HW7

	Exothermic reactions	Endothermic reactions
	Gives out heat to surrounding, Products more stable than reactants	Takes in heat from surrounding, Products less stable than reactants
Enthalpy change ΔH – energy (in form of heat) exchanged with surrounding at constant pressure kJ/mol	Negative	Positive
	Standard enthalpy change, ΔH°, enthalpy change for molar amounts of substance under standard conditions enthalpy change of reaction, ΔH_r, enthalpy change for molar amounts of reactants when chemical reaction occurred. Enthalpy change of formation, ΔH_f, heat given out or taken in when one mole of substance is formed from its elements in their standard states. $\Delta Hr = \sum \Delta Hf$ (products) - $\sum \Delta Hf$ (reactants) Bond enthalpy – heat taken in when one mole of covalent bonds, in a gaseous molecule, is broken. Bond breaking – endothermic process, bond making exothermic process. $\Delta Hr = \sum$ (bond broken) - $\sum$ (bond formed)	
Entropy, S J/Kmol	Gas>liquids>solid Depends on the number of ways the energy can be distributed among the particles	
Gibbs free energy, G, $\Delta G = \Delta H - T\Delta S$ Help us to determine if the process will be spontaneous	If $\Delta S > 0$ and $\Delta G < 0$ the process is spontaneous at any temperatures	If $\Delta S < 0$ and $\Delta G > 0$ the process is nonspontaneous at any temperatures

Consider the following reaction under standard conditions (298K):
C_2H_4 (g) + H_2 (g) → C_2H_6 (g)

Given:
Average bond enthalpies (kJ/mol)
C=C 614
C-C 347
C-H 413
H-H 436

Standard enthalpies of formation (kJ/mol)

C_2H_4 (g) 52.5
H_2 (g) 0
C_2H_6 (g) -84.7

Standard entropy change for the reaction -112 J/mol K.

Questions:
Using the bond enthalpy data estimate the reaction enthalpy change
Using the standard enthalpies of formation calculate the reaction enthalpy change
State whether the reaction is spontaneous or not under standard conditions.