(10) 1. List the condition(s) under which each of the following equations can be applied. If no conditions are required, write "none."

$dG \leq 0$ (for spontaneity and equilibrium)

$\Delta G^\circ = -RT\ln K$

$\overline{C}_P = \overline{C}_V + R$

$\ln K = -\frac{\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R}$

$dG = VdP - SdT$

(9) 2. Describe three independent ways that would allow you to measure the ΔG° value of a chemical process.

(10) 3. A quantity of 0.20 mole of CO_2 was heated to 350°C with an excess of graphite in a closed container until the following equilibrium is reached:

$$C(s) + CO_2(g) \rightleftharpoons 2CO(g)$$

(a) If the average molar mass of the gases was 35 g mol^{-1}, calculate the mole fractions of CO and CO_2. (b) Calculate K_P for the reaction if the total pressure is 11 atm. (*Hint*: The average molar mass is the sum of the products of the mole fraction of each gas times its molar mass. C = 12.01 g mol^{-1}; O = 16.00 g mol^{-1}.)

(15) 4. An electrochemical cell consists of a Ag electrode in contact with 346 mL of 0.100 M $AgNO_3$ solution and a Mg electrode in contact with 288 mL of 0.100 M $Mg(NO_3)_2$ solution at 25°C. (a) Calculate the emf (E) of the cell. (b) A current is drawn from the cell until 1.20 g of Ag has been deposited at the Ag electrode. Calculate the value of E at this stage of operation. Assume the volumes to remain unchanged. (The standard reduction potentials for Ag and Mg are 0.800 V and –2.372 V, respectively. The molar mass of Ag is 107.9 g mol^{-1}. Keep in mind that the number of Mg^{2+} ions produced at the anode is equal to half of the Ag^+ ions reduced at the cathode.)