After studying chapter 10 you should be able to:
- State the general properties of energy.
- Define and give examples of state functions.
- Compare and contrast temperature and heat.
- Compare and contrast endothermic and exothermic processes.
- Describe in words and diagrams the heat changes that occur in melting, freezing, boiling, and condensing, and calculate the heat changes involved.
- Calculate the heat changes that occur in chemical and physical processes
- Apply Hess' Law of heat summation to find heat changes for chemical and physical properties.
- State two reasons why reactions occur.
- State the reasons that enthalpy changes occur in chemical reactions.
- Calculate enthalpies of formation and use them to calculate enthalpies of reaction.
- Explain how the quality of energy changes as it is used.
- Describe and give examples of changes in entropy.
- Calculate Gibb's free energies, and relate Gibb's free energy to the spontaneity of reactions.

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1. State the first and second laws of thermodynamics.

2. Does condensation absorb or release energy? Explain what happens to the particles during condensation.

3. Why do phase changes occur with no change in temperature?

4. Calculate the energy needed to change 9.00 g of ice at –20.0 ^{o}C to steam at 100.0^{o}C.   Include a sketch.

Use the following constants in your problems:
C_p(ice) = 2.077 J /g ^{o}C
C_p(water) = 4.180 J /g ^{o}C
C_p(water vapor) = 2.042 J /g ^{o}C
ΔH_{fus} = 6.00 kJ/mol= 333 J/g
ΔH_{vap} = 40.6 kJ/mol = 2254 J/g

5. State the primary reason that enthalpy changes may occur during chemical reactions.

6. Calculate the enthalpy change for the reaction $A_2 + B \rightarrow A_2B$
   Given the following information: $AB + A \rightarrow A_2 + B$ $\Delta H^{\circ} = + 27.1$ kJ/mol
                                      $A_2B \rightarrow A + AB$ $\Delta H^{\circ} = -30.4$ kJ/mol

   a) Is the total reaction endothermic or exothermic? How can you tell?
   b) What is the value of ΔH° for the reverse reaction $A_2B \rightarrow A_2 + B$?

7. Calculate the heat of reaction for the formation of 2 mol of solid carbon from the decomposition of carbon dioxide. Use the following data.
   $C(s) + CO_2(g) \rightarrow 2 CO(g)$ $\Delta H = +173$ kJ
   $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ $\Delta H = -567$ kJ

8. With one exception, the standard heats of formation of all the following substances are identical: $Na(s)$, $O_2(g)$, $Br_2(l)$, $CO(g)$, $Fe(s)$, $He(g)$. What is the standard heat of formation for all but one? What is the exception? Explain the difference.

9. Use enthalpy of formation data to find the enthalpy change for the reaction: (Refer to http://fizyczna.chem.pg.gda.pl/files/2012/10/chf_epm_cr_00.pdf or your lilac handout)
 $C_2H_4(g) + 3 O_2(g) \rightarrow 2 CO_2(g) + 2 H_2O(g)$
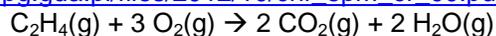
10. What is meant by entropy change? What is the symbol?

11. Give examples of situations which result in a change in entropy, and explain them.

12. Why is the "quality" of energy decreasing in the universe?

13. What is the entropy change for the single displacement reaction between NaCl and F_2? The entropy of sodium fluoride is 51.5 J/mol K—refer to (Refer to http://fizyczna.chem.pg.gda.pl/files/2012/10/chf_epm_cr_00.pdf or your lilac handout)

14. A large box is divided into two compartments with a door between them. Equal quantities of two different monatomic gases are placed in the two compartments, as shown in the overhead view in a. The door between the compartments is opened and the gas particles immediately start to mix as shown in b. Why would it be highly unlikely for the situation in b to progress to the situation shown in c?

A   B   C
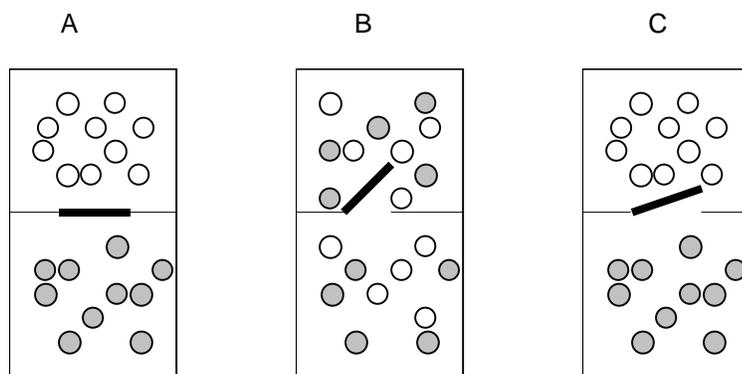
15. How is Gibb's Free Energy calculated?

16. How can you predict whether a reaction will proceed spontaneously?
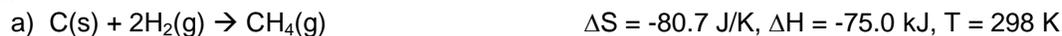
17. Under what conditions might you expect an endothermic reaction to proceed spontaneously?

18. Why are "matter spread" and "energy spread" considered to be driving forces?

19. Use http://fizyczna.chem.pg.gda.pl/files/2012/10/chf_epm_cr_00.pdf to calculate ΔH_f°, ΔS°, and ΔG_f° for the following reaction:

$$2 \ H_2S(g) + 3O_2(g) \rightarrow 2 \ H_2O(g) + 2 \ SO_2(g)$$
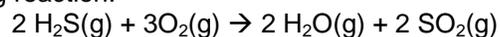
20. Calculate the free energy changes for each of the following reactions. Determine if each reaction will be spontaneous.

a) $C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\qquad\qquad$ $\Delta S = -80.7$ J/K, $\Delta H = -75.0$ kJ, $T = 298$ K

b) $3 Fe_2O_3(s) \rightarrow 2 Fe_3O_4(s) + \frac{1}{2} O_2(g)$ $\qquad$ $\Delta S = 134.2$ J/K, $\Delta H = 235.8$ kJ, $T = 298$ K

c) For reaction b, find the minimum temperature above which the reaction will become spontaneous. Calculate ΔG for this reaction if 1 mol of oxygen gas is formed at 298 K.

21. Use the following terms to fill in the blanks provided. Each word may be used once, more than once, or not at all.

Decrease	Independent	Positive
Dependent	Maximum	Temperature
Equilibrium	Minimum	Zero
Increase	Negative	

The enthalpy of a system is a) _____ of the path taken in going from the initial

state to the final state. Natural processes tend to move toward states of b) _____

disorder and c) _____ energy. In an exothermic reaction, the ΔH will have a

d) _____ sign. Endothermic reactions and weak exothermic reactions often can take

place spontaneously if the e) _____ is raised. The sign of ΔS is

f) _____ when the system is becoming less disordered. When the free energy of a

system is zero, the system is at g) _____. In a spontaneous reaction, the sign of

the free energy will be h) _____.

Answers to selected questions
1. 27600 J
7. +788 kJ
9. 1323 kJ
13. -21.2 J
19. -1036 kJ, -152.9 J/K, -990.5 kJ
20. (a) -51.0 kJ; spon
b) +195.8 kJ; non
c) T = 1757 K;
+391.6 kJ per mole O_2 formed