1. Consider the reaction:

$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g)$$

Estimate ΔG° at each temperature and predict whether or not the reaction is spontaneous. (Assume that ΔH° and ΔS° do not change too much within the given temperature range.)

Species	$\Delta H_f^\circ\ (kJ/mol)$	$\Delta S_f^\circ\ (J/mol \cdot K)$
$NO(g)$	91.3	210.8
$O_2(g)$	0	205.2
$NO_2(g)$	33.2	240.1

(a) 298 K

(b) 715 K

(c) 855 K

$\Delta H_{rxn}^\circ = (2 \times 33.2\ kJ/mol) - (2 \times 91.3\ kJ/mol + 0\ kJ/mol) = -116.2\ kJ/mol$

$\Delta S_{rxn}^\circ = (2 \times 240.1 J/mol \cdot K) - (2 \times 210.8\ J/mol \cdot K + 1 \times 205.2\ J/mol \cdot K) = -146.6\ J/K = -0.1466\ kJ/(Kmol)$

Then: Use $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$.

a) $-72.5\ kJ/mol$
b) $-11.4\ kJ/mol$
c) $9.1\ kJ/mol$

2. Calculate ΔG for the following reaction at 25°C under the following conditions:

$$2\,H_2S(g) + SO_2(g) \longrightarrow 3\,S(s, \text{rhombic}) + 2\,H_2O(g) \qquad \Delta G^\circ_{rxn} = -102\ kJ/mol$$

$P_{H_2S} = 2.00\ atm$, $P_{SO2} = 1.50\ atm$, $P_{H_2O} = 0.0100\ atm$

Recognize here that you are not at standard conditions (the pressures are not 1 atm).

So $\Delta G_{rxn} = \Delta G^\circ_{rxn} + RTlnQ$

$Q = \frac{P^2_{H_2O}}{P^2_{H_2S}P_{SO_2}} = 1.67 \times 10^{-5}$

Be careful with the units. R has units of $J/mol \cdot K$, so you should either convert the energies to J or R to units of $kJ/mol \cdot K$.

$\Delta G_{rxn} = -102,000\ J/mol + (8.314\ J/mol \cdot K)(298\ K)ln(1.67 \times 10^{-5}) = -129253\ J/mol = -129\ kJ/mol$

3. Calculate the equilibrium constant for the following reaction at 298 K.

$$N_2O_4(g) \rightleftharpoons 2\,NO_2(g) \qquad \Delta G^\circ_{rxn} = 2.8\ kJ/mol$$

Again, be careful with the units of R in this problem.

$\Delta G^\circ_{rxn} = -RTlnK$

$lnK = \frac{-\Delta G^\circ_{rxn}}{RT} = \frac{-2.8 \times 10^3\ J/mol}{(8.314\ J/mol \cdot K)(298\ K)} = -1.13$

$K = e^{-1.13} = 0.32$

4. Balance the following reaction in acidic solution.

$$Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$$

Oxidation states: $Zn = 0$, $Zn^{2+} = +2$, $Sn^{2+} = +2$, $Sn(s) = 0$

So Zn(s) is oxidized and Sn^{2+} is reduced.

Mass Balance (no need to worry about oxygen or hydrogen):
Oxidation: $Zn(s) \longrightarrow Zn^{2+}$
Reduction: $Sn^{2+} \longrightarrow Sn(s)$

Add electrons to balance the charge for each reaction:
Oxidation: $Zn(s) \longrightarrow Zn^{2+} + 2\,e^-$
Reduction: $Sn^{2+} + 2\ e^- \longrightarrow Sn(s)$

Then make the number of electrons equal in each reaction (they already are).

Then add the reactions together:

$Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$

5. Balance the following reaction in basic solution.

$$H_2O_2(aq) + ClO_2(aq) \longrightarrow ClO_2^-(aq) + O_2(g)$$

Oxidation States: In H_2O_2, H = +1, O = -1
In ClO_2, O = -2, Cl = +4
In ClO_2^-, O = -2, Cl = +3
In O_2, O = 0

Oxidation: $H_2O_2 \longrightarrow O_2$
Reduction: $ClO_2 \longrightarrow ClO_2^-$

Mass balance

$H_2O_2 \longrightarrow O_2 + 2\,H^+$
$ClO_2 \longrightarrow ClO_2^-$

Add the base to both sides to neutralize H^+ where needed.

$H_2O_2 + 2\,OH^- \longrightarrow O_2 + 2\,H^+ + 2\,OH^-$
$ClO_2 \longrightarrow ClO_2^-$

Then combine H^+ with OH^- to make water.

$H_2O_2 + 2\,OH^- \longrightarrow O_2 + 2\,H_2O$
$ClO_2 \longrightarrow ClO_2^-$

Charge balance

$H_2O_2 + 2\,OH^- \longrightarrow O_2 + 2\,H_2O + 2\,e^-$
$ClO_2 + e^- \longrightarrow ClO_2^-$

Need to multiply the bottom reaction by two so that each reaction has two electrons.

$H_2O_2 + 2\,OH^- \longrightarrow O_2 + 2\,H_2O + 2\,e^-$
$2\,ClO_2 + 2\,e^- \longrightarrow 2\,ClO_2^-$

Now add together and cancel things that appear on both sides

$2\,OH^- + H_2O_2 + 2\,ClO_2 \longrightarrow O_2 + 2\,H_2O + 2\,ClO_2^-$

Don't forget to check. Count the atoms and total charge on each side. They should match.