

Chem1B, General Chemistry II

1.) Radium-226 undergoes radioactive α -decay.a.) Write the **balanced** nuclear reaction.

Loses two neutrons and two protons (He)

b.) If the half-life of radium-226 is 1.25×10^9 years, calculate the amount of radium-226 remaining in a rock that is 1.5×10^9 years old and originally 15 kg.

Determine the rate constant for the first order reaction.

$$t_{1/2} = (0.693)/k, \text{ so } k = (0.693)/(1.25 \times 10^9 \text{ years}) = 5.544 \times 10^{-10} \text{ yr}^{-1}$$

For a first order reaction:

$$\ln(N_t/N_0) = - (5.544 \times 10^{-10} \text{ yr}^{-1})(1.5 \times 10^9 \text{ yr}) = -0.8316$$

$$N_t/N_0 = e^{-0.8316} = 0.435$$

$$N_t = (2.296)(N_0) = (0.435)(15 \text{ kg}) = 6.53 \rightarrow \mathbf{6.5 \text{ kg of radium-226}}$$

c.) Based on your answer from (b), how many kg of the produced element in (a) after decay are present after 1.5×10^9 years?

Determine how many kg of radon-226 were lost.

$$15 \text{ kg initial} - 6.5 \text{ kg remaining} = 8.5 \text{ kg lost}$$

Use mass numbers as mass ratio.

$$8.5 \text{ kg} = (226/222)(\text{kg Rn}) \rightarrow 8.3 \rightarrow \mathbf{8 \text{ kg radon-222 formed}}$$

2.) A solution of 0.10 M $\text{CuSO}_4(\text{aq})$ is electrolyzed for 1.00 hours by a current of 1.62 A. Calculate the amount of Cu(s), in g, deposited at the cathode.

Calculate the total coulombs put into the system.

$$(1.00 \text{ hours}) \times (60 \text{ min./1 hr.}) \times (60 \text{ sec./1 min.}) = 3600 \text{ sec}$$

$$1.62 \text{ A} = 1.62 \text{ C/sec} \times (3600 \text{ sec}) = 5832 \text{ C}$$

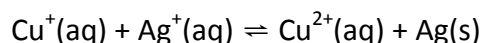
Convert to moles e^-

$$5832 \text{ C} \times (1 \text{ mole } e^- / 96485 \text{ C}) = 0.06044 \text{ moles } e^-$$

Convert to g Cu (recall: $\text{Cu}^{2+}(\text{aq}) + 2 e^- \rightarrow \text{Cu}(\text{s})$)

$$0.06044 \text{ moles } e^- \times \frac{1 \text{ mole Cu}}{2 \text{ moles } e^-} \times \frac{63.546 \text{ g Cu}}{1 \text{ mole Cu}} = 1.920 \rightarrow \mathbf{1.92 \text{ g Cu}}$$

3.) Consider the following equilibrium at 298K:

Calculate the reduction potential (E°_{red}) for $\text{Cu}^{2+}(\text{aq}) + e^- \rightarrow \text{Cu}^+(\text{aq})$, given that $\text{Ag}^+ + 1 e^- \rightarrow \text{Ag}(\text{s})$ has a reduction potential of +1.98 V and the equilibrium constant (K_{eq}) is 6.29×10^{30} .Find ΔG°

$$\Delta G^\circ = -RT \ln(K) = -(8.314 \text{ J/(mol}\cdot\text{K)}) (298 \text{ K}) (\ln (6.29 \times 10^{30})) = -1.757 \times 10^5 \text{ J}$$

Solve for E°_{cell}

$$\Delta G^\circ = -nF E^\circ_{\text{cell}}$$

$$-1.757 \times 10^5 \text{ J} = -(1 e^-)(96485 \text{ C/mol } e^-)(E^\circ_{\text{cell}}) \rightarrow E^\circ_{\text{cell}} = 1.821 \text{ V}$$

Solve for E°_{red} of Cu, behaving as **anode** (oxidized).

$$1.821 \text{ V} = 1.98 \text{ V} - E^\circ_{\text{red, Cu}} \rightarrow E^\circ_{\text{red, Cu}} = 1.98 \text{ V} - 1.821 \text{ V} = 0.158 \rightarrow \mathbf{0.16 \text{ V}}$$

Chem1B, General Chemistry II

4.) Consider the following equilibrium:

a.) Calculate ΔG° at 533 K, in kJ/mol, at equilibrium.At equilibrium, $\Delta G = 0$, so $\Delta G^\circ = -RT \ln(K)$

$$\text{Use } \Delta G^\circ = -RT \ln(K) = -(8.314 \text{ J/(mol}\cdot\text{K)}) (533 \text{ K}) (\ln(4.56 \times 10^8)) = -8.835 \times 10^4 \text{ J}$$

$$\Delta G^\circ = -88.4 \text{ kJ/mol}$$

b.) Calculate ΔS_{rxn} , given that $\Delta S^\circ \text{CO} = 197.7 \text{ J/K}$, $\Delta S^\circ \text{O}_2 = 205.1 \text{ J/K}$, and $\Delta S^\circ \text{CO}_2 = 213.7 \text{ J/K}$.

$$\text{Use } \Delta S_{\text{rxn}} = \sum m(S^\circ_{\text{products}}) - \sum n(S^\circ_{\text{reactants}})$$

$$\Delta S_{\text{rxn}} = [(2)(213.7 \text{ J/K})] - [(2)(197.7 \text{ J/K}) + (1)(205.1 \text{ J/K})]$$

$$\Delta S_{\text{rxn}} = [427.4 \text{ J/K}] - [600.5 \text{ J/K}] = -173.1 \rightarrow -173.1 \text{ J/K}$$

*Check: should be decreasing since 3 moles of gas become 2 moles of gas.*c.) Calculate the heat of formation of $\text{CO}_2\text{(g)}$ in kJ/mol, given that for CO , $\Delta H_f = -110.5 \text{ kJ/mol}$.Use $\Delta G = \Delta H - T\Delta S$ to solve for ΔH (heat of reaction).

$$-88.4 \text{ kJ/mol} = \Delta H - (533 \text{ K})(-173.1 \times 10^{-3} \text{ kJ/K})$$

$$\Delta H = -88.4 \text{ kJ/mol} - 92.26 \text{ kJ} = -180.66 \text{ kJ/mol}$$

Use $\Delta H_{\text{rxn}} = \sum m(H^\circ_{\text{products}}) - \sum n(H^\circ_{\text{reactants}})$ to solve for ΔH_f° for $\text{CO}_2\text{(g)}$

$$-180.66 \text{ kJ/mol} = [(2)(x)] - [(2)(-110.5 \text{ kJ/mol}) + (1)(0)]$$

$$2x = -401.66 \text{ kJ/mol}$$

$$x = -200.83 \rightarrow \Delta H_f^\circ = -200.8 \text{ kJ/mol}$$

5.) Given the following:

a.) Identify the **anode**.Cd (more negative reduction potential)b.) Identify the **cathode**.Au (more positive reduction potential)c.) Write the **balanced** overall redox reaction between Cd and Au.

Six electron transfer to balance, with Cd being oxidized and Au being reduced

d.) Calculate E°_{cell} , in V.

$$\text{Use } E^\circ_{\text{cell}} = E^\circ_{\text{red, cathode}} - E^\circ_{\text{ox, anode}} = +1.52 \text{ V} - (-0.403 \text{ V}) = 1.923 \rightarrow 1.92 \text{ V}$$

e.) Calculate ΔG° in kJ/mol for the above reaction.

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -(6 \text{ moles e}^-)(96485 \text{ C/mol e}^-)(1.92 \text{ J/C}) = -1.111 \times 10^6 \text{ J} \rightarrow$$

$$-1.11 \times 10^3 \text{ kJ/mol}$$

6.) Consider a radioactive atom of uranium-238 (mass: 238.0289 amu).

a.) Write the **balanced** nuclear equation for the α -decay of uranium-238.

α -decay is loss of 2 protons, 2 neutrons.



b.) Calculate the energy change, in J, associated with this nuclear reaction (the elemental product weighs 232.0381 amu and the decayed particle weighs 4.00150 amu, $m_e = 5.48580 \times 10^{-4}$ amu, $m_n = 1.00866$ amu, $m_p = 1.00728$ amu)

Use $\Delta E = c^2 (\Delta m)$

mass of U nucleus = 238.0289 amu - (92 e^-) (5.48580 $\times 10^{-4}$ amu/ e^-) = 237.97843 amu

mass of Th nucleus = 232.0381 amu - (90 e^-) (5.48580 $\times 10^{-4}$ amu/ e^-) = 231.98872 amu

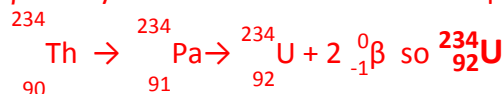
mass of He nucleus = 4.00150 amu (no e^-)

$\Delta m = (231.98872 \text{ amu} + 4.00150 \text{ amu}) - (237.97843 \text{ amu}) = -1.98820 \text{ amu} = -1.99820 \text{ g}$

$\Delta E = (2.9979 \times 10^8 \text{ m/s})^2 (-1.99820 \times 10^{-3} \text{ kg}) = -1.78687 \times 10^{14} \rightarrow -1.7869 \times 10^{14} \text{ J}$

c.) Uranium-238 can undergo sequential radioactive decay. Using your answer from (1), determine the final product when the product of (1) undergoes a further **two** β -decay reactions.

β -decay is the loss of 1 neutron \rightarrow 1 proton + 1 e^-



7.) Consider the following equilibrium:



a.) Calculate the equilibrium constant (K_p) for this reaction.

Use $\Delta G^\circ = -RT \ln(K) = 102.6 \times 10^3 \text{ J/mol} = - (8.314 \text{ J/(mol}\cdot\text{K)})(298 \text{ K})(\ln K)$

$\ln(K) = -41.411$, so $K = e^{-41.411} = 1.0356 \times 10^{-18} \rightarrow K_p = 1.036 \times 10^{-18}$

b.) Calculate the temperature, in K, at which this equilibrium will begin to be spontaneous.

Spontaneous = - ΔG , so find when $\Delta G < 0$ using $\Delta G = \Delta H^\circ - T\Delta S^\circ$

$0 = -67.68 \times 10^3 \text{ J/mol} - (T)(-120.6 \text{ J/K})$

$T = (67.68 \times 10^3 \text{ J/mol}) / (120.6 \text{ J/K}) = 561.19 \rightarrow \text{lower than } 561.2 \text{ K}$

c.) Using your answer from (b), find the partial pressure of $\text{NO}_2(\text{g})$ when the system is at equilibrium at that temperature if all the partial pressures are equal.

At equilibrium, $\Delta G = 0$. Find K at $T = 561.2 \text{ K}$.

$102.6 \times 10^3 \text{ J/mol} = - (8.314 \text{ J/(mol}\cdot\text{K)})(561.2 \text{ K})(\ln(K))$

$\ln(K) = -21.989 \rightarrow K = e^{-21.989} = 2.8182 \times 10^{-10}$

Use the equilibrium constant expression, where all pressures = x

$2.8182 \times 10^{-10} = \frac{(x)^2}{(x)^2} = x^{-1}$ so $x = 3.5482 \times 10^9 \rightarrow 3.548 \times 10^9 \text{ atm}$