

Exam #1

1. Chloromethane, CH_3Cl , is used as a topical anesthetic. What quantity of heat must be absorbed by the liquid to convert 150g of liquid to a vapor at -24.09°C ?

$$\Delta H_{\text{vap}} = 21.4 \text{ kJ/mol} \quad T_{\text{bp}} = -24.09^\circ\text{C}$$

a. 0.17 kJ

b. 3210 kJ

c. 63.6 kJ

d. 7.00 kJ

e. 5.77 kJ

$$\text{CH}_3\text{Cl molar mass} = 35.45 + 12.0 + 3.0 = 50.45 \text{ g/mol}$$

$$q = \Delta H_{\text{vap}} \cdot n$$

$$= \frac{21.4 \text{ kJ}}{\text{mol}} \times 150 \text{ g} \times \frac{1 \text{ mol}}{50.45 \text{ g}} = 63.6 \text{ kJ}$$

2. 74.8J of heat is required to raise the temperature of 18.69g of silver from 10.0 to 27.0°C . What is the specific heat of silver?

a. 0.24 J/g°C

b. 4.25 J/g°C

c. 23.7 J/g°C

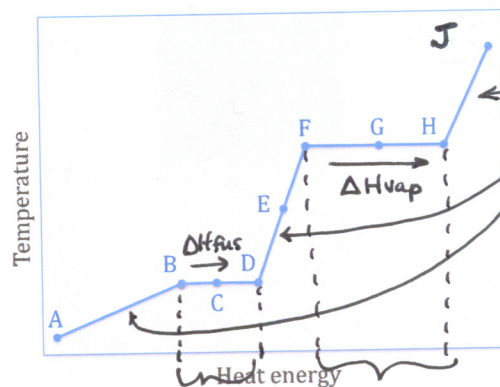
d. 82.24 J/g°C

e. 0.015 J/g°C

$$q = mc\Delta T$$

$$c = \frac{q}{m\Delta T} = \frac{74.8 \text{ J}}{(18.69 \text{ g})(27^\circ\text{C} - 10^\circ\text{C})} = 0.24 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

3. Indicate which part of the following diagram represents the region where the heat of vaporization is used to calculate energy flow.



a. line segment B→D

b. line segment D→B

c. line segment E→F

d. line segment G→H

e. line segment H→G

4. Considering the heating curve above, which is a larger value, ΔH_{fusion} or $\Delta H_{\text{vaporization}}$?

a. ΔH_{fusion}

b. $\Delta H_{\text{vaporization}}$

5. Which substance is the easiest to heat up (i.e., takes the least energy to heat)?

a. aluminum $0.9 \text{ J/g}^\circ\text{C}$

b. gold $0.129 \text{ J/g}^\circ\text{C}$

c. graphite $0.711 \text{ J/g}^\circ\text{C}$

d. diamond $0.519 \text{ J/g}^\circ\text{C}$

e. silicon $0.703 \text{ J/g}^\circ\text{C}$

← lowest heat capacity = least energy needed to heat material

6. Hydrogen and oxygen react to form liquid water in an exothermic process. If these reactants are considered the system, what change occurs to the entropy of the surroundings when the reaction occurs?

a. entropy of the surroundings is not affected by the reaction

b. entropy of the surroundings increases

c. entropy of the surroundings decreases

d. entropy change for the system must be known to answer this question

$\Delta H_{\text{rxn}} = -$, surrounding $T \uparrow$, $\Delta S_{\text{sur}} \uparrow$

7. The reaction to form a Cl_2 molecule from two Cl atoms involves forming a Cl-Cl bond.



This reaction:

a. is exothermic and heat energy is released

b. is exothermic and heat energy is absorbed

c. is endothermic and heat energy is released

d. is endothermic and heat energy is absorbed

→ Bond formation is an exothermic process (always).

8. When the reaction of two Cl atoms to form Cl-Cl occurs,



which of the following is true?

a. the entropy of the system increases

b. the entropy of the system decreases

c. the entropy change will be zero because the same atoms are present

d. you can't predict what direction of entropy change

2 molecules of gas → 1 molecule of gas } entropy decreases when # of gas molecules decrease

9. A reaction has $\Delta H = +730 \text{ kJ}$ and $\Delta S = -58 \text{ J/K}$. In what temperature range is the reaction thermodynamically favored?

a. at high temperatures

b. at low temperatures

c. the reaction is always favored

d. the reaction will never be favored

e. cannot be determined from the information given

ΔH is favorable when negative
 ΔS is favorable when positive
 In this case, both ΔH and ΔS are not favorable, so

10. It's cold out there! When I boil water to make hot chocolate at 100°C , what occurs?

- a. the kinetic energy of the water increases
- b. the entropy of the water increases
- c. the entropy of the water decreases
- d. the enthalpy of the water decreases
- e. both the entropy and the kinetic energy of water increase**
- f. the enthalpy of the water decreases and the entropy increases

11. Which of the following reactions involve a decrease in entropy?

- r1. $2 \text{Mg}(s) + \text{O}_2(g) \rightarrow 2 \text{MgO}(s)$ decrease in entropy (fewer gas molecules in product)
- r2. $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$ increase in entropy (more gas molecules in product)
- r3. $2 \text{CH}_2\text{H}_2(g) + 5 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 2 \text{H}_2\text{O}(l)$ decrease in entropy (fewer gas molecules in product)
- a. only r1
- b. only r2
- c. only r3
- d. r1 and r3**
- e. r2 and r3
- Handwritten note: 7 gas molecules \rightarrow 4 gas molecules*

12. If a reaction is thermodynamically favored, which of the following must definitely be true: (rxn and sys refer to the same thing)

- a. $\Delta S_{\text{universe}} > 0$ and $\Delta G_{\text{rxn}} < 0$**
 - b. $\Delta S_{\text{universe}} < 0$ and $\Delta G_{\text{rxn}} < 0$
 - c. $\Delta S_{\text{sys}} > 0$
 - d. $\Delta S_{\text{sys}} < 0$
 - e. $\Delta S_{\text{universe}} < 0$
- Handwritten note: The 2nd Law of Thermodynamics states that the entropy of the universe must always increase ($\Delta S_{\text{uni}} > 0$). ΔG is defined so that when $\Delta G_{\text{rxn}} < 0$, $\Delta S_{\text{uni}} > 0$.*

13. A reaction has $\Delta H = -87 \text{ kJ}$ and $\Delta S = -152 \text{ J/K}$. What is the entropy change for the universe for this reaction when performed at 298 K ?

- a. -140 J/K
- b. $+140 \text{ J/K}$**
- c. -152 J/K
- d. $+152 \text{ J/K}$
- e. $+292 \text{ J/K}$

$$\Delta S_{\text{universe}} = \Delta S_{\text{sys}} - \frac{\Delta H_{\text{rxn}}}{T}$$

$$= \frac{-152 \text{ J}}{\text{K}} - \frac{(-87,000 \text{ J})}{298 \text{ K}} = 139.9 \text{ J/K}$$

14. Consider a reaction for which $\Delta H = -126 \text{ kJ}$ and $\Delta S = -105 \text{ J/K}$. At what temperature does this reaction become spontaneous?

- a. above 1200 K
- b. below 1200 K**
- c. above 1.2 K
- d. below 1.2 K
- e. above 83 K

enthalpy favorable (ΔH^-) } $\Delta G = \Delta H - T\Delta S$
 entropy unfavorable (ΔS^+) } \downarrow
 enthalpy dominates @ low T 's,
 so the ΔG_{rxn} will be -
 below the T where $\Delta G_{\text{rxn}} = 0$

$$\Delta G = \Delta H - T\Delta S$$

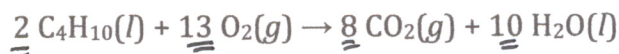
$$\frac{\Delta G - \Delta H}{\Delta S} = -T$$

$$T = \frac{(-126,000 \text{ J}) - 0}{(-105 \text{ J/K})} = 1200 \text{ K}$$

$$\frac{\Delta H - \Delta G}{\Delta S} = T$$

when $\Delta G = 0$, reaction @ equilibrium

15. Given the following enthalpies of formation, determine the enthalpy change for the combustion of butane.



Compound	ΔH_f
$\text{C}_4\text{H}_{10}(\text{l})$	-147.6 kJ
$\text{CO}_2(\text{g})$	-393.5 kJ
H_2O	-285.8 kJ

$\Delta H_f = 0$ for element in its standard state. Thus ΔH_f for $\text{O}_2 = 0$ kJ.

a. -5711 kJ

b. +5711 kJ

c. -532 kJ

d. +532 kJ

e. -2956 kJ

f. +2956 kJ

ΔH is negative (combustion is an exothermic process)

$$\Delta H = \sum \Delta H_{\text{products}} - \sum \Delta H_{\text{reactants}}$$

$$= \left[(8 \times -393.5 \text{ kJ}) + (10 \times -285.8 \text{ kJ}) \right] - \left[(2 \times -147.6 \text{ kJ}) + (13 \times 0 \text{ kJ}) \right]$$

$$= -5710.8 \text{ kJ}$$

16. What is the enthalpy change for the reaction of $\text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{CO}_2(\text{g})$

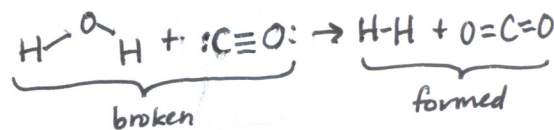
Bond energies:

O-H 463 kJ

C-O triple bond 1072 kJ

H-H 436 kJ

C=O double bond 799 kJ



a. -36 kJ

b. +36 kJ

c. -300 kJ

d. +300 kJ

$$\Delta H = \sum \text{bonds broken} - \sum \text{bonds formed}$$

$$(2 \text{H}-\text{O} + \text{C}\equiv\text{O}) - (\text{H}-\text{H} + 2 \text{C}=\text{O})$$

$$\Delta H = \left[(2 \times 463 \text{ kJ}) + (1072 \text{ kJ}) \right] - \left[(436 \text{ kJ}) + (2 \times 799 \text{ kJ}) \right] = -36 \text{ kJ}$$

17. Consider the reaction that occurs when two proteins come together and dimerize (two individual molecules bind and form one unit). For this reaction, $\Delta H = 17$ kJ and $\Delta S = 65$ J/K. What is ΔG at 298 K?

a. -23 kJ

b. -19.4 kJ

c. -36.4 kJ

d. +19.4 kJ

e. -2.37 kJ

$$\Delta G = \Delta H - T\Delta S$$

$$= 17 \text{ kJ} - \left(298 \text{ K} \times \frac{0.065 \text{ kJ}}{\text{K}} \right)$$

18. A hemoglobin molecule (molar mass = 60,000g/mole) can bind up to four oxygen molecules. In a coffee-cup calorimetry experiment (constant pressure) a 0.085L solution containing 6.0g of deoxyhemoglobin (hemoglobin with no bound oxygen) was combined with an excess of oxygen. The heat capacity of the calorimeter is negligible. What is the enthalpy of reaction per mole of oxygen bound if the temperature rose by 0.044°C. Assume the solution is dilute so that the specific heat of the solution is equal to that of water. ($C_{\text{water}} = 4.18 \text{ J/g}^\circ\text{C}$)

- a. -16.7 J/mol
- b. -167.0 kJ/mol**
- c. +167.0 kJ/mol
- d. -156.3 kJ/mol
- e. +156.3 kJ/mol
- f. -1.10 kJ/mol

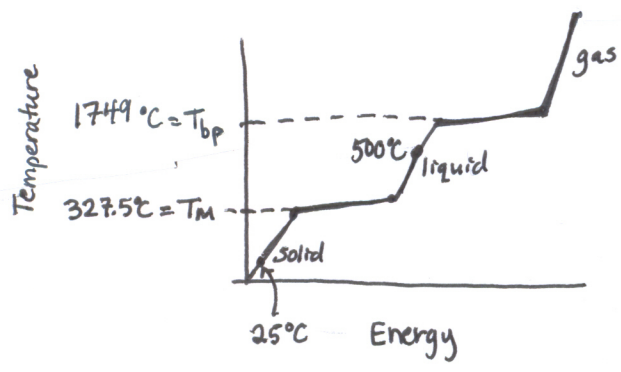
$\Delta T = 0.044^\circ\text{C}$
 exothermic reaction means ΔH will be negative.

$$\Delta H = \frac{-q}{n} = \frac{-m C \Delta T}{n} = \frac{-(6g + 85g) \left(\frac{4.18 \text{ J}}{\text{g}^\circ\text{C}} \right) (0.044^\circ\text{C})}{\left(6g \times \frac{1 \text{ mol}}{60,000g} \right)}$$

$$\Delta H = -167,367.2 \text{ J/mol} = -167 \text{ kJ/mol}$$

19. How much heat is required to raise the temperature of 97g of lead from 25°C to 500°C? Show all of your work and circle your final answer.

- $T_m = 327.5^\circ\text{C}$ $T_{bp} = 1749^\circ\text{C}$
- $C_p = 26.65 \text{ J/mol K}$ $\Delta H_{\text{fus}} = 4.77 \text{ kJ/mol}$ $\Delta H_{\text{vap}} = 179.5 \text{ kJ/mol}$



$$q_{\text{total}} = q_1 + q_2 + q_3$$

- $q_1 = \text{heat to go from } 25^\circ\text{C to } 327.5^\circ\text{C } (T_m)$
- $q_2 = \text{heat to go from solid} \rightarrow \text{liquid } (@ 327.5^\circ\text{C})$
- $q_3 = \text{heat to go from } 327.5^\circ\text{C} \rightarrow 500^\circ\text{C}$

$$n = 97g \times \frac{1 \text{ mol}}{207.2g} = 0.468 \text{ mol}$$

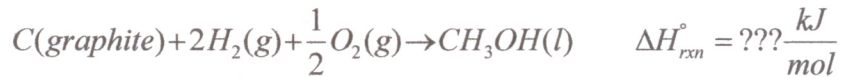
$$q_1 = m C \Delta T = (0.468 \text{ mol}) \left(\frac{26.65 \text{ J}}{\text{mol K}} \right) (327.5^\circ\text{C} - 25^\circ\text{C}) = 3772.8 \text{ J} = 3.77 \text{ kJ}$$

$$q_2 = \Delta H_{\text{fus}} \cdot n = \frac{4.77 \text{ kJ}}{\text{mol}} \times 0.468 \text{ mol} = 2.23 \text{ kJ}$$

$$q_3 = m C \Delta T = (0.468 \text{ mol}) \left(\frac{26.65 \text{ J}}{\text{mol K}} \right) (500^\circ\text{C} - 327.5^\circ\text{C}) = 2151.45 \text{ J} = 2.15 \text{ kJ}$$

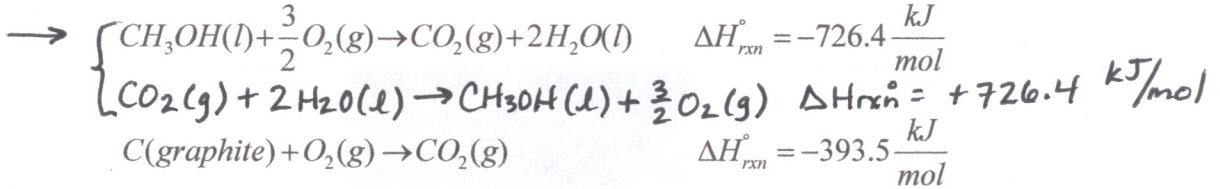
$$q_{\text{total}} = 3.77 \text{ kJ} + 2.23 \text{ kJ} + 2.15 \text{ kJ} = 8.15 \text{ kJ}$$

20. Calculate the enthalpy of formation of methanol (CH₃OH) from its elements:

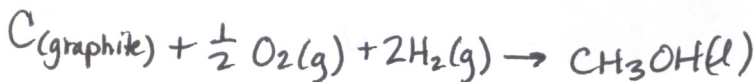
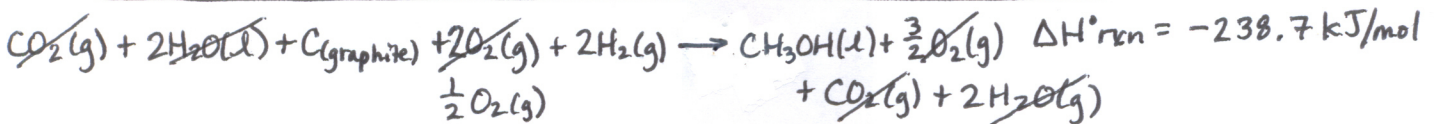
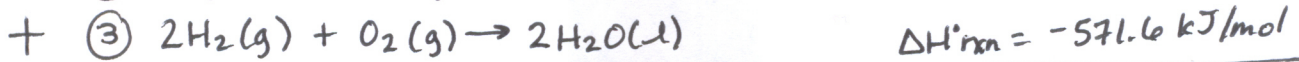
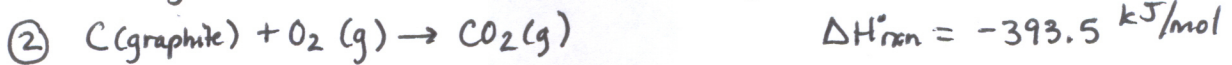
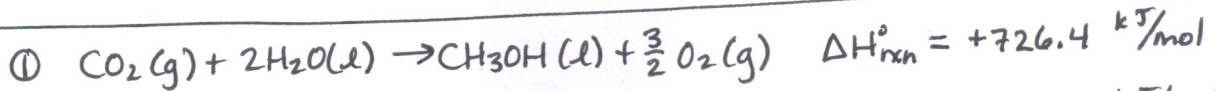
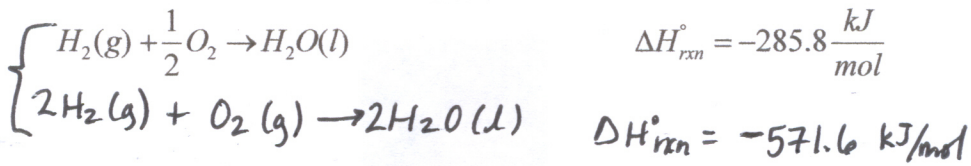


using the following information. Show all of your work and circle your final answer. Use Hess' Law

needs to be reversed because CH₃OH is the product you are interested in.



need to multiply by 2 to get the H₂O's to cancel when you add the reactions



$$\Delta H_{\text{rxn}}^\circ = -238.7 \frac{\text{kJ}}{\text{mol}}$$

Extra credit challenge question:

The fuel value of a hamburger is approximately 3.6 kcal/g. If a man eats a 1 pound hamburger for lunch and if none of the energy is stored in his body, estimate the amount of water that would have to be lost in perspiration (i.e., sweat!) to keep his body temperature constant. For water, $\Delta H_{\text{vap}} = 2.41 \text{ kJ/g}$

$$1 \text{ lb.} = 453.6 \text{ g} \quad 1 \text{ kcal} = 4.184 \text{ kJ}$$

$$\text{The amount of energy provided by a 1 pound burger} = 453.6 \text{ g} \times \frac{3.6 \text{ kcal}}{\text{g}} = 1632.96 \text{ kcal}$$

$$1632.96 \text{ kcal} \times \frac{4.184 \text{ kJ}}{\text{kcal}} = 6832.3 \text{ kJ} \leftarrow \text{This is the amount of "heat energy" you will use up.}$$

To vaporize water, it must go from 37°C (98.6°F) to 100°C (water's boiling point) and then it must change phase from liquid → gas (at 100°C).

$$\text{Total } q \text{ used} = \underbrace{(mC\Delta T)}_{\text{heating}} + \underbrace{(\Delta H_{\text{vap}} \cdot m)}_{\text{l} \rightarrow \text{g}}$$

↑
6832.3 kJ

Extra Credit Answer, continued

$$q = (m c \Delta T) + (\Delta H_{\text{vap}} \times m)$$

We know q , c , ΔH_{vap} and ΔT and we want to know m , so we solve for m :

$$q = m (c \Delta T + \Delta H_{\text{vap}})$$

$$m = \frac{q}{(c \Delta T + \Delta H_{\text{vap}})} = \frac{(6832.3 \text{ kJ})}{\left(\frac{0.00418 \text{ kJ}}{\text{g}^\circ\text{C}}\right)(100^\circ\text{C} - 37^\circ\text{C}) + \left(\frac{2.41 \text{ kJ}}{\text{g}}\right)}$$

$$m = 2555.7 \text{ g} \times \frac{1 \text{ mL}}{1 \text{ g}} = 2555.7 \text{ mL} = 2.6 \text{ L}$$

↑
density of water

2.6 L of water