1) For the reaction

$CO(g) + 3H_2(g) \rightarrow CH_4(g) + H_2O(g)$

ΔH^o = **-206.1** kJ and ΔS^o = **-214.7** J/K
The equilibrium constant for this reaction at **344** K is:
A) 1.74×10^{46}
B) 1.21×10^{20}
C) 3.46×10^{42}
D) 8.28×10^{-21}

The equilibrium constant is related to the change in the standard Gibbs Free Energy for a reaction:
$\Delta G^o_{rxn} = \Delta H^o_{rxn} - T\,\Delta S^o_{rxn} = -RT \ln K_{eq}$

Use ΔH^o and ΔS^o to calculate ΔG^o for the reaction at 344 K:
$\Delta G^o = \Delta H^o - T\,\Delta S^o$
... = **-206.1** kJ - (**344** K)(**-0.2147** kJ/ K) = **-132.24** kJ = **-1.3224e+5** J

Carry extra digits here, to avoid roundoff problems when taking the antilog later.

Use ΔG^o to calculate ln K_{eq}:
$\Delta G^o_{rxn} = -RT \ln K_{eq}$

OR

$\ln K = -(\Delta G^o)/RT$
... = - (**-1.3224e+5** J/mol K) / (8.314 J/mol K)(**344**K)
... = **46.239**

Raise to the base "e" to clear the logarithm:
$K = e^{46.239}$ = **1.21e+20**

Note that K_{eq} is greater than 1.
Products are favored at equilibrium.

2) For the reaction

$N_2(g) + O_2(g) \rightarrow 2NO(g)$

ΔH^o = **181** kJ and ΔS^o = **24.9** J/K
K would be greater than 1 at temperatures:_____

A) below than 7.27×10^3 K
B) above than 7.27×10^3 K
C) above 273.15 K
D) below 298.15 K

The variation in ΔG^o with temperature can be calculated using the following equation, assuming that ΔH^o and ΔS^o are constant.

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

Since ΔH^o and ΔS^o are both positive, it is the negative $-T\Delta S^o$ term that can cause ΔG^o to become negative. Since this term is larger at higher temperatures, **ΔG^o will be negative at relatively high temperatures** and positive at relatively low temperatures.

Solving for the crossover temperature at which $\Delta G^o = 0$

$\Delta G^o = \Delta H^o - T\Delta S^o = 0$

$T = \Delta H^o / \Delta S^o =$
181 kJ / **2.49e-2** kJK^{-1} = **7.27e+3** K

Therefore, ΔG^o will be negative above **7.27e+3** K.

Since $\Delta G^o = -RT \ln K$, when $\Delta G^o < 0$, $K > 1$ so the reaction will have $K > 1$ above **7.27e+3** K.