Chemical Reactions Chapter 17 Study Guide (Unit 10)

Name:_____ Hr:_____

- Understand and be able to explain all of the key concepts.
- Define and understand all of the survival words
- Memorize the names and symbols for these elements: (Ag, Al, Ar, As, Au, B, Ba, Be, Br, C, Ca, Cd, Cs, Cl, Co Cr, Cu, F, Fe, Fr, H, He, Hg, I, K, Kr, Li, Mg, Mn, N, Na, Ne, Ni, O, P, Pb, Rn, S, Sc, Si, Sn, Sr, Ti, U, Xe, Zn)
- Know the charges (oxidations numbers) of elements found in the s & p blocks.
- Know the polyatomic ions with charges that are on your polyatomic ion list.
- Know which elements are transition metals and therefore require the charge in parenthesis when being named. [ex. Sn (II) or Co (III)]
- Know the charges of Ag^{+1}, Zn^{+2}, Cd^{+2}, Al^{+3}, and Ga^{+3} (exceptions to the transition area)
- Know how to predict the products of any reaction
- Know how to balance a chemical reaction
- Know how to draw a Lewis structure (have-want-need-bond)
- Review all classwork (including your lab notebook) and quizzes.

Key Concepts

Lewis Structures

Be able to draw basic Lewis Structures for determining which types and how many bonds are broken and formed during a chemical reaction.
- The HAVE / WANT / NEED /BONDS calculation:
  - Have: All of the valence electrons contained within the molecule
  - Want: All of the maximum allowable valence electrons within the molecule
  - Need: Want – Have
  - Bonds: Need/2

**Helpful Hint: When doing the H/W/N/B calculation for polyatomic ions. The HAVE should account for any additional or less electrons that the molecule has. The WANT will not change.

17.1 The Flow of Energy – Heat and Work
- Heat always flows from a warmer object to a cooler object. Remember that hot objects' particles have a higher average kinetic energy compared to cooler objects which have a lower average kinetic energy.
- A systems gains heat in an endothermic process and loses heat in an exothermic process
- Heat flow is measure with two common units, the calorie or the joule. The symbol for heat is q.
- The heat capacity of an object depends on both how much mass it has and its chemical composition. This is different from specific heat capacity (c) in that specific heat capacity is only measures or specifically measures 1 gram of a substance. A chemist can compare and contrast the specific heat capacities of two different substances to determine which one requires more or less heat to change the temperature by one degree.

17.2 Measuring and Expressing Enthalpy Changes
- In calorimetry, the heat released by a system equals the heat absorbed by its surrounding. Conversely, the heat absorbed by a system equals the heat released by its surroundings. $q_{system} = - q_{surroundings}$.
- The enthalpy change for a reaction can be treated like any other reactant or product. When the enthalpy change is written in the reactants it represents an endothermic reaction. When the enthalpy change is written in the products it represents an exothermic reaction.
- The enthalpy change in a reaction (ΔH) can be calculated by adding the energy needed to break the bonds in the reactants with the energy released when new bonds form in the products (remember this is negative because it is energy being released.)

17.3 Heat in Changes of State
- The heat absorbed by a melting (fusing) solid is exactly the same as the heat lost when the liquid solidifies; that is, $\Delta H_{fusion} = - \Delta H_{solidification}$.
- The heat absorbed by a vaporizing liquid is exactly the same as the heat lost when the vapor condenses; that is, $\Delta H_{vaporization} = - \Delta H_{condensation.}$
- Heat is either released or absorbed during the dissolving process. $\Delta H_{solution}$
- The ΔH in a thermochemical equation can be used as a conversion factor.

17.4 Calculating Heats of Reaction
- You can calculate the heat of a reaction ~~by applying Hess's law of heat summation or~~ by using stand heats of formation.

Survival Words
- Calorimeter (511)
- Calorimetry (511)
- Chemical potential energy (505)
- Endothermic process (506)
- Enthalpy (511)
- Exothermic process (506)
- Heat (505)
- Heat capacity (517)
- Heat of combustion (517)
- Heat of reaction (514)
- Law of conservation of energy (506)
- Specific heat capacity (508)
- Standard heat of formation (530)
- Surroundings (506)
- System (506)
- Thermochemical equation (514)
- Thermochemistry (505)

Key Equations

$$Q = m \, C \, \Delta T$$

$$Percent\,Yield = \left(\frac{actual\ yield}{theoretical\ yield} \right) x100$$

$$\Delta H^0 = \Delta H_f^0 (\text{products}) - \Delta H_f^0 (\text{reactants})$$

Review Questions

1. If a steak is removed from the freezer and then submerged in hot water, then the
   a. temperature of the water will increase
   b. water will begin to freeze
   c. temperature of the steak will decrease and the water's temperature will increase
   d. temperature of the water will decrease and the steak's temperature will increase

2. When a chemist holds a flask that has an exothermic reaction in it, the substance inside the flask will _____ and the chemist's had will feel _____.
   a. Lose heat; hot
   b. Gain heat; hot
   c. Lose heat; cold
   d. Gain heat; cold

3. Energy (heat) lost by the surroundings is always
   a. equal to the heat gained by the system
   b. equal to the heat lost by the system
   c. greater than the heat gained by the system
   d. cannot be determined

4. What is the amount of heat required to raise the temperature of 500.0 g of aluminum by 20° C if the specific heat capacity of aluminum is $0.2 \frac{cal}{g\,°C}$?

q = mcΔT

$$q = (500.0 \; g) \left(0.2 \; \frac{cal}{g°C} \right) (20°C) = 2{,}000 \; cal$$

5. What is the specific heat for Substance Z if 100 grams of Z absorbed 500 calories of heat, resulting in a temperature increase of 30°C?

q = mcΔT

$$500 \; cal = (100 \; g) \left(c \; \frac{cal}{g°C} \right) (30°C)$$

After rearranging the equation to solve for specific heat capacity (c), $c = 0.17 \frac{cal}{g\,°C}$

6. How much energy is released with 24 g of Na_2O reacts in the following reaction? (Esper's students: make a BCA table)

$$Na_2O \text{ (s)} + 2 \text{ HI (g)} \rightarrow 2 \text{ NaI (s)} + H_2O \text{ (l)} \qquad \Delta H = -120 \text{ Cal}$$
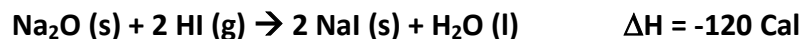
	Na_2O	+ 2 HI	$\rightarrow$ 2 NaI	+ H_2O	+ 120 cal
B	0.39				0
C	-0.39				+46.45
A	0				46.45 cal

Work:
24 g Na_2O / 62 g [molar mass] = 0.39 mol

$$\frac{1 \; mol \; Na2O}{120 \; cal} = \frac{0.39 \; mol}{x \; cal}$$

x = 46.45 cal

7. During a phase change
    a. The temperature of a substance changes
    b. The space between the particles changes This is a change in the phase energy "account"
    c. The speed of the particles remains the same
    d. The energy in the particles remains the same

8. [T/F] The heat of solution can be positive, negative, or zero. According to 17.3, heat of solution is positive or negative.

9. When 50 grams of solid sodium hydroxide are dissolved in water, how much heat is released? ($\Delta H_{solution}$ = -44.51 kJ/mol) ← per mole unit is a conversion factor
10.
    50 g NaOH / 40 g [molar mass] = 1.25 mol NaOH

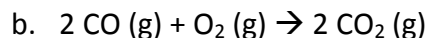$$\frac{1 \; mol \; NaOH}{44.51 \; kJ} = \frac{1.25 \; mol}{x \; kJ}$$

x = 55.64 kJ

11. Using a table that lists the standard heats of formation, you can calculate the change in enthalpy for a given chemical reaction. The change in enthalpy is equal to
    a. ΔH_f of products minus ΔH_f of reactants
    b. ΔH_f of products plus ΔH_f of reactants
    c. ΔH_f of reactants minus ΔH_f of products
    d. ΔH_f of reactants plus ΔH_f of products

    Esper's class will not be tested on this question

12. Calculate the ΔH using the standard heats of formation (given in the chart to the right) for the following reactions:

a. $2 H_2O_2 (l) \rightarrow 2 H_2O (l) + O_2 (g)$
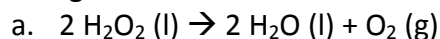
[Products] – [Reactants]
[2(-285.8) + 1(0)] – [2(-187.8)] = **- 196 kJ**

Substance	ΔH$_f$ (kJ/mol)
H_2O (l)	-285.8
H_2O_2 (l)	-187.8
O_2 (g)	0
CO (g)	-110.5
CO_2 (g)	-393.5

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

[Products] – [Reactants]
[2(-393.5)] – [2(-110.5) + 1(0)] = **- 566 kJ**

13. Using the average bond energies table below, calculate the enthalpy change in the following reactions. (Hint: Draw the Lewis Structures first!)
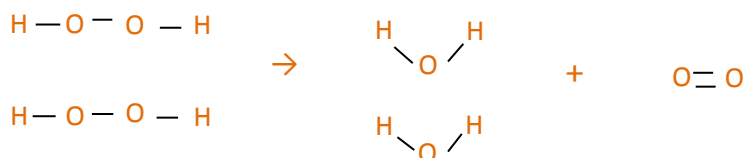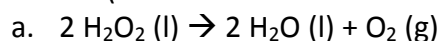
a. $2 H_2O_2 (l) \rightarrow 2 H_2O (l) + O_2 (g)$

H — O — O — H

→   H   H
        \O/   +   O꞊O

H — O — O — H   H   H
                        \O/

Breaking bonds requires energy (+) and forming bonds releases energy (-)

[Bonds breaking] – [bonds making]
[4 O-H + 2 O-O] – [4 O-H + 1 O=O]
[4(463) + 2(146)] – [4(463) + 1(495)]
2144 kJ – 2347 kJ
- 203 kJ

b. $2 CO (g) + O_2 (g) \rightarrow 2 CO_2 (g)$ (sorry for the imperfect Lewis Structures!)

C ☰ O   O ꞊ O   O ꞊ C ꞊ O

c ☰ o   →   O ꞊ C ꞊ O

Breaking bonds requires energy (+) and forming bonds releases energy (-)

[Bonds breaking] – [bonds making]
[2 C≡O + 1 O=O] – [4 C=O]
[2(1072) + 1(495)] – [4(799)]
2639 – 3196
- 557 kJ

Average Bond Energies (kJ/mol)

H–H	436 kJ/mol	C–H	413 kJ/mol	C=C	614 kJ/mol
H–Cl	431 kJ/mol	C–C	348 kJ/mol	C=C	839 kJ/mol
H–F	567 kJ/mol	C–N	293 kJ/mol	C=O	799 kJ/mol
N–H	391 kJ/mol	C–O	358 kJ/mol	O=O	495 kJ/mol
N–O	201 kJ/mol	C–F	485 kJ/mol	C≡O	1072 kJ/mol
O–H	463 kJ/mol	C–Cl	328 kJ/mol	C=N	615 kJ/mol
O–O	146 kJ/mol	C–S	259 kJ/mol	N=N	418 kJ/mol
F–F	155 kJ/mol	Cl–Cl	242 kJ/mol	N≡N	941 kJ/mol

14. According to the chart to the right, the material requiring the *most* energy to raise its temperature from 10°C to 50°C is _____water_____?

This means its resistance to change temperature is [high / low] compared to the other substances on the table.

Material	Specific heat (J/kgºC)
water	4,184
aluminum	900
steel	470
silver	235
oil	1,900
concrete	880
glass	800
gold	129
wood	2,500

15. Consider the following heating curve:
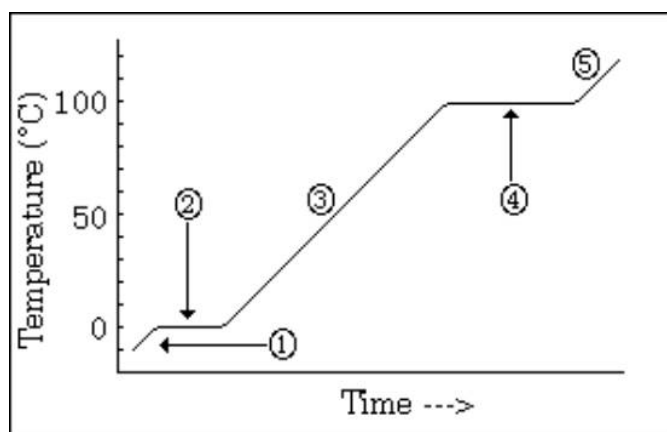
a. Which section represents the solid? ____1____
b. Which section represents the gas? ___5____
c. Which section represents the liquid? ___3____
d. Which section represents the melting/freezing point? ___2____
e. Which section represents the boiling/condensing point? ___4____
f. In which sections is the temperature stays the same? ___2, 4____
g. In which sections is the temperature changing? __1, 3, 5____
h. In which sections is the equation q=mcΔT used to calculate energy? ___1, 3, 5____
i. In which sections is a conversion factor used to calculate energy? __2, 4____

16. How many calories are required to raise the temperature of a 50 gram sample of iron from 25°C to 45°C if iron has a specific heat capacity of 0.11 $\frac{cal}{g\,°C}$?

q = mcΔT

$$q = (50.0\ g\,)\left(0.11\ \frac{cal}{g\,°C}\right)(20°C) = 110\ cal$$

17. How many calories are required to raise the temperature of 30 grams of ice from a temperature of -20°C to 150°C? (Use the heating curve in question 14 as a reference.)

$C_{H20\ (g)}$ =0.48 $\frac{cal}{g\,°C}$; $C_{H20\ (s)}$ =0.50 $\frac{cal}{g\,°C}$; $C_{H20\ (l)}$ =1.0 $\frac{cal}{g\,°C}$; $\Delta H_{vaporization}$ = 540 $\frac{cal}{g}$; ΔH_{fusion} = 800 $\frac{cal}{g}$

Draw a picture of the heating curve (or use the one above) to plot the 50 gram sample's journey from -20°C to 150°C.

Leg 1:

q = mcΔT

$$q = (30.0 \ g) \left(0.50 \ \frac{cal}{g \ ^\circ C}\right)(20 ^\circ C) = 300 \ cal$$

Leg 2:

ΔH$_{vap}$ and ΔH$_{fus}$ values are given for every gram in this problem.

$$\frac{1 \ g \ H20}{800 \ cal} = \frac{30 \ g \ H2O}{x \ cal}$$

x = 24000 cal

Leg 3:

q = mcΔT

$$q = (30.0 \ g) \left(1.0 \ \frac{cal}{g \ ^\circ C}\right)(100 ^\circ C) = 3000 \ cal$$

Leg 4:

ΔH$_{vap}$ and ΔH$_{fus}$ values are given for every gram in this problem.

$$\frac{1 \ g \ H20}{540 \ cal} = \frac{30 \ g \ H2O}{x \ cal}$$

x = 16200 cal

Leg 5:

q = mcΔT

$$q = (30.0 \ g) \left(0.48 \ \frac{cal}{g \ ^\circ C}\right)(50 ^\circ C) = 720 \ cal$$

Σ Legs = 22,620 cal

18. It takes 200 calories of heat to raise the temperature of 14 grams of canola oil by 30°C.
    a. Calculate the specific heat capacity of canola oil.

    q = mcΔT

    $$200\ cal = 14\ g\ \left(?\ \frac{cal}{g\ ℃}\right)30℃$$

    After rearranging the equation to solve for specific heat capacity (c), c =0.48 $\frac{cal}{g\ ℃}$

    b. How does the specific heat capacity of canola oil compare to that of water (is it higher, lower, equal)?

    Liquid water's specific heat capacity is 1.0 $\frac{cal}{g\ ℃}$, so canola oil's specific heat capacity of 0.48 $\frac{cal}{g\ ℃}$ is much lower than that of water. (Canola oil's specific heat capacity is the same, however, as water *vapor's* specific heat capacity.)

    c. If you have 100 g of canola oil and 100 g of water and they are heated by the same source (both are absorbing equal amounts of heat), which substance will reveal the greatest change in temperature during 10 minutes of heating? Explain your reasoning?

    If you are heating liquid water and canola oil for 10 minutes, the canola oil will have the greatest change in temperature, because it only requires 0.48 calories for every gram to change the temperature by 1 degree verses water's 1.0 calorie.

19. A 5 gram sample of Ben and Jerry's Chubby Hubby Ice Cream is placed in a bomb calorimeter with 500.00 grams of water at an initial temperature of 25°C. After combustion of the ice cream the water has a temperature of 52°C.
    a. How much energy was absorbed by the water?

    $$q = mcΔT$$

    $$q = 500.0\ g\ \left(1.00\ \frac{cal}{g\ ℃}\right)27℃ = 13,500\ cal$$

    b. How much energy was released by the ice cream?

The Law of Conservation of Energy states that all energy being released into the surroundings (the water above) is equal to the energy released from the system (the ice cream), so 13,500 cal were released by the ice cream.

    c. What is the caloric value in calories/gram?
       13500 cal / 5 g = 2700 cal/g

20. A chunk of ice whose temperature is -40°C is added to an insulated cup filled with water at 0°C.
   a. In what direction does heat flow? From the water to the ice, because the ice is at a higher temperature
   b. Which of the following statements is true? Explain why!!
      i. The ice melts until it reaches the temperature of the water.
      ii. The liquid water freezes so the chunk of ice gets larger.
         Energy moves from higher temperature areas to lower temperature areas until there is no difference in temperature. This means that energy from the water will heat up the ice until the water and the ice have the same temperature. In this case, the water is at 0°C, so after losing more energy, the water would be solid. The final temperature will be between 0 and -40°C, in which water would be solid.

21. After absorbing 180 J of energy as heat, the temperature of a 35 g block of copper was 65°C. What was the initial temperature of the copper block?

$$C_{H20\,(l)} = 4.184 \frac{Joules}{g\,°C}; \quad C_{Cu\,(s)} = 0.385 \frac{Joules}{g\,°C}; \quad \Delta H_{fusion} = 333 \frac{Joules}{g}.$$

$$q = mc\Delta T$$

$$180\ joules = 35.0\ g\left(0.385\ \frac{joules}{g\,°C}\right)(65°C - T_i)$$

$$\frac{180\ joules}{13.475\ \frac{joules}{°C}} = \frac{13.475\ \frac{joules}{°C}(65°C - T_i)}{13.475\ \frac{joules}{°C}}$$

$$13.36\,°C = (65°C - T_i)$$

$$13.36\,°C = (65°C - T_i)$$

$$-51.64\ °C = -T_i$$
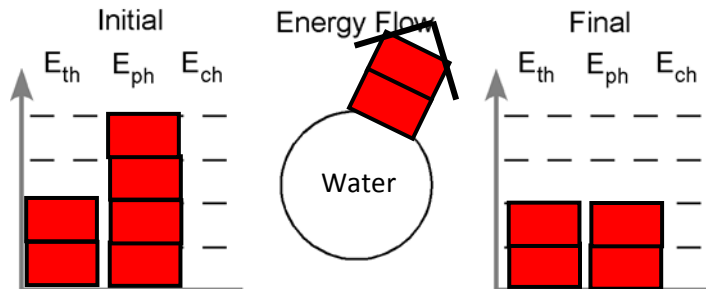
$$T_i = 51.64\ °C$$

22. Make an "energy bar chart" for the following scenarios. Then, indicate if the process is endothermic or exothermic.
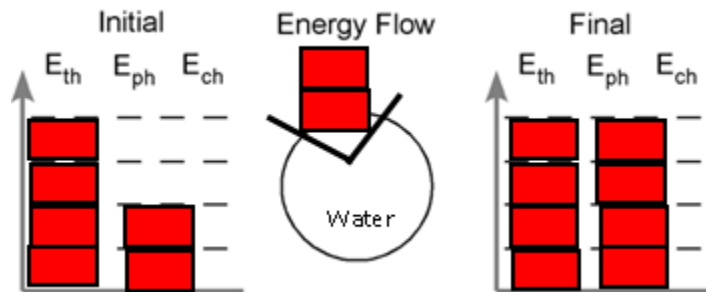
a. An ice cube melts in a glass of iced tea. [**Endothermic** / **Exothermic**]
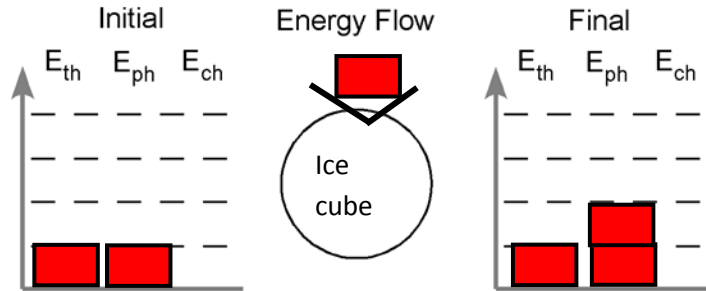
b. Water vapor in the air condenses onto a car windshield. [**Endothermic** / **Exothermic**]

c. Water at $100^{\circ}C$ is heated on a stove. [**Endothermic** / **Exothermic**]

d. A tray of water ($20^{\circ}C$) is placed in a freezer and turns into ice cubes (- $4^{\circ}C$). [**Endothermic / Exothermic**]