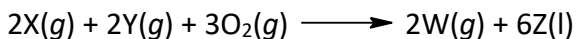


THERMODYNAMICS (no calculator)

You can do all of these problems with or without a calculator. Answers given were generated without a calculator.

1) Given the reaction:

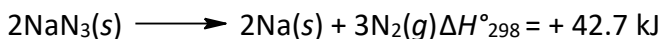


a) Use the enthalpies of formation given to calculate ΔH° for the reaction. **(-1250 kJ)**

| Compound | ΔH°_f (kJ/mol) |
|----------|-----------------------------|
| X(g) | -50 |
| Y(g) | -75 |
| W(g) | +150 |
| Z(l) | -300 |

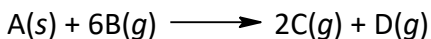
b) Calculate ΔE°_{298} for the reaction. **(-1237.5 kJ)**

2) Given the thermochemical equation:



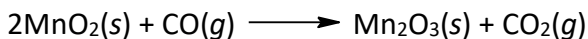
Calculate $\Delta H^\circ_{f,298}$ for $\text{NaN}_3(s)$. **(-21.35 kJ/mol)**

3) A 10.00 mmol sample of A(s) was reacted with excess B(g) in a bomb calorimeter. The total heat capacity of the calorimeter was 10.00 kJ/°C. The temperature of the calorimeter and its contents increased from 24.00°C to 26.00°C. Calculate ΔH°_{298} for the reaction:

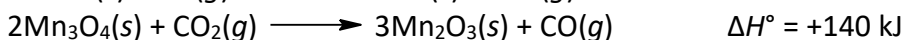
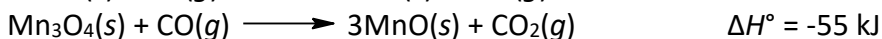
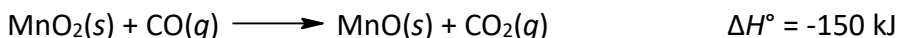


(-2007.5 kJ)

4) Determine ΔH° for the reaction

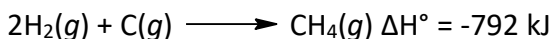


Given the following data:



(about -217 kJ)

5) Given that $\Delta H^\circ_f(\text{CH}_4(g)) = -75 \text{ kJ/mol}$ and



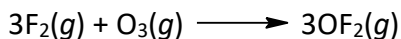
Calculate $\Delta H^\circ_{\text{sub}}$ for C(graphite) in kJ/mol. **(+717 kJ/mol)**

6) For oxygen difluoride, $\text{OF}_2(g)$, the Gibbs free energy of formation is +41 kJ/mol.

a) Is the preparation of $\text{OF}_2(g)$ from its elements at 25°C a spontaneous reaction? EXPLAIN.
(NO)

b) For ozone, $\text{O}_3(g)$, the Gibbs free energy of formation is +163 kJ/mol. Why isn't the Gibbs free energy of formation of ozone = 0? **(See your class notes)**

c) Is it theoretically possible to prepare $\text{OF}_2(g)$ at 25°C by the reaction:

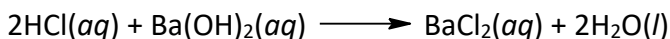


(YES)

7) The standard enthalpy of formation of $\text{CS}_2(l)$ at 298 K is +90 kJ/mol. The absolute molar entropy (S°) of C(graphite) at 298 K is 6 J/mol·K, of S(rhombic) is 32 J/mol·K, and of $\text{CS}_2(l)$ is 151 J/mol·K. Calculate the standard Gibbs free energy of formation of $\text{CS}_2(l)$ at 298 K.

($\Delta G^\circ_{f,298} \cong +66 \text{ kJ/mol}$)

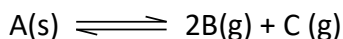
8) When 100 mL of 0.1250 M HCl at 24.00°C was mixed with 250 mL of 0.1000 M $\text{Ba}(\text{OH})_2$ at 24.00°C (assume the calorimeter constant = 0 J/°C), the temperature rose to 24.50°C. Calculate ΔH° for the reaction



Assume the specific heat for the solution is 4.184 J/g·°C. In calculating heat flow assume no heat is lost to the surroundings and ignore the small amount of water formed in the reaction.

(-117 kJ)

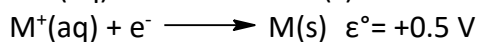
9) For the reaction



$$\Delta G^\circ_{298} = +23.0 \text{ kJ and } \Delta H^\circ_{298} = +150 \text{ kJ}$$

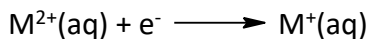
- Calculate the approximate value of K_p at 298 K. (**1×10^{-4}**)
- Calculate P_B and P_C at equilibrium at 298 K. (**$P_B \approx 6 \times 10^{-2}$ bar and $P_C \approx 3 \times 10^{-2}$ bar**)
- Calculate the approximate value of K_p at 500 K. (**1×10^6**)
- Calculate the absolute molar entropy of B at 298 K given that the absolute molar entropies of C(g) and A(s) are 200 J/mol·K and 150 J/mol·K respectively. (**$185 \text{ J/mol}\cdot\text{K}$**)

10) Given the following standard electrode potentials and Gibbs free energies of formation at 298 K:



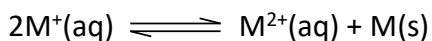
(You may use the value for the Faraday as 1×10^5 Coulombs for these calculations.)

- Calculate ϵ° for the reaction



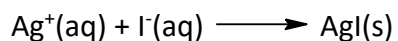
$$(\epsilon^\circ = +0.1 \text{ V})$$

- Calculate ΔG° and K_c at 298 K for the reaction



$$(\Delta G^\circ \approx -40 \text{ kJ and } K_c \approx 1 \times 10^7)$$

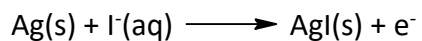
11) For the reaction



$$\varepsilon^\circ_{298} = +0.95 \text{ V and } \Delta G^\circ_{328} = -90 \text{ kJ}$$

(You may use the value for the Faraday as 1×10^5 Coulombs for these calculations.)

- Calculate ΔG° for the above reaction at 298 K. **(-95 kJ)**
- Calculate ΔS° and ΔH° for this reaction at 298 K. Assume that ΔS° and ΔH° are independent of temperature. **($\Delta S^\circ \approx -170 \text{ J/K}$ and $\Delta H^\circ \approx -146 \text{ kJ}$)**
- Given that the Gibbs free energy of formation of $\text{Ag}^+(\text{aq})$ at 298 K is +77 kJ/mol, calculate ΔG°_{298} and ε°_{298} for the half-reaction



$$\text{(}\varepsilon^\circ_{298} = +0.18 \text{ V and } \Delta G^\circ_{298} = -18 \text{ kJ)}$$