Unit 6 Review for Midterms

Thermodynamics

You should be able to:

- Understand and explain exothermic and endothermic processes
- Explain how the strength of intermolecular forces determines if forming a solution is endo- or exothermic
- Describe an endo- or exothermic process with an energy diagram
- Explain the relationship between collisions and thermal energy transfer
- Define thermal equilibrium
- Perform calorimetry calculations for a one-species system and two-species system
- Use the positive or negative signs for heat accurately
- Explain changes in heat absorbed or released by a system undergoing phase changes
- Use the terms vaporization, fusion, condensation, solidification correctly
- Understand the relationship between heat energy and enthalpy
- Calculate the enthalpy change of a reaction based on bond energies
- Calculate enthalpy change of a reaction based on the standard enthalpies of formation
- Use Hess's law
- Manipulate equations and enthalpies

1. Upon adding solid potassium hydroxide pellets to water, the following reaction takes place:

   $KOH(s) \rightarrow K^+ + OH^- + 43$ kJ/mol

   Answer the following questions regarding the addition of 14.0 g of KOH to water:
   a. Does the beaker get warmer or colder?
   b. Is the reaction endothermic or exothermic?
   c. What is the enthalpy change for the dissolution of the 14.0 g of KOH?
2. When 1.00 L of 1.00 M $Ba(NO_3)_2$ solution at 25.0˚C is mixed with 1.00 L of 1.00 M Na_2SO_4 solution at 25.0˚C in a calorimeter, the white solid $BaSO_4$ forms and the temperature of the mixture increase to 28.1˚C.
   a. Write the balanced equation for this reaction.
   b. Write the net-ionic equation.
   c. Assuming that the calorimeter absorbs only a negligible quantity of heat, and that the specific heat capacity of the solution is 4.18 J/g˚C, and that the density of the final solution is 1.00 g/mL, calculate the enthalpy change per mole of $BaSO_4$ formed.
3. When 1 mole of methane (CH_4) is burned, 890 kJ/mol of energy is released as heat.
   a. Write the balanced thermochemical equation for this reaction.
   b. Calculate ΔH for a process in which a 5.8 gram sample of methane is burned.
4. Calculate the total energy needed to turn 89.70 grams of ice at -40.00 °C into steam at 350.0 °C. (ΔH_{fus} = 6.01 kJ/mol, C_{liquid} = 4.184 J/g˚C, ΔH_{vap} = 40.7 kJ/mol, C_{gas} = 2.02 J/g˚C, C_{ice} = 2.11 J/ g˚C)
5. Given the information below, calculate the ΔH°_{rxn} for the following reaction:

$$3\ Al(s) + 3\ NH_4ClO_4(s) \rightarrow Al_2O_3(s) + AlCl_3(s) + 3\ NO(g) + 6\ H_2O(g)$$

Substance	ΔH_f° (kJ/mol)
$NH_4ClO_4(s)$	−295
$Al_2O_3(s)$	−1676
$AlCl_3(s)$	−704
$NO(g)$	90.0
$H_2O(g)$	−242

6. Calculate the H for this overall reaction $2\ H_3BO_3(aq) \rightarrow B_2O_3(s) + 3\ H_2O(\ell)$ given the following equations:

$$H_3BO_3(aq) \rightarrow HBO_2(aq) + H_2O(\ell) \qquad \Delta H = -0.02\ kJ/mol_{rxn}$$

$$H_2B_4O_7(aq) + H_2O(\ell) \rightarrow 4\ HBO_2(aq) \qquad \Delta H = -11.3\ kJ/mol_{rxn}$$

$$H_2B_4O_7(aq) \rightarrow 2\ B_2O_3(s) + H_2O(\ell) \qquad \Delta H = 17.5\ kJ/mol_{rxn}$$

7. For the following reaction and data given below:

$$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$$

Bond Type	Bond Energy
H−H	432 kJ/mol
F−F	154 kJ/mol
H−F	565 kJ/mol

   a. Draw dot structures for all reactants and products.
   b. Use the bond energies to calculate the enthalpy of the reaction.

8.

AP Questions

1.

$$O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$$

Consider the reaction represented above.

(a) Referring to the data in the table below, calculate the standard enthalpy change, ΔH°, for the reaction at 25°C. Be sure to show your work.

	$O_3(g)$	$NO(g)$	$NO_2(g)$
Standard enthalpy of formation, ΔH_f°, at 25°C (kJ mol^{-1})	143	90.	33

(d) Use the information in the table below to write the rate-law expression for the reaction, and explain how you obtained your answer.

Experiment Number	Initial $[O_3]$ (mol L^{-1})	Initial [NO] (mol L^{-1})	Initial Rate of Formation of NO_2 (mol L^{-1} s^{-1})
1	0.0010	0.0010	x
2	0.0010	0.0020	$2x$
3	0.0020	0.0010	$2x$
4	0.0020	0.0020	$4x$

(e) The following three-step mechanism is proposed for the reaction. Identify the step that must be the slowest in order for this mechanism to be consistent with the rate-law expression derived in part (d). Explain.

Step I: $O_3 + NO \rightarrow O + NO_3$

Step II: $O + O_3 \rightarrow 2\ O_2$

Step III: $NO_3 + NO \rightarrow 2\ NO_2$

2.

Answer the following questions that relate to the chemistry of nitrogen.

(a) Two nitrogen atoms combine to form a nitrogen molecule, as represented by the following equation.

$$2\ N(g) \rightarrow N_2(g)$$

Using the table of average bond energies below, determine the enthalpy change, ΔH, for the reaction.

Bond	**Average Bond Energy (kJ mol^{-1})**
N – N	**160**
N = N	**420**
N ≡ N	**950**

(b) The reaction between nitrogen and hydrogen to form ammonia is represented below.

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g) \qquad \Delta H^\circ = -92.2\ kJ$$

(d) When $N_2(g)$ and $H_2(g)$ are placed in a sealed container at a low temperature, no measurable amount of $NH_3(g)$ is produced. Explain.

3.

$$2\ Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s) \qquad \Delta H^\circ_f = -824\ kJ\ mol^{-1}$$

Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 75.0 g sample of Fe(s) is mixed with 11.5 L of $O_2(g)$ at 2.66 atm and 298 K.

(a) Calculate the number of moles of each of the following before the reaction begins.

(i) Fe(s)
(ii) $O_2(g)$

(b) Identify the limiting reactant when the mixture is heated to produce $Fe_2O_3(s)$. Support your answer with calculations.

(c) Calculate the number of moles of $Fe_2O_3(s)$ produced when the reaction proceeds to completion.

The reaction represented below also produces iron(III) oxide. The value of ΔH° for the reaction is −280. kJ per mole of $Fe_2O_3(s)$ formed.

$$2\ FeO(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$$

Calculate the standard enthalpy of formation, , ΔH°_f, of FeO(s).