Chapter 10: Thermochemistry Study Guide

Key Objectives/ Main things I need to know for the test…

1. Vocabulary: heat, temperature, 1st Law of Thermodynamics, specific heat capacity, exothermic, endothermic, potential energy, kinetic energy, thermal energy, chemical energy, calorie, joule, calorimeter, heat of reaction, state function, bond energy, activation energy, catalyst, equilibrium, enthalpy, Hess's law, heat of formation, entropy, Gibb's Free energy

2. Collision theory and potential energy diagrams

3. Rates of Reactions and understanding of basic equilibrium.

4. Enthalpy and state functions

5. Distinguish between heat and temperature and are they intensive or extensive properties

6. Define the terms system and surroundings as they relate to chemical processes.

7. Distinguish between endothermic and exothermic reactions and the sign of ΔH

8. Apply the concept of specific heat capacity to examples.

9. Solve enthalpy calculations using:
   a. Stoichiometry
   b. Calorimetry
   c. Hess's Law
   d. Heats of Formation
   e. Bond Enthalpies

10. Solve calculation for Entropy and Gibbs Free Energy and how to determine if a reaction is spontaneous or not.

Practice

Matching:	Calorimetry	Chemical energy
Hess's Law	Heat of reaction	Catalyst
Enthalpy	Activation energy	1st law of thermodynamics
Specific heat capacity	Equilibrium	

1. ________________________________ the energy associated with chemical bonds and attractions between particles of a system.
2. ________________________________ states that regardless of the multiple stages or steps of a reaction, the total enthalpy change for the reaction is the sum of all changes.
3. ________________________________ the minimum amount of energy needed to initiate a reaction.
4. ________________________________ the state in which the rates of opposing processes are equal.
5. ________________________________ a substance that speeds a chemical reaction by lowering the activation energy of the reaction.
6. ________________________________ energy can be transformed from one form to another but it cannot be created nor destroyed.
7. ________________________________ process that measures of the amount of heat released or absorbed in a chemical reaction, change of state, or formation of a solution.
8. ________________________________ the heat absorbed or released as a result of a chemical reaction.
9. ________________________________ a thermodynamic quantity equivalent to the total heat content of a system. It is equal to the internal energy of the system plus the product of pressure and volume.
10. ________________________________ the amount of heat required to raise the temperature of 1 gram of a substance 1oC.

True or False: Write T or F next to each statement. (if false, correct to make true)

11. The container of an endothermic reaction will feel hot when the reaction is complete.
12. Energy is released when breaking the bonds in chemical reaction.

13. The enthalpy of the products in an exothermic reaction is greater than the enthalpy of the reactants.

14. If the equation of an exothermic reaction is reversed, the sign of ΔH becomes positive.

15. In a calorimeter, the heat lost in a reaction is less than the heat gained by the water.

16. For a given substance, the kinetic energy increases in the following order: gas→ liquid → solid

17. The reaction $NH_4Cl(s)$ + heat → $NH_3(g)$ + $HCl(g)$, is an example of an endothermic reaction.

Multiple-Choice:

18. What does thermochemistry study?
    a. Rate of reaction
    b. completeness of reaction
    c. Heat effects in chemical reactions
    d. mole ratios

19. Which serves as a unit of heat?
    a. Degree Celsius
    b. Kelvin
    c. Pascal
    d. Joule

20. An exothermic reaction
    a. Is indicated by a positive enthalpy change
    b. Requires added energy to proceed
    c. Releases heat to its surroundings
    d. Does not involve energy transfer

21. In a combustion reaction, the energy of the products is
    a. Less than the energy required of the reactants
    b. More than the energy required of the reactants
    c. Equal to the energy required of the reactants
    d. Impossible to determine

22. Hess's law makes it possible to
    a. Balance tricky equations
    b. Indirectly determine the enthalpy change of a reaction
    c. Predict the products in exothermic reactions
    d. Determine the specific heat of a substance with a calorimeter

23. The size of a temperature increase in a substance depends primarily upon the
    a. Specific heat and mass of the substance
    b. Original temperature of the substance
    c. Size of the calorimeter
    d. Temperature of the surroundings

24. The breakdown of glucose in your body is
    a. An endothermic reaction
    b. An exothermic reaction
    c. An unnecessary reaction
    d. The body's way of storing energy

25. Which of the following change of state operations is exothermic?
    a. Gas to liquid
    b. solid to liquid
    c. liquid to gas
    d. solid to gas

26. What happens if ΔH if a chemical equation is reversed?
    a. It becomes zero
    b. it changes in sign
    c. It remains the same
    d. depends on the reaction

Equation	***ΔH° (kJ)***
(1) $W(s) \rightarrow W(g)$	+60
(2) $W(g) + X(g) \rightarrow Y(g)$	–40
(3) $2Y(g) \rightarrow Z(g)$	–90
(4) $Z(s) \rightarrow Z(g)$	+30

Use the table above to answer questions 27 – 30.

27. Which reactions in the table above are exothermic?
    a. 1 and 4 only
    b. all of the equations
    c. 2 and 3 only
    d. none of the equations

28. Use the table to determine the value of ΔH for the reaction W (g) → W (s)
    a. +60 kJ
    b. +30 kJ
    c. -30 kJ
    d. -60 kJ

29. What is the value of ΔH for the reaction W (s) + X (g) → Y (g)?
    a. +20
    b. -20 kJ
    c. +100 kJ
    d. -100 k

30. Use the table to determine the value of ΔH for the reaction 2 Z (g) → 4 W (s) + 4 X (g)
    a. -340 kJ
    b. +340 kJ
    c. -100 kJ
    d. +100kJ

31. The heat of fusion of ice is 334 Joules per gram. Adding 334 Joules to one gram of ice at STP will cause the ice to
    a. Increase in temperature
    b. Decrease in temperature
    c. Change to water at a higher temperature
    d. Change to water at the same temperature

32. A solid material X is placed in liquid Y. Heat will flow from Y to X when the temperature of
    a. Y is 20°C and X is 30°C
    b. Y is 15°C and X is 10°C
    c. Y is 10°C and X is 20°C
    d. Y is 30°C and X is 40°C

Short Answer:

1. Draw and label an exothermic potential energy diagram. Label axes, E_A, TS, reactants, products, and ΔH.

2. Why do almost all chemical reactions involve changes in energy?

3. How much heat will be released if 30.0g of octane (C_8H_{18}) is burned in excess oxygen?

   $C_8H_{18} + 12\ \frac{1}{2}\ O_2 \rightarrow 8\ CO_2 + 9\ H_2O$ $\qquad$ $\Delta H = -53483.4$ kJ

4. If 1335 kJ of heat is released, what mass of H2O will be formed?

   $CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (l) + 890$ kJ

5. Given the equation: $H_2O(g) + Cu(s) + 86.6\text{ kJ} \rightarrow H_2(g) + CuO(s)$, state whether the reaction is endothermic or exothermic and whether the reactants or the products contain more stored heat energy. Then find the quantity of heat transferred when 0.750 gram of copper reacts completely.

6. Liquid water in a container is heated from 35.5 °C to 75.8 °C. Calculate the amount of heat absorbed by the water, given 357 g of the water.

7. What is the specific heat of carbon tetrachloride (CCl_4) if it requires 77 J of heat to raise the temperature of a 6.5 g sample from 21°C to 35°C?

8. If 6.00 grams of solid aluminum at 1°C absorbs 8.6 joules of heat, the temperature of the aluminum will change by how many degrees? (specific heat of Al = 0.897 J/g°C)

9. A piece of iron with a mass of 35g is heated and then placed in a beaker containing 600.g of water. The temperature of the water increases from 25°C to 33°C. The specific heat of iron is 0.451J/g °C. What was the original temperature of the piece of iron?

10. Calculate ΔH for the reaction $CH_4\ (g) + NH_3\ (g) \rightarrow HCN\ (g) + 3\ H_2\ (g)$ given the following:

$N_2\ (g) + 3H_2\ (g) \rightarrow 2\ NH_3\ (g)$	ΔH= -91.8 kJ
$C\ (s) + 2H_2\ (g) \rightarrow CH_4\ (g)$	ΔH= -74.9 kJ
$H_2\ (g) + 2C\ (s) + N_2\ (g) \rightarrow 2HCN$	ΔH= +270.3 kJ

11. Calculate ΔH for the reaction $2\ Al\ (s) + 3\ Cl_2\ (g) \rightarrow 2\ AlCl_3\ (s)$ from the data:

$2\ Al\ (s) + 6\ HCl\ (aq) \rightarrow 2$ AlCl3 $(aq) + 3\ H_2\ (g)$ $\Delta H = -1049$ kJ
$HCl\ (g) \rightarrow HCl\ (aq)$ $\Delta H = -74.8$ kJ
$H_2\ (g) + Cl_2\ (g) \rightarrow 2\ HCl\ (g)$ $\Delta H = -1845$ kJ
$AlCl_3\ (s) \rightarrow AlCl_3\ (aq)$ $\Delta H = -323$ kJ

12. Given the following heats of formation, solve the enthalpy of the reactions below.

Compound	**ΔH_f (kJ/mol)**	**Compound**	**ΔH_f (kJ/mol)**
$CH_4(g)$	-74.8	$HCl(g)$	-92.3
$CO_2(g)$	-393.5	$H_2O(g)$	-241.8
$NaCl(s)$	-411.0	$SO_2(g)$	-296.1
$H_2O(l)$	-285.8	$NH_4Cl(s)$	-315.4
$CO(g)$	-110.5	$NO_2(g)$	+33.9
$NH_3(g)$	-46.2	$ZnS(s)$	-202.9
$NaOH$ (s)	-426.7		

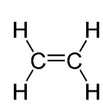
a) $NaOH(s) + HCl(g) \rightarrow NaCl(s) + H_2O(g)$

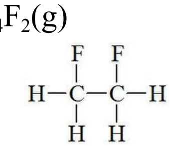
b) $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$

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

13. Using the table of bond enthalpy values from lecture slides, determine the ΔH of the following reactions.

a. $C_2H_4(g)\ +\ F_2(g)\ \rightarrow C_2H_4F_2(g)$





b. Combustion of C_5H_{12}

14. Using table values in Appendix A of your textbook (copy online) calculate the change in **enthalpy** for the following reaction:

$$C_2H_{4(g)} + 3\ O_{2(g)} \quad \square \quad 2\ CO_{2(g)} \quad + \quad 2\ H_2O_{(g)}$$

15. Using table values in Appendix A of your textbook (copy online) calculate the change in **entropy** for the following reaction:

$$C_2H_{4(g)} + 3\ O_{2(g)} \quad \square \quad 2\ CO_{2(g)} \quad + \quad 2\ H_2O_{(g)}$$

16. Using your answers from #14 and #15, calculate Gibbs energy for the reaction, if it occurs at 200˚C.

$$C_2H_{4(g)} + 3\ O_{2(g)} \quad \square \quad 2\ CO_{2(g)} \quad + \quad 2\ H_2O_{(g)}$$

a. $\Delta G = ?$

b. Is this reaction spontaneous?

c. At what temperature does this reaction become spontaneous?