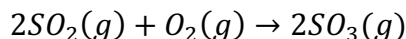


- One of the possible initial steps in the formation of acid rain is the oxidation of the pollutant SO_2 to SO_3 by the reaction:



Compound	ΔH_f° (kJ/mol)	ΔS° (J/mol*K)
SO_2 (g)	-296.8	248.2
O_2 (g)	0	205.2
SO_3 (g)	-395.8	256.8

- Calculate ΔG° and determine whether the reaction is product-favored at equilibrium at 25°C .

$$\begin{aligned}\Delta_r H^\circ &= \left[2 \text{ mol } \text{SO}_3 * -395.8 \frac{\text{kJ}}{\text{mol}} \right] - \left[\left(2 \text{ mol } \text{SO}_2 * -296.8 \frac{\text{kJ}}{\text{mol}} \right) + \left(1 \text{ mol } \text{O}_2 * 0 \frac{\text{kJ}}{\text{mol}} \right) \right] \\ &= -198 \frac{\text{kJ}}{\text{mol}}\end{aligned}$$

$$\begin{aligned}\Delta_r S^\circ &= \left[2 \text{ mol } \text{SO}_3 * 256.8 \frac{\text{J}}{\text{mol} * \text{K}} \right] \\ &\quad - \left[\left(2 \text{ mol } \text{SO}_2 * 248.2 \frac{\text{J}}{\text{mol} * \text{K}} \right) + \left(1 \text{ mol } \text{O}_2 * 205.2 \frac{\text{J}}{\text{mol} * \text{K}} \right) \right] \\ &= -188 \frac{\text{J}}{\text{mol} * \text{K}} = -0.188 \frac{\text{kJ}}{\text{mol} * \text{K}}\end{aligned}$$

$$\begin{aligned}\Delta_r G^\circ &= \Delta_r H^\circ - T \Delta_r S^\circ = -198 \frac{\text{kJ}}{\text{mol}} - (298.15 \text{ K}) \left(-0.188 \frac{\text{kJ}}{\text{mol} * \text{K}} \right) \\ &= -141.95 \frac{\text{kJ}}{\text{mol}}\end{aligned}$$

Now that we have calculated $\Delta_r G^\circ$, let's calculate K to find out whether it is reactant or product-favored.

$$\Delta_r G^\circ = -141.95 \frac{\text{kJ}}{\text{mol}} = -RT \ln K$$

$$-141.95 \frac{\text{kJ}}{\text{mol}} = - \left(8.314 * 10^{-3} \frac{\text{kJ}}{\text{mol} * \text{K}} \right) (298.15 \text{ K}) \ln K$$

$$K = 7.41 * 10^{24}$$

This K is massive and thus the reaction is product favored

- b. Estimate the temperature at which the reaction switches between product-favored and reactant-favored at equilibrium.

The Gibbs free energy will be 0 when the reaction switches between product-favored and reactant-favored at equilibrium. Thus,

$$\Delta_r G^\circ = 0 = \Delta_r H^\circ - T \Delta_r S^\circ$$

$$T = \frac{\Delta_r H^\circ}{\Delta_r S^\circ} = - \frac{198 \frac{\text{kJ}}{\text{mol}}}{-0.188 \frac{\text{kJ}}{\text{mol} * \text{K}}} = 1053 \text{ K}$$

2. For a certain chemical reaction, $\Delta H^\circ = -35.4 \text{ kJ}$ and $\Delta S^\circ = -85.5 \text{ J/K}$.
- a. Is the reaction exothermic or endothermic?

Change in enthalpy (ΔH) is negative, so reaction is exothermic.

- b. Does the reaction lead to an increase or decrease in the randomness/disorder of the system?

Change in entropy (ΔS) is negative, so there is a decrease in the randomness/disorder of system.

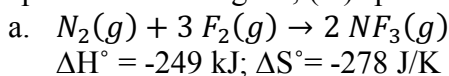
- c. Calculate the ΔG° for the reaction at 298 K.

$$\Delta_r G^\circ = \Delta H - T \Delta S = -35.4 \text{ kJ} - 298 \text{ K} * \left(\frac{-0.0855 \text{ kJ}}{\text{K}} \right) = -9.9 \text{ kJ}$$

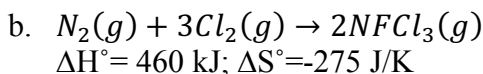
d. Is the reaction spontaneous at 298 K under standard conditions?

Yes, because Gibbs Free Energy is negative.

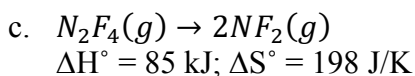
3. Classify each of the following reactions as one of the four possible types: (i) spontaneous at all temperatures, (ii) not spontaneous at any temperature; (iii) spontaneous at low T but not spontaneous at high T; (iv) spontaneous at high T but not spontaneous at low T.



Type (iii). It is spontaneous at **VERY** low temperatures but not spontaneous at high temperatures.



Type (ii). It is not spontaneous at any temperature, because we are always subtracting a negative number from a positive number.



Type (iv).

True or False

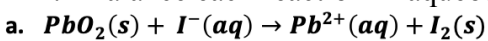
(**T** / **F**) For a process that occurs at constant temperature, the change in Gibbs free energy depends on changes in the enthalpy and entropy of the system.

This is true as seen in the equation $\Delta G = \Delta H - T\Delta S$

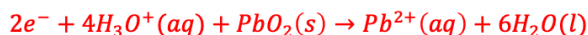
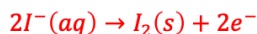
(**T** / **F**) If ΔG is large and negative for a certain reaction, the rate at which the reaction occurs is fast.

Activation energy will still determine the kinetics, ΔG determines the spontaneity.

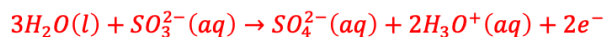
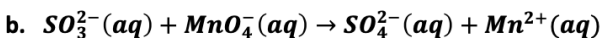
4. Balance each reaction in aqueous acidic solution.



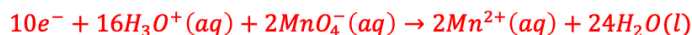
Balance each half reaction, first by mass and then by charge. Add H_3O^+ and H_2O where needed.



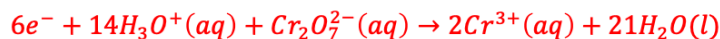
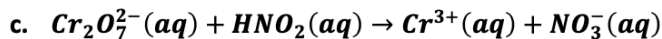
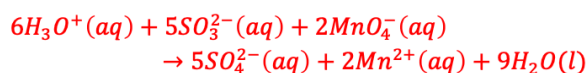
Since $2e^-$ are already on both sides, add the two half reactions together to get the overall balanced redox reaction:



Scale the half reactions to get the same number of electrons, then add them:



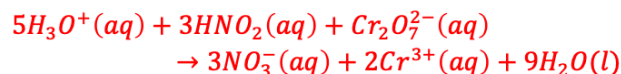
Overall reaction:

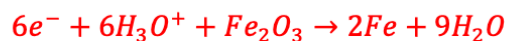


Multiply the second half reaction by 3 to get:



Add it to the first half reaction to get:

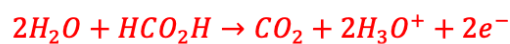
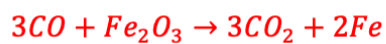




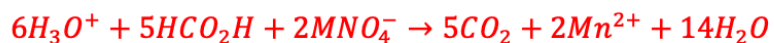
Multiply the second half reaction by 3:



Add the two half reactions to get:



Multiply the first by 5 and the second by 2 and then add:



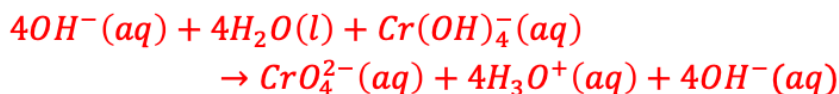
5. Balance each reaction in basic aqueous solution



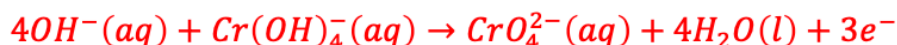
Balance each half reaction, first by mass and then by charge. Add OH^- and H_2O where needed. Balancing in basic solution is a bit trickier when deciding whether to add OH^- or H_2O to a given side. One thing that might work is to balance as if you were in acidic solution (add H_3O^+ to balance hydrogen) and then add an appropriate number of OH^- on both sides. Water should be created on the one side with H_3O^+ and OH^- will be left on the other side



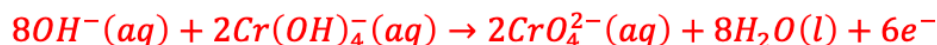
Add 4 OH^- to both sides:



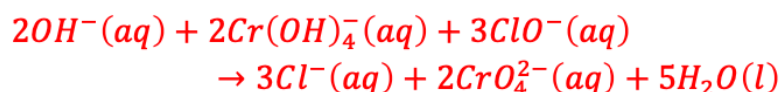
Simplify the equation and add electrons to balance charge:

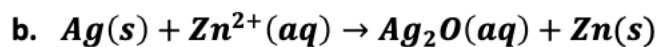


Multiply the first half reaction by 3 and the second by 2:

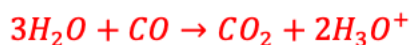
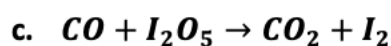


Add the two together to get the overall balanced redox equation:





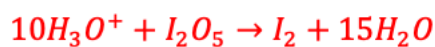
Add the two half reactions together:



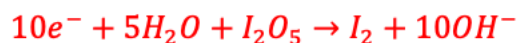
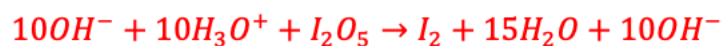
Add 2 OH⁻ to both sides



Now, the second half reaction:

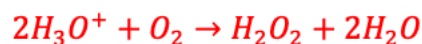
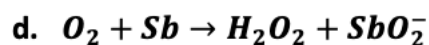


Add 10 OH⁻ to both sides:

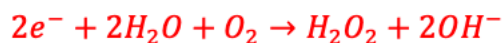
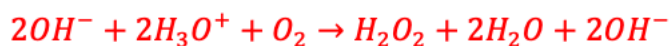


Multiply the first half reaction by 5 and add it to the second:





Add 2 OH⁻ to both sides:



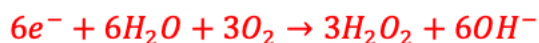
The second half reaction:



Add 4 OH⁻ to both sides:



Multiply the first half reaction by 3 and the second by 2:

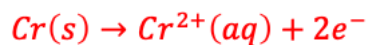


Add the two together to get the overall balanced redox reaction:



6. Consider a voltaic cell involving chromium (II) and gold (I)

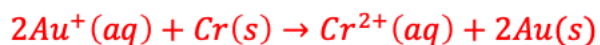
a. Balance the following reaction: $Au^+(aq) + Cr(s) \rightarrow Au(s) + Cr^{2+}(aq)$



Multiply the first half reaction by 2:



Add the two half reactions together:



- b. Sketch this cell, identifying the cathode and anode, the flow of electrons and the flow of cations and anions from the salt bridge (composed of $NaNO_3$).

