CHEMISTRY 123-07

Practice exam #4 – answer key November 01, 2010

PART I: MULTIPLE CHOICE

- 1. An endothermic reaction causes the surroundings to
 - a. warm up.
 - b. become acidic.
 - c. condense.
- 2. Which of the following is an example of potential energy?
 - a. Hitting a baseball
 - b. Running around bases
 - c. Pitching a baseball
- 3. The First Law of Thermodynamics states that:
 - a. Molecules move faster as temperature increases.
 - b. The total energy of the universe is constant.
 - c. Energy transfers from hotter objects to cooler objects.
 - d. Samples with different temperatures that come in contact with one another will reach thermal equilibrium.
 - e. All of the above.

4. Which property can be used to distinguish one substance from another substance?

- a. Temperature
- b. Enthalpy
- c. Internal energy
- 5. Which term refers to a quantity of heat transferred at constant pressure?
 - a. Entropy
 - b. Enthalpy
 - c. Work
- 6. If a 10.0 g sample of each substance below has 250 J applied to it, which substance will have the greatest increase in temperature?
 - a. Iron (specific heat = $0.46 \text{ J/g.}^{\circ}\text{C}$)
 - b. Water (specific heat = $4.184 \text{ J/g.}^{\circ}\text{C}$)
 - c. Copper (specific heat = $0.39 \text{ J/g.}^{\circ}\text{C}$)
 - d. <u>Aluminum (specific heat = $0.92 \text{ J/g.}^{\circ}\text{C}$)</u>
 - e. Lead (specific heat = $0.13 \text{ J/g.}^{\circ}\text{C}$)
- 7. The temperature of a 15.5 g sample of a metal increases by 35.0 °C when 125 J of energy is applied to it. What is the identity of the metal?
 - a. Silver (specific heat = $0.23 \text{ J/g.}^{\circ}\text{C}$)
 - b. Copper (specific heat = $0.39 \text{ J/g.}^{\circ}\text{C}$)
 - c. Iron (specific heat = $0.46 \text{ J/g.}^{\circ}\text{C}$)
 - d. Lead (specific heat = $0.13 \text{ J/g.}^{\circ}\text{C}$)
 - e. Aluminum (specific heat = $0.92 \text{ J/g.}^{\circ}\text{C}$)
- 8. Copper metal has a specific heat of 0.385 J/g⋅°C. Calculate the amount of heat required to raise the temperature of 22.8 g of Cu from 20.0°C to 875°C.

a.	$1.97 \times 10^{-5} \text{ J}$	d.	7.51 kJ
b.	$1.0 \times 10^{-2} \text{ J}$	e.	10.5 kJ
c.	329 J		

9. A glass containing 200. g of H_2O at 20°C was placed in a refrigerator. The water loses 11.7 kJ as it cools to a constant temperature. What is its new temperature? The specific heat of water is 4.184 J/g·°C.

a.	0.013°C	d.	14°C
b.	<u>4°C</u>	e.	34°C

c. 6°C

- d. decrease in temperature.
- e. release CO₂
- d. A bat lying on the ground
- e. Sliding into home plate

d. Specific heat capacity

Specific heat capacity

e. Expansion

e. Kinetic energy

d.

- 10. Suppose a 50.0 g block of silver (specific heat = $0.2350 \text{ J/g} \cdot ^{\circ}\text{C}$) at 100°C is placed in contact with a 50.0 g block of iron (specific heat = $0.4494 \text{ J/g} \cdot ^{\circ}\text{C}$) at 0°C, and the two blocks are insulated from the rest of the universe. The final temperature of the two blocks.
 - a. will be higher than 50°C.
 - will be lower than 50°C. b.
 - c. will be exactly 50°C.
 - d. is unrelated to the composition of the blocks.
 - e. cannot be predicted.
- 11. A 0.1326 g sample of magnesium was burned in an oxygen bomb calorimeter. The total heat capacity of the calorimeter plus water was 5,760 J/°C. If the temperature increase of the calorimeter with water was 0.570°C, calculate the enthalpy of combustion of magnesium.

d.

e.

106 kJ/mol

-602 kJ/mol

$$Mg(s) + 1/2O_2(g) \rightarrow MgO(s)$$

a. -3280 kJ/mol

- b. -24.8 kJ/mol
- c. 435 kJ/mol
- 12. To which one of these reactions occurring at 25°C does the symbol ΔH_f^o [H₂SO₄(1)] refer?
 - a. $2H(g) + S(g) + 4O(g) \rightarrow H_2SO_4(l)$
 - b. $H_2(g) + S(g) + 2O_2(g) \rightarrow H_2SO_4(l)$
 - c. $H_2SO_4(l) \rightarrow H_2(g) + S(s) + 2O_2(g)$
 - d. $H_2SO_4(l) \rightarrow 2H(g) + S(s) + 4O(g)$
 - e. $H_2(g) + S(s) + 2O_2(g) \rightarrow H_2SO_4(l)$
- 13. Octane (C_8H_{18}) undergoes combustion according to the following thermochemical equation:

 $\Delta H_{rxn}^{o} = -11.020 \text{ kJ/mol}$ $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

Given that $\Delta H_f^o[CO_2(g)] = -393.5 \text{ kJ/mol}$ and $\Delta H_f^o[H_2O(l)] = -285.8 \text{ kJ/mol}$, calculate the standard enthalpy of formation of octane.

- a. -210 kJ/mol
- b. -11,230 kJ/mol
- c. 22,040 kJ/mol
- d. -420 kJ/mol
- e. 420 kJ/mol

14. Given $2Al(s) + (3/2)O_2(g) \rightarrow Al_2O_3(s)$, $\Delta H_f^o = -1,670 \text{ kJ/mol for } Al_2O_3(s)$.

Determine ΔH° for the reaction $2Al_2O_3(s) \rightarrow 4Al(s) + 3O_2(g)$

- 3,340 kJ/mol a.
- b. 1,670 kJ/mol
- c. -3,340 kJ/mol
- d. -1,670 kJ/mol
- e. -835 kJ/mol
- 15. Given the thermochemical equation $2SO_2 + O_2 \rightarrow 2SO_3$, $\Delta H_{rxn}^o = -198$ kJ/mol, what is the standard enthalpy change for the decomposition of one mole of SO₃?
 - a. 198 kJ/mol -99 kJ/mol

- d. 396 kJ/mol
- e. -198 kJ/mol

99 kJ/mol c.

b.

- 16. For the reaction C(graphite) + $O_2(g) \rightarrow CO_2(g)$, $\Delta H^\circ = -393$ kJ/mol. How many grams of C(graphite) must be burned to release 275 kJ of heat?
 - a. 22.3 g d. 17.1 g b. 0.70 g 8.40 g e.
 - c. 12.0 g

- 17. Find the heat absorbed from the surroundings when 15 g of O₂ reacts according to the equation $O + O_2 \rightarrow O_3$, $\Delta H_{rxn}^o = 103$ kJ/mol.
 - a. $4.6 \times 10^{-3} \text{ kJ}$ d. 32 kJ

 b. 48 kJ e. 110 kJ

 c. 96 kJ e. 110 kJ
- 18. At 25°C, the standard enthalpy of formation of KCl(s) is -435.87 kJ/mol. When one mole of KCl(s) is formed by reacting potassium vapor and chlorine gas at 25°C, the standard enthalpy of reaction is -525.86 kJ/mol. Find Δ H° for the sublimation of potassium, K(s) \rightarrow K(g), at 25°C.
 - a. -345.88 kJ/mol
 - b. 45.00 kJ/mol
 - c. <u>345.88 kJ/mol</u>
 - d. 89.99 kJ/mol
 - e. -525.86 kJ/mol
- 19. 10.1 g CaO is dropped into a styrofoam coffee cup containing 157 g H₂O at 18.0°C. If the following reaction occurs, then what temperature will the water reach, assuming that the cup is a perfect insulator and that the cup absorbs only a negligible amount of heat? [specific heat of water = $4.18 \text{ J/g} \cdot ^{\circ}\text{C}$]

 $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$ $\Delta H = -64.8 \text{ kJ/mol}$

- a. <u>18.02°C</u>
- b. 35.8°C
- c. 311°C
- d. 42.2°C
- e. 117°C
- 20. A gas is compressed in a cylinder from a volume of 20 L to 2.0 L by a constant pressure of 10.0 atm. Calculate the amount of work done on the system.

d. $-1.81 \times 10^4 \text{ J}$

e. 180 J

- a. $1.01 \times 10^4 \,\mathrm{J}$
- b. -180 J
- c. $1.81 \times 10^4 \text{ J}$
- 21. Which of the following processes always results in an increase in the energy of a system?
 - a. The system loses heat and does work on the surroundings.
 - b. The system gains heat and does work on the surroundings.
 - c. The system loses heat and has work done on it by the surroundings.
 - d. The system gains heat and has work done on it by the surroundings.
 - e. None of these is always true.
- 22. For which of these reactions will the difference between ΔH° and ΔE° be the smallest?
 - a. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
 - b. $4PH_3(g) \rightarrow P_4(g) + 6H_2(g)$
 - c. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
 - d. $CO_2(g) + 2H_2O(l) \rightarrow CH_4(g) + 2O_2(g)$
 - e. $P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$

PART II: CONCEPTS

23. For each of the following, determine the value of ΔH_2 in terms of ΔH_1 .

 $A + B \longrightarrow 2C \qquad \Delta H_1$ a. $2C \longrightarrow A + B \qquad \Delta H_2 = ?$

<u>Answer:</u> The second equation is obtained by reversing the first one. Therefore $\Delta H_2 = -\Delta H_1$

 $A \longrightarrow B + 2C \qquad \Delta H_1$ b. $1/2 B + C \longrightarrow 1/2 A \qquad \Delta H_2 = ?$

<u>Answer:</u> The second equation is obtained by reversing the first one and dividing it by 2. Therefore $\Delta H_2 = -\frac{1}{2} \Delta H_1$

PART III: HESS'S LAW

24. Calculate ΔH for the following reaction:

 $\operatorname{Fe}_2O_3(s) + 3\operatorname{CO}(g) \longrightarrow 2\operatorname{Fe}(s) + 3\operatorname{CO}_2(g)$

Use the following reactions:

1) 2 Fe (s)	+	$3/2 O_2(g) \longrightarrow Fe_2O_3(s)$	$\Delta H = -824.2 \text{ kJ}$
2) CO (g)	+	$\frac{1}{2}O_2(g) \longrightarrow CO_2(g)$	$\Delta H = -282.7 \text{ kJ}$

<u>Answer:</u> We need to reverse equation (1) since Fe_2O_3 (s) is on the reactant side in the equation in question. Equation (2) needs to be multiplied by 3 since the studied equation involves 3 moles of CO (g). As a result:

1)
$$\operatorname{Fe}_2O_3(s) \longrightarrow 2 \operatorname{Fe}(s) + 3/2 O_2(g) \qquad \Delta H = +824.2 \text{ kJ}$$

2) $3 \operatorname{CO}(g) + 3/2 O_2(g) \longrightarrow 3 \operatorname{CO}_2(g) \qquad \Delta H = -848.1 \text{ kJ}$
Add equations (1) and (2):
 $\operatorname{Fe}_2O_3(s) + 3 \operatorname{CO}(g) + 3/2 O_2(g) \longrightarrow 2 \operatorname{Fe}(s) + 3/2 O_2(g) + 3 \operatorname{CO}_2(g)$
 $\operatorname{Fe}_2O_3(s) + 3 \operatorname{CO}(g) \longrightarrow 2 \operatorname{Fe}(s) + 3 \operatorname{CO}_2(g)$

 $\Delta H = + 824.2 \text{ kJ} + (-848.1 \text{ kJ}) = -23.9 \text{ kJ}$

PART IV: STANDARD MOLAR ENTHALPIES OF FORMATION

25. Write an equation for the formation of each of the following compounds from their constituent elements, in their standard states.

a. NO₂ (g) $\frac{1}{2} N_2 (g) + O_2 (g) \longrightarrow NO_2 (g)$ b. C₂H₄ (g) c. MgCO₃ (s) Mg (s) + C (graphite) + 3/2 O₂ (g) \longrightarrow MgCO₃ (s)

d. CH₃OH (1)

 $C (graphite) + 2 H_2 (g) + \frac{1}{2} O_2 (g) CH_3OH (l)$

26. Use data on ΔH_f° (Table 6.2 and Appendix 4) to calculate ΔH for each of the following reactions.

a. $2 H_2S(g) + 3 O_2(g) \rightarrow 2 H_2O(l) + 2 SO_2(g)$

<u>Answer:</u> We know that, in general, ΔH for any reaction can be determined if we know the ΔH_f^o values of all reactants and products, using the equation:

 $\Delta H = \Sigma n \times \Delta H_{f}^{o} (products) - \Sigma m \times \Delta H_{f}^{o} (reactants)$

From Table 6.2:	$\Delta H_{f}^{o} (H_{2}O(l)) = -286 \text{ kJ/mol}$
From Appendix 4:	$\Delta H_{f}^{o}(H_{2}S(g)) = -21 \text{ kJ/mol}$ $\Delta H_{f}^{o}(SO_{2}(g)) = -297 \text{ kJ/mol}$

We also know that for elements, in their standard states, $\Delta H_{f}^{o} = 0 \text{ kJ/mol}$. Therefore $\Delta H_{f}^{o} (O_{2} (g)) = 0 \text{ kJ/mol}$

Using the above equation:

 $\Delta H = 2 x (-286) + 2 x (-297) - 2 x (-21) = -1124 kJ$ $\Delta H = -1124 kJ$

b. $N_2O_4(g) + 4H_2(g) \longrightarrow N_2(g) + 4H_2O(g)$

Answer: We use the same general equation:

$\Delta H = \Sigma n x \Delta H_{f}^{o} (products) - \Sigma m x \Delta H_{f}^{o} (reactants)$

From Appendix 4: $\Delta H_f^{o}(N_2O_4(g)) = +10 \text{ kJ/mol}$

For elements, in their standard states, $\Delta H_{f}^{o} = 0$ kJ/mol. Therefore

 $\begin{array}{l} \Delta H_{f}^{\;o}\left(H_{2}\left(g\right)\right)=0\;kJ/mol\\ \Delta H_{f}^{\;o}\left(N_{2}\left(g\right)\right)=0\;kJ/mol \end{array}$

Using the above equation:

 $\Delta H = 4 \text{ x} (-242) - 1 \text{ x} 10 = -978 \text{ kJ}$ $\Delta H = -978 \text{ kJ}$