UNITS 1 & 2 FREE RESPONSE REVIEW

General Chemistry

1. [5 POINTS]

A sample of an unknown hydrocarbon was burned in air.

(a) When the hydrocarbon sample was burned in a reaction that went to completion, 2.2 grams of water and 3.6 liters of carbon dioxide were produced under standard conditions. What is the empirical formula of the hydrocarbon? **[2 POINTS]**

$$3.6\ L\ CO_2 \times \frac{1\ mol\ CO_2}{22.4\ L\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} = 0.12\ mol\ C \quad \div 0.12 = 1$$

$$2.2\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.016\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} = 0.24\ mol\ H \quad \div 0.12 = 2$$

$\boxed{CH_2}$

(b) If the molar mass of the compound was determined to be ~~54.0~~ 28.1 g/mol, what is the molecular formula of the hydrocarbon? **[2 POINTS]**

$$\frac{28.1\ g/mol}{14.026\ g/mol} \approx 2$$

(molecular mass is twice as big as empirical) $\Rightarrow 2 \times CH_2 = \boxed{C_2H_4}$

(c) Write the balanced equation for the combustion reaction that took place in (a). **[1 POINT]**

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$$

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UNIT 3 FREE RESPONSE REVIEW
Electrochemistry

2. [9 POINTS]

A strip of Ni Metal is placed in a 1.0 M solution of $Ni(NO_3)_2$ and a strip of Ag metal is placed in a 1.0 M solution of $AgNO_3$. An electrochemical cell is created when the two solutions are connected by a salt bridge and the two metal strips are connected by wires to a voltmeter.

Standard Reduction Potentials at 25°C	
Reduction Half-Reaction	E°(V)
$Au^{3+} + 3e^- \rightarrow Au(s)$	+ 1.50
$Ag^+ + e^- \rightarrow Ag(s)$	+ 0.80
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	– 0.25
Na+ + $e^- \rightarrow Na(s)$	– 2.71

(a) Write the balanced chemical equation for the overall reaction that occurs in the cell, and calculate the cell potential, E°. **[2 POINTS]**

$$(Ag^+ + e^- \rightarrow Ag) \times 2 = 2Ag^+ + 2e^- \rightarrow 2Ag$$

$$+ \quad Ni \rightarrow Ni^{2+} + 2e^-$$

$$E^\circ_{cell} = E^\circ_{ox} + E^\circ_{red} = 0.25 + 0.80 = \boxed{+1.05\ V}$$

$$2Ag^+(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$$

(b) Calculate how many grams of metal will be deposited on the cathode if the cell is allowed to run at constant current of 1.5 amperes for 8.00 minutes. **[2 POINTS]**

2 s.f. 2 s.f.

$$8.00\ min \times \frac{60\ s}{1\ min} \times \frac{1.5\ C}{1\ s} \times \frac{1\ mol\ e^-}{96{,}485\ C} \times \frac{1\ mol\ Ag}{1\ mol\ e^-} \times \frac{107.87\ g\ Ag}{1\ mol} = \boxed{0.80\ g\ Ag}$$

$$Amp = \frac{C}{s}$$

are you SAFe? MMM

(c) Calculate the value of the standard free energy change, ΔG°, for the cell reaction. **[2 POINTS]**

$$\Delta G^\circ = -nFE^\circ_{cell}$$

↑ 2 b/c 2 mol of e^- are transferred in balanced rxn

$$= -\left(2\frac{mol\ e^-}{mol\ rxn}\right)\left(96{,}485\frac{C}{s}\right)(1.05\ V) = -203{,}000\frac{J}{mol_{rxn}} = \boxed{-203\frac{kJ}{mol_{rxn}}}$$

V = J/C

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(d) Predict: will the cell potential, E_{cell}, at 25°C for the cell shown above be higher, lower, or equal to the standard cell potential $E°_{cell}$, if the initial concentration of $Ni(NO_3)_2$ is 0.100 M and the initial concentration of $AgNO_3$ is 1.20 M. **[2 POINTS]**

Decreasing $[Ni^{2+}]$, a product, and increasing $[Ag^+]$, a reactant, will both cause Q to be lower than under std. conditions. Since the rxn is further from equilibrium, E_{cell} will be higher than $E°_{cell}$.

(e) Is the reaction in the cell thermodynamically favorable under conditions described in part (d)? Justify your answer. **[1 POINT]**

Yes, b/c the $E°_{cell}$ is positive, so E_{cell} will have an even higher (more positive) voltage, which indicates a thermodynamically favorable rxn.

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UNIT 4 FREE RESPONSE REVIEW
Thermochemistry and Thermodynamics

3. [8 POINTS]

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

The above reaction for the combustion of methane gas has a standard entropy change, ΔS°, with a value of −242.7J/mol•K. The following data are also available.

Compound	$\Delta H°_f$ (kJ/mol)
$CH_4(g)$	−74.8
$H_2O(l)$	−285.9
$CO_2(g)$	−393.5

(a) What are the values of $\Delta H°_f$ and $\Delta G°_f$ for $O_2(g)$? **[1 POINT]**

$\Delta H°_f$ and $\Delta G°_f$ for $O_2(g)$ are both equal to zero. The enthalpy and free energy change of formation of any element in its standard state is zero.

(b) Calculate the standard change in enthalpy, ΔH°, for the combustion of methane. **[2 POINTS]**

$\Delta H = \Sigma \Delta H_f \text{products} - \Sigma \Delta H \text{reactants}$

$= [(-393.5) + (2(-285.9))] - [(-74.8) + (2(0))]$

$= [-965.3 kJ] - [-74.8 kJ]$

$\boxed{\Delta H = -890.5 kJ}$

(c) Calculate the standard free energy change, ΔG°, for the combustion of methane. **[2 POINTS]**

$\Delta G° = \Delta H° - T\Delta S°$

$= (-890,500 J) - (298 K)(-242.7 J/K)$

$\boxed{\Delta G° = -818,200 J = -818.2 kJ}$

*When no temperature is given assume 298K (25°C).

(d) How would the value of ΔS° for the reaction be affected if the water produced in the combustion remained in the gas phase? **[1 POINT]**

ΔS° would become less negative/more positive. $H_2O(g)$ has more entropy than $H_2O(l)$. Gas phase is more disordered than liquid phase.

(e) A 20.0g sample of $CH_4(g)$ underwent combustion in a bomb colorimeter with excess oxygen gas.

(i) Calculate the mass of carbon dioxide produced. **[1 POINT]**

$20.0 g\ CH_4 \times \frac{1 \text{ mol } CH_4}{16 \text{ g } CH_4} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} \times \frac{44.0 \text{ g } CO_2}{1 \text{ mol } CO_2} = \boxed{55.0 g\ CO_2}$

(ii) Calculate the heat released by the reaction. **[1 POINT]**

$20.0 g\ CH_4 \times \frac{1 \text{ mol } CH_4}{16 \text{ g } CH_4} \times \frac{890.5 \text{ kJ}}{1 \text{ mol } CH_4} = \boxed{1{,}110 kJ \text{ of heat released}}$

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UNIT 5 FREE RESPONSE REVIEW
Kinetics

4. [8 POINTS]

$$A + 2B \rightarrow 2C$$

The following results were obtained in experiments designed to study the rate of the reaction above:

Experiment	Initial [A] (mol/L)	Initial [B] (mol/L)	Initial rate of disappearance of A (M/sec)
1	0.05	0.05	3.0×10^{-3}
2	0.05	0.10	6.0×10^{-3}
3	0.10	0.10	1.2×10^{-2}
4	0.20	0.10	2.4×10^{-2}

(a) Determine the order of the reaction with respect to A. Justify your response. **[1 POINT]**

Exp 3+4, when [A] is doubled and [B] is held constand the rate also doubles. Therefore, this reaction is first-order with respect to A.

(b) Determine the order of the reaction with respect to B. Justify your response. **[1 POINT]**

Exp 1+2, when [B] is doubled and [A] is held constant the rate also doubles. Therefore, this reaction is first-order with respect to B.

(c) Write the rate law for the reaction. **[1 POINT]**

Rate = k[A][B]

(d) Calculate the value of the rate constant, k, for the reaction. Include the units. **[2 POINTS]**

Rate = k[A][B]

$$k = \frac{Rate}{[A][B]} = \frac{(1.20 \times 10^{-2}\ M/sec)}{(0.10M)(0.10M)} = 1.2\ M^{-1}sec^{-1}$$

* You could use ANY experiment.

(e) If another experiment is attempted with [A] and [B], both 0.02-molar, what would be the initial rate of disappearance of A? **[1 POINT]**

Rate = k[A][B]

$$= (1.2\ M^{-1}sec^{-1})(0.02M)(0.02M)$$

$$= 4.8 \times 10^{-4}\ M/sec$$

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UNIT 5 FREE RESPONSE REVIEW
Kinetics

(f) The following reaction mechanism was proposed for the reaction above:

$$A + B \rightarrow C + D$$

$$D + B \rightarrow C$$

(i) Show the mechanism is consistent with the balanced reaction. **[1 POINT]**

Add the reactions together

$A + B \rightarrow C + \cancel{D}$
$\cancel{D} + B \rightarrow C$
$A + 2B \rightarrow 2C$ → original reaction

(ii) Show which step is the rate-determining step, and explain your choice. **[1 POINT]**

The first step of the mechanism is the slow, rate-determining step b/c its rate law is the same as the experimentally determined rate law.

$Rate = k[A][B]$

(iii) In the mechanism, would D be classified as a catalyst or an intermediate? Justify your response. **[1 POINT]**

D is an intermediate b/c it is produced in an early step and consumed in a later step.

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UNIT 6 FREE RESPONSE REVIEW
Equilibrium

5. [7 POINTS]

At 25°C, the solubility product constant, K_{sp}, for nickel hydroxide, $Ni(OH)_2$, is $1.6x10^{-14}$.

(a) Write a balanced equation for the solubility equilibrium for $Ni(OH)_2$. **[1 POINT]**

$Ni(OH)_2(s) \rightleftharpoons Ni^{2+}(aq) + 2OH^-(aq)$

(b) What is the molar solubility of $Ni(OH)_2$ in pure water at 25°C? **[2 POINTS]**

$K_{sp} = [Ni^{2+}][OH^-]^2$

$1.6x10^{-14} = (x)(2x)^2$

$1.6x10^{-14} = 4x^3$

$x = 1.6x10^{-5} M$ → Molar solubility of $Ni(OH)_2$. Same as $[Ni^{2+}]$ b/c ratio is 1:1.

* One Ni^{2+} is produced for every $Ni(OH)_2$ so the molar solubility of $Ni(OH)_2$ will be the same as the concentration of Ni^{2+}!

(c) Predict whether a precipitate will form when 200.0 milliliters of a $5.0x10^{-5}$ M KOH solution is mixed with 300.0 milliliters of a $2.0x10^{-4}$ molar $Ni(NO_3)_2$ solution at 25°C. Show calculations to support your prediction. **[2 POINTS]**

STEP #1
Moles of OH^- ⇒ $M = \frac{mol}{L}$ ⇒ $5.0x10^{-5} M = \frac{x}{0.200L}$ ⇒ $1.0x10^{-5}$ mol OH^-

Moles of Ni^{2+} ⇒ $M = \frac{mol}{L}$ ⇒ $2.0x10^{-4} M = \frac{x}{(0.300L)}$ ⇒ $6.0x10^{-5}$ mol (Ni^{2+})

STEP #2
Add the volumes
$(0.200L + 0.300L) = 0.500L$

STEP #3
$[Ni^{2+}] = \frac{6.0x10^{-5} mol}{(0.500L)} = 1.2x10^{-4} M$

$[OH^-] = \frac{1.0x10^{-5} mol}{(0.500L)} = 2.0x10^{-5} M$

STEP #4
$Q = [Ni^{2+}][OH^-]^2$

$Q = (1.2x10^{-4} M)(2.0x10^{-5})^2 = 4.8x10^{-14}$

K < Q, so precipitate will form.

(d) At 25°C, 100mL of a saturated $Ni(OH)_2$ solution was prepared.

(i) Calculate the mass of $Ni(OH)_2$ present in the solution. **[1 POINT]**

$\frac{1.6x10^{-5} mol}{L} \times \frac{92.69g}{1 mol} = \frac{0.0015 g}{L}$ ⇒ $(0.0015 g/L)(0.100L) = 0.00015 g$

(ii) If the solution is allowed to evaporate to a final volume of 50mL, what will the $[OH^-]$ be at this volume? **[1 POINT]**

B/c the solution is saturated, evaporation will NOT change the concentrations.

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