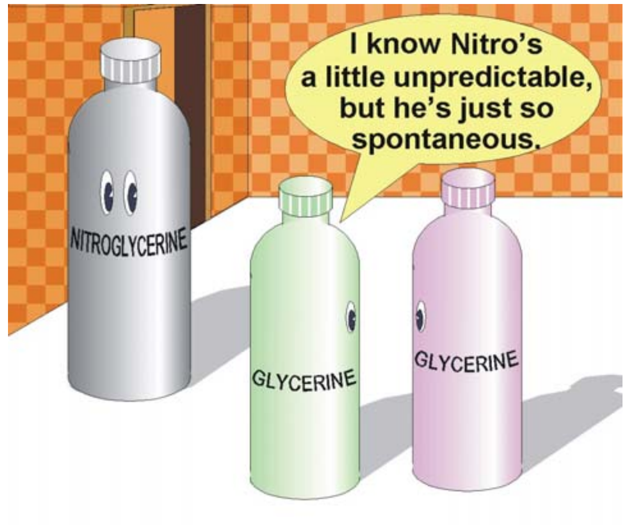
Name ________________________________

Date __________________ Period _____

Being Spontaneous

Whether a reaction proceeds spontaneously or not depends on the balance between two natural tendencies: [1] The drive toward greater stability (reduced potential energy); and [2] The drive toward less organization (increased entropy). In nature, systems tend toward the lowest possible energy or enthalpy (H). Exothermic reactions are favored (ΔH is negative). In nature, systems also tend toward greater randomness (disorder) or entropy (S). When entropy increases, ΔS, the change in entropy is positive. High entropy is favored by increased temperature. The Gibbs free energy change (ΔG) predicts whether or not a reaction is spontaneous. It takes into acount the change in enthalpy and the change in entropy. The Gibbs free enrgy change is the difference between the energy change (ΔH) and the product of the absolute or Kelvin temperature (T) and the entropy change (ΔS)

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



Chemical dating preferences

For a system at equilibrium, $\Delta G = 0$. In order for a system to change spontaneously, the resulting ΔG must be negative. If the drive toward lower energy and higher entropy cannot be satisfied at the same time, the type of change that will be favored will depend on the temperature. At low temperatures, the term $T\Delta S$ will be small, and ΔH will have the greatest effect on the free energy. At high temperatures, the term $T\Delta S$ will be large, and ΔS will have the greatest effect on the free energy.

Based on your reading above, fill in the table below and answer the questions that follow.

Reaction Conditions			Is the reaction spontaneous? (*Yes*, *No*, *Likely*, or *Unlikely*)
Temperature	**ΔH**	**ΔS**	
High	+	–	
High	+	+	
High	–	–	
High	–	+	
Low	+	–	
Low	+	+	
Low	–	–	
Low	–	+	

Continue ☞

1. At low temperatures, water is a solid (ice), while at high temperatures, water is a gas (steam). Explain this based on the change in Gibbs free energy (ΔG). ______

2. Wood, a carbohydrate and a solid, burns to form water vapor and carbon dioxide, both gases.
   a. Is the reaction exothermic or endothermic? ______
   b. Is the entropy during the reaction increasing or decreasing? ______
   c. What do your answers to the previous two questions tell you about ΔG? Should the reaction be spontaneous at room temperature? ______
   d. Is the reaction spontaneous at room temperature? How do you explain this? (*HINT:* What other energy considerations besides ΔG determine if a reaction will occur?) ______

3. Referring to the table to the right, determine the entropy of formation in each of the examples below:

Sample Problem

Find the entropy of formation of carbon dioxide.

$\Delta G_f° = -394.2$ kJ/mol	$\Delta G = \Delta H - T\Delta S$
$\Delta H_f° = -393.3$ kJ/mol	-394.2 kJ/mol = -393.3 kJ/mol – 298K(ΔS)
T = 298 K (see table)	-0.9 kJ/mol = -298K(ΔS)
	$\Delta S = 0.003$ kJ/mol•K

a. Ethene

b. Water vapor

STANDARD ENERGIES OF FORMATION OF COMPOUNDS AT 1 atm AND 298 K

Compound	Heat (Enthalpy) of Formation* kJ/mol ($\Delta H_f°$)	Free Energy of Formation kJ/mol ($\Delta G_f°$)
Aluminum oxide $Al_2O_3(s)$	–1674.1	–1580.9
Ammonia $NH_3(g)$	–46.0	–16.3
Barium sulfate $BaSO_4(s)$	–1471.8	–1361.0
Calcium hydroxide $Ca(OH)_2(s)$	–985.2	–897.9
Carbon dioxide $CO_2(g)$	–393.3	–394.2
Carbon monoxide $CO(g)$	–110.4	–137.1
Copper (II) sulfate $CuSO_4(s)$	–770.8	–661.3
Ethane C_2H_6 (g)	–84.4	–33.0
Ethene (ethylene) $C_2H_4(g)$	52.3	68.1
Ethyne (acetylene) $C_2H_2(g)$	226.6	209.0
Hydrogen fluoride $HF(g)$	–270.9	–273.0
Hydrogen iodide $HI(g)$	26.3	1.7
Iodine chloride $ICl(g)$	18.0	–5.4
Lead (II) oxide $PbO(s)$	–215.3	–188.1
Magnesium oxide $MgO(s)$	–601.1	–568.9
Nitrogen monoxide $NO(g)$	90.3	86.5
Nitrogen dioxide $NO_2(g)$	33.0	51.4
Potassium chloride $KCl(s)$	–436.4	–408.8
Sodium chloride $NaCl(s)$	–410.9	–383.7
Sulfur dioxide $SO_2(g)$	–296.4	–299.7
Water $H_2O(g)$	–241.6	–228.2
Water $H_2O(\ell)$	–285.5	–237.0

* Minus sign indicates an exothermic reaction.

Sample equations:

$2Al(s) + {}^3/_2\, O_2(g) \rightarrow Al_2O_3(s) + 1674.1$ kJ

$2Al(s) + {}^3/_2\, O_2(g) \rightarrow Al_2O_3(s) \quad \Delta H = -1674.1$ kJ/mol