$$C_2H_2(g) + 2\,H_2(g) \rightarrow C_2H_6(g)$$

Information about the substances.

Substance	$S°$ (J/mol K)	$\Delta H_f°$ (kJ/mol)	Bond	Bond Energy (kJ/mol)
$C_2H_2(g)$	200.9	226.7	C−C	347
$H_2(g)$	130.7	0	C=C	611
$C_2H_6(g)$	--------	−84.7	C−H	414
			H−H	436

1. If the value of the standard entropy change, ΔS°, for the reaction is −232.7 joules per mole Kelvin, calculate the standard molar entropy, S°, of C_2H_6 gas.

$$\Delta S = \sum S°_{products} - \sum S°_{reactants}$$ "Big Mamma"

$$-232.7 \frac{J}{mol_{rxn} \cdot K} = X - (200.9 + 2(130.7))$$

$$X = 229.6 \frac{J}{mol_{rxn} \cdot K}$$

2. Calculate the value of the standard free-energy change, ΔG°, for the reaction. What does the sign of ΔG° indicate about the reaction above?

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

° = "naught"
Standard cond.
25°C

$$\Delta H = (-84.7) - (226.7) = -311.4 \frac{kJ}{mol_{rxn}}$$

$$\Delta G = \left(-311.4 \frac{kJ}{mol_{rxn}}\right) - (298\,K)\left(-.2327 \frac{kJ}{mol_{rxn} \cdot K}\right)$$

$$\Delta G = -242.1 \frac{kJ}{mol_{rxn}}$$

3. Calculate the value of the equilibrium constant, K, for the reaction at 298 K.

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

$$-242.1 \frac{kJ}{mol_{rxn}} = -\left(8.31 \times 10^{-3} \frac{kJ}{mol \cdot K}\right)(298K) \ln K$$

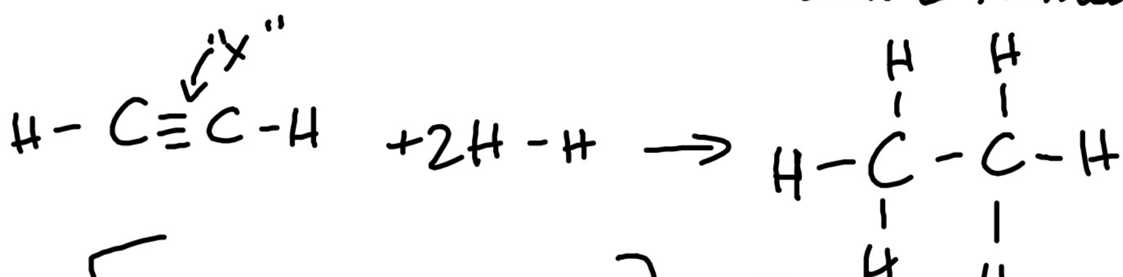
$$\ln K = 97.76$$

$$K = e^{97.76}$$

$$K = 2.87 \times 10^{42}$$

4. Calculate the value of the C ≡ C bond energy in C_2H_2 in kilojoules per mole.

$$\Delta H = \Sigma_{\text{Bond E Broken}} - \Sigma_{\text{Bond E Formed}}$$



$$-311.4 = \underbrace{\left[x + 2(414) + 2(436)\right]}_{\text{Broken}} - \underbrace{\left[347 + 6(414)\right]}_{\text{Formed}}$$

$$x + 1700 \quad - 2831$$

$$x = 819.6 \frac{kJ}{mol}$$