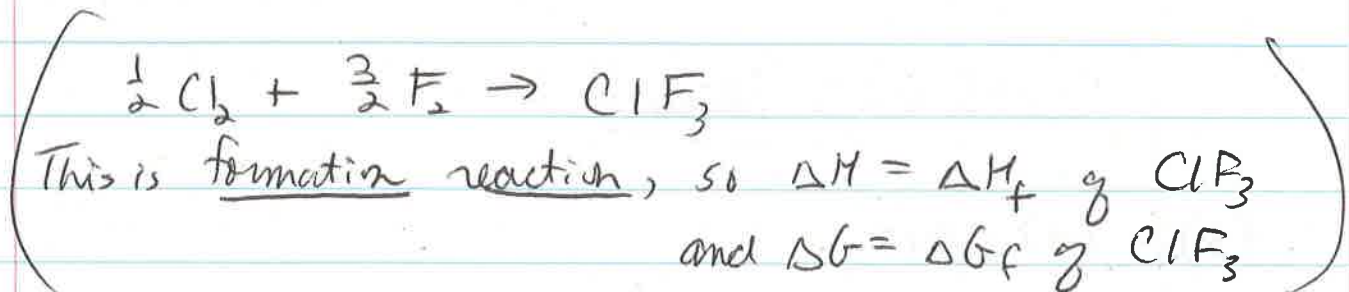
$1709 + (-4x) = -114.1$

Bond energy of NO in NO_2

$x = \frac{3987}{4} \text{ kJ/mol}$
 455.8 kJ/mol

455.8 kJ/mol

Sample #4



$$\Delta G = \Delta H - T\Delta S$$
$$-246 \text{ kJ} = -326.4 \text{ kJ} - (298) \Delta S$$

$$-0.270 \text{ kJ/K} = \Delta S$$

$$-270 \text{ J/K} = \Delta S$$

(c) ΔS would be even more negative, as this would be an even greater decrease in entropy.

Now: 4 moles gas \Rightarrow 2 moles gas | decrease in entropy ΔS
Change: 4 mole gas \Rightarrow liquid | LARGER decrease in entropy

(d) (i) Although both substances are gases with lots of particle motion, ClF_3 molecules are larger and would have more internal motion (atoms wiggling inside the molecule)

(ii) $\Delta S = \sum S^\circ(\text{prod}) - \sum S^\circ(\text{react})$
 $-270 = [2(281.50)] - [222.96 + 3S^\circ_{\text{F}_2}]$

$$S^\circ \text{ of } \text{F}_2 = 203.3 \text{ J/mol K}$$

Sample #5

Given



$$\Delta H = -2,183.5 \text{ kJ}$$

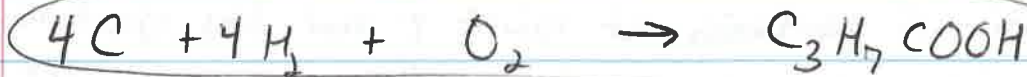
(a) $\Delta H_{\text{rxn}} = \sum \Delta H_f(\text{prod}) - \sum \Delta H_f(\text{react})$

$$-2183.5 = [4(-393.5) + 4(-285.8)] - [\Delta H_f + 5(0.0)]$$

$$\Delta H_f = -1574 + -1143.4 + 2183.5$$

$$\Delta H_f = -533.9 \text{ kJ/mol for } \text{C}_3\text{H}_7\text{COOH}$$

(b)



Note

ΔH_f of $\text{C}_3\text{H}_7\text{COOH}$

Find it for this one

(c) $\Delta S_f = \sum S^\circ(\text{prod}) - \sum S^\circ(\text{react})$

$$= [226.3] - [4(5.69) + 4(130.6) + 205.0]$$

$$= -523 \text{ J/K}$$

(don't really need ΔS for the combustion reaction)

(d) ΔH_f for formation reaction? Answer from (a)

$$\Delta G_f = \Delta H_f - T\Delta S_f$$

$$\Delta G_f = (-533.9 \text{ kJ/mol}) - (298\text{K})(-0.523 \text{ kJ/K})$$

$$= -378 \text{ kJ}$$

Sample #6

Given: reaction is SPONTANEOUS,

$$\text{so } \Delta G < 0 \text{ (neg)}$$

- (a) Reaction goes from
15 moles gas \Rightarrow 8 moles of gas + liquid
higher entropy lower entropy

There is a DECREASE in entropy, so $\Delta S < 0$ (neg)

- (b) See above - it says reaction is spontaneous

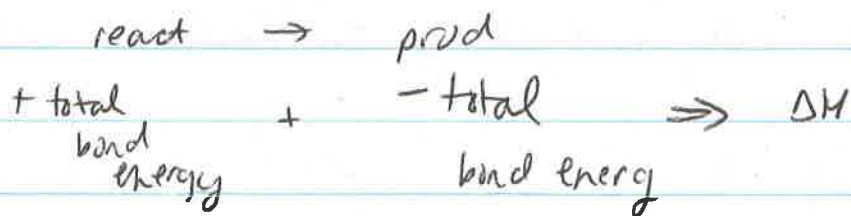
which means ΔG is NEGATIVE

Reaction is spontaneous at lower T and less spontaneous at higher T

- (c) ΔH must be NEGATIVE.

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

~~In order~~ In order for ΔG to be negative, given that $-T\Delta S$ is a positive term, ΔH must be NEGATIVE



∴ ΔH is negative, this means the products have greater bond energy than the reactants. More energy is RELEASED when products^{bonds} form than is absorbed to break reactant bonds