1) Consider the reaction

$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(s) + H_2(g)$$

Using standard absolute entropies at 298K, calculate the entropy change for the **system** when **1.74** moles of **Fe(s)** react at standard conditions.

	S^o (J/K mol)		
$FeCl_2(s)$	118.0	$H_2(g)$	130.7
Fe(s)	27.8	HCl(aq)	56.5

First calculate ΔS^o_{rxn} from standard absolute entropies (Appendix L Book):
$\Delta S^o_{rxn} = \Sigma S^o(\text{products}) - \Sigma S^o(\text{reactants})$

	Moles	*Species*	S^o *(J/K mol)*
Products	**1**	**$FeCl_2(s)$**	**118.0**
	1	**$H_2(g)$**	**130.7**
Reactants	**1**	**Fe(s)**	**27.8**
	2	**HCl(aq)**	**56.5**

ΔS^o_{rxn} = (**1** mol)(**118.0** J/K mol) + (**1** mol)(**130.7** J/K mol) - [(**1** mol)(**27.8** J/K mol) + (**2**mol)(**56.5** J/K mol)]
... = **107.9** J/K

Now calculate the entropy change for 1.74 moles of Fe(s) reacting:
ΔS = **1.74** mol **Fe(s)** x [**107.9** JK^{-1} / **1** mol **Fe(s)**]
... = **187.7** J/K

2) Consider the reaction

$$4HCl(g) + O_2(g) \rightarrow 2H_2O(g) + 2Cl_2(g)$$

Using standard thermodynamic data at 298K, calculate the entropy change for the **surroundings** when **1.51** moles of **HCl(g)** react at standard conditions.

	ΔH^o_f (kJ/mol)		
$H_2O(g)$	-241.8	$Cl_2(g)$	0.0
$HCl(g)$	-92.3	$O_2(g)$	0.0

The entropy of the SURROUNDINGS changes when a reaction gives off or absorbs heat.
The relationship is:
$\Delta S^o_{surroundings} = - (\Delta H^o_{rxn} / T)$

You must first calculate ΔH^o_{rxn} from standard enthalpies of formation (linked under "Tables" above):
$\Delta H^o_{rxn} = \Sigma \Delta H^o_f(\text{products}) - \Sigma \Delta H^o_f(\text{reactants})$

	Moles	*Species*	*ΔH^o_f (kJ/mol)*
Products	2	$H_2O(g)$	-241.8
	2	$Cl_2(g)$	0.0
Reactants	4	$HCl(g)$	-92.3
	1	$O_2(g)$	0.0

ΔH^o_{rxn} = **2** mol (**-241.8** kJ/mol) + **2** mol (**0.0** kJ/mol) - [**4** mol (**-92.3** kJ/mol) + **1** mol (**0.0** kJ/mol)]
... = **-114.4** kJ

... = **-114.4** kJ x 1000 J/kJ = **-1.144e+5** J

Now calculate $\Delta S^o_{surroundings}$:
$\Delta S^o_{surroundings}$ = - (ΔH^o_{rxn} / T)
= - (**-1.144e+5** J)/ 298 K = **383.9** JK^{-1}

If 1.51 moles of HCl(g) react:
1.51 mol **HCl(g)** x (**383.9** JK^{-1} / **4** mol **HCl(g)**) = **144.9** JK^{-1}

Note that the reaction is **exothermic (heat is given off by the system),** and the entropy of the surroundings therefore **increases** at constant temperature.

3) When **12.0** mL of a **1.21x 10^{-4}** M **sodium sulfide** solution is combined with **22.0** mL of a **4.71 x 10^{-4}** M **iron(II) sulfate** solution does a precipitate form ?
FeS Ksp = 4.9 x 10^{-18}

Final volume 34.0 mL
Final concentrations: S^{2-} = $(C_1 V_1)/V_2$ = (1.21 e-4 M x 12.0 mL)/34.0 mL = 4.27 e-5 M
Fe^{2+} = $(C_1 V_1)/V_2$ = (4.71e-4 M x 22 mL)/ 34 mL = 3.05 e-4 M

[S2-].[Fe2+] = 4.27 e-5 x 3.05 e-4 = 1.3 e-8 >> 4.9 e-18

4) Solid **sodium sulfide** is slowly added to **75.0** mL of a **0.0404** M **manganese(II) acetate** solution. The concentration of **sulfide** ion required to just initiate precipitation is:
___**?**___ M.
MnS Ksp = 5.1 x 10^{-15}

Ksp = $[Mn^{2+}] . [S^{2-}]$ = 5.1 x 10^{-15}

$[S^{2-}]$ = 5.1 x 10^{-15} / .0404 = 1.3 x 10^{-13} M