AP Chemistry
Unit 5- Homework Problems
Thermodynamics

Specific Heat Problems

1. How many J would it take to raise the temperature of 200 grams of water from 5 ^oC to 85 ^oC?

Q = (200 g)(4.18)(80 ^oC) = 66880 J

2. How many J would problem number 1 be if it was aluminum instead of water?

Q = (200 g)(0.897)(80 ^oC) = 14352 J

3. How many grams of copper could be heated from 20 ^oC to 75 ^oC if 1200 J are applied to it?

1200 J = (x)(0.385)(55 ^oC) = 56.7 g

4. What is the specific heat capacity of a substance if 750 J caused 100 grams of it to go from 90 ^oC to 135 ^oC?

750 J = (100 g)(x)(45 ^oC) = 0.167 J/g^oC

5. What would the final temperature be if 500 J are applied to 150 grams of ice at –90 ^oC?

500 J = (150 g)(2.06)(x – (-90 ^oC)) =-88.4 ^oC

6. What would the temperature *change by* if a 90 gram piece of hot iron cooled by losing 200 J.

-200 J = (90 g)(0.449)(x) = -4.95 ^oC

7. What was the initial temperature if 500 J were applied to 250 g of mercury and the final temperature was 50 ^oC?

500 J = (250 g)(0.140)(50 ^oC - x) = 35.7 ^oC

Latent Heat Problems

1. Why can we not use the equation $Q = mc\Delta T$ for phase changes of a substance; why won't it work?

There is no temperature change at a phase change

2. What is the Latent Heat of Fusion of a substance?

Energy needed to melt or freeze

3. What is the Latent Heat of Vaporization of a substance?

Energy needed to boil or condense

4. How many J are needed to melt 30 grams of copper?

Q = (30 g)(209 J/g) = 6270 J

5. How much heat is needed to boil 63 grams of ethanol?

Q = (63 g)(838 J/g) = 52.8 kJ

6. What is the Latent Heat of Fusion of a substance if 1200 J melts 40 grams of it?

1200 J = (40 g)(x) = 30 J/g

7. What is the Latent Heat of Vaporization of a substance if 50 J boiled 3 grams of it?

50 J = (3 g)(x) = 16.7 J/g

8. How many grams of copper could be melted if 700 J are applied to it?

700 J = (x)(209 J/g) = 3.35 g

9. How many grams of water could be boiled if 8000 J are applied to it?

8000 J = (x)(2260 J/g) = 3.54 g

Combination Problems

1. How much energy is needed to heat 200 grams of diethyl ether, $(C_2H_5)_2O$, from 5 ^oC to its boiling point and boil it?

A) Heat to boiling point $Q = mc(T_f - T_i) = (200\ g)(2.33)(34.6 – 5\ ^oC) = 13.8\ kJ$

B) Boil $Q = mL_v = (200\ g)(357\ J/g) = 71.4\ kJ$

Total = 13.8 + 71.4 = 85.2 kJ

2. How much energy is needed to turn 400 grams of liquid benzene, C_6H_6 , at 20 ^oC to gaseous benzene at 150 ^oC?

A) Heat to boiling point $Q = mc(T_f - T_i) = (400\ g)(1.74)(80 – 20\ ^oC) = 41.8\ kJ$

B) Boil $Q = mL_v = (400\ g)(393\ J/g) = 157.2\ kJ$

C) Heat gas from bp to 150 ^oC $Q = mc(T_f - T_i) = (400\ g)(1.06)(150-80) = 29.7\ kJ$

Total = 41.8 + 157.2 + 29.7 = 228.7 kJ

3. How much energy is needed to turn a 75 g block of ice at -40 ^oC to steam at 250 ^oC?

A) Heat to melting point $Q = mc(T_f - T_i) = (75\ g)(2.06)(0 – (-40\ ^oC)) = 6.2\ kJ$

B) Melt $Q = mL_f = (75\ g)(333\ J/g) = 25\ kJ$

C) Heat liq from mp to bp $Q = mc(T_f - T_i) = (75\ g)(4.18)(100 – 0\ ^oC) = 31.4\ kJ$

B) Boil $Q = mL_v = (75\ g)(2260\ J/g) = 169.5\ kJ$

C) Heat gas from bp to 250 ^oC $Q = mc(T_f - T_i) = (75\ g)(1.86)(250 – 100\ ^oC) = 20.9\ kJ$

Total = 6.2 + 25 + 31.4 + 169.5 + 20.9 kJ = 253 kJ

4) Which part of #3 contributes the most energy to the overall process?

The boiling part (going from liquid to gas or vice versa) requires the most energy

Latent Heat Graphing

1. What do you call each of the following changes of state:

A. Solid to liquid	melt
B. Liquid to gas	boil
C. Gas to solid	deposition
D. Solid to gas	sublimation
E. Liquid to solid	freeze
F. Gas to liquid	condense

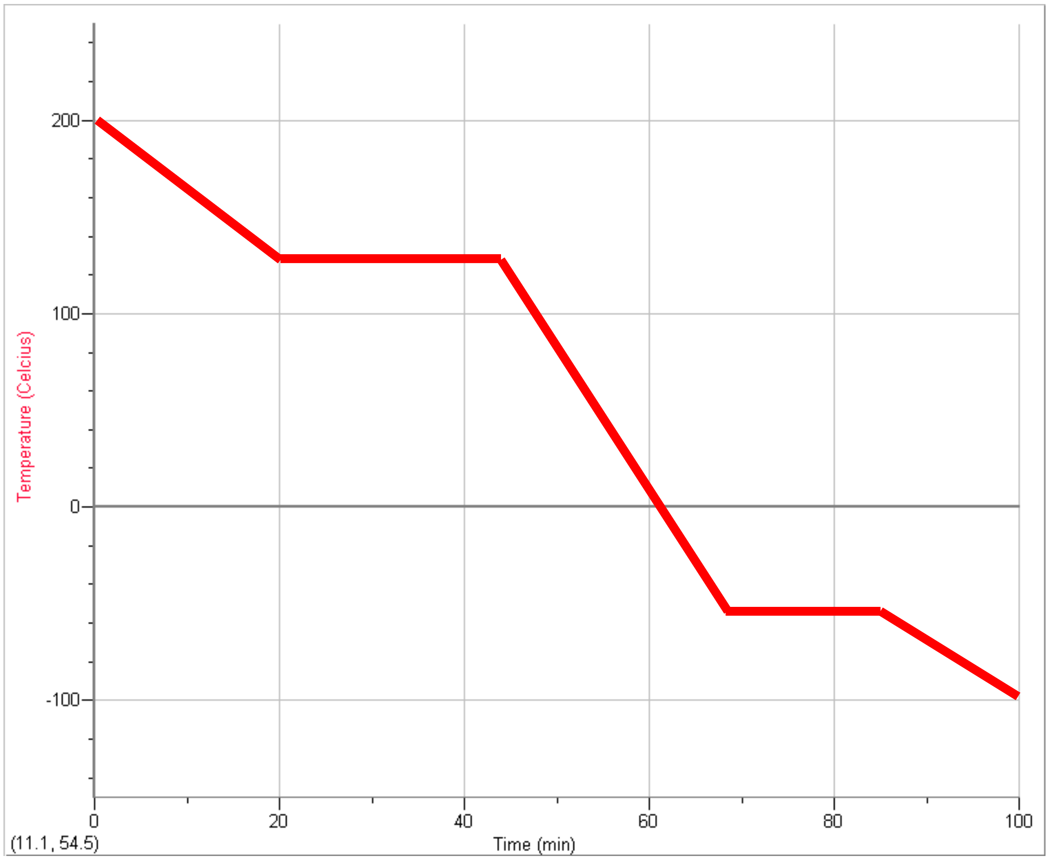
2. The melting point of a solid is the same as what other point?

Freezing point

3. The condensing point of a gas is the same as what other point?

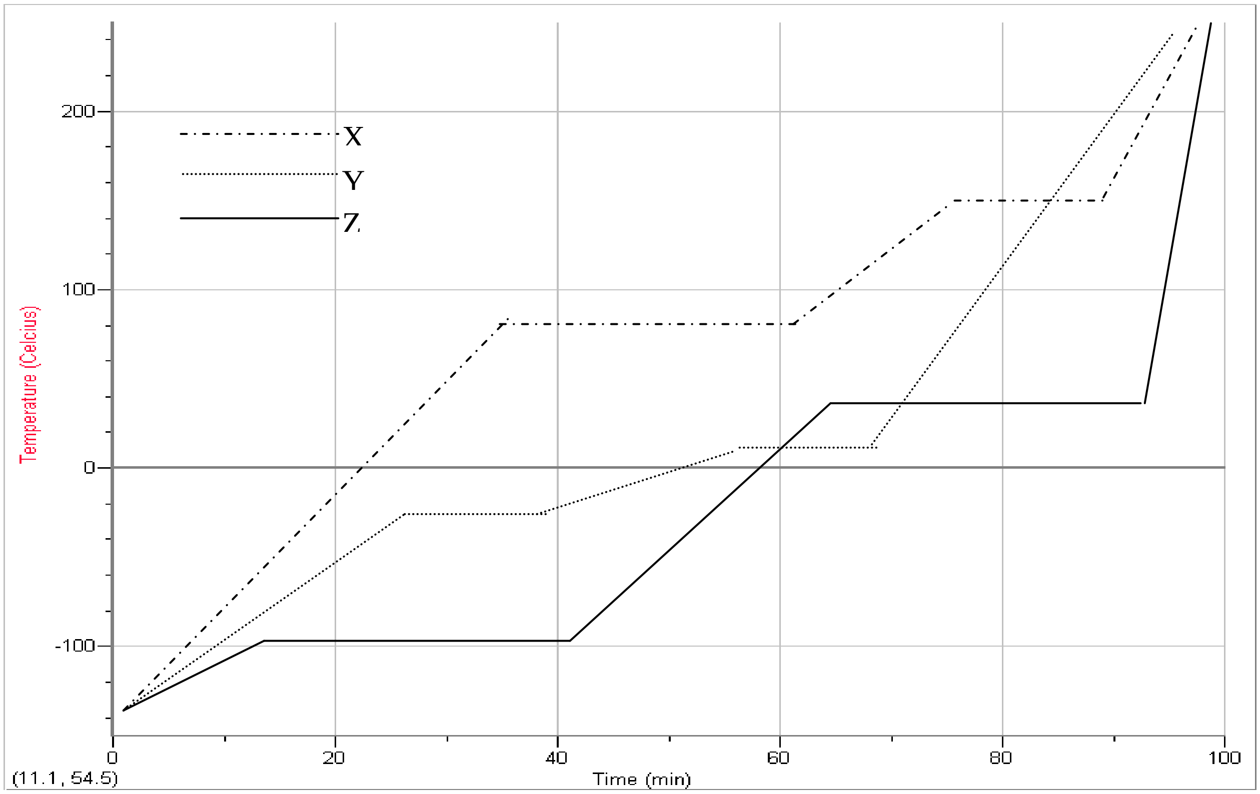
Boiling point

4. Octane, the major component in gasoline freezes at –57 °C and boils at 125 °C. If gaseous octane was cooled from 200 °C to –100 °C, draw what the graph would look like:



5. The following graph shows three different substances and their phase changes. Answer the following questions about them:

A. What is the melting point of substance Y? @ -25 ^{o}C
B. What is the boiling point of substance Z? @ 40 ^{o}C
C. Which substance is still a solid when the other two substances have turned to gases? X
D. Which substance has the lowest boiling point? Y
E. Which substance has the lowest freezing point? Z



6. Water boils at 100 °C. Why are steam burns potentially more dangerous than boiling water burns?

Under normal conditions, liquid water can only get up to 100 ^{o}C. Steam, because it is a gas, can get hotter than 100 ^{o}C and is potentially more dangerous because of it.

Calorimetry

1. What is the specific heat capacity of water? 4.18 J/g^{o}C

2. A piece of ice is placed on the sidewalk on a hot summer day. Describe what happens to energy content of the piece of ice, does it gain or lose energy? What happens to the air around the ice, does it gain or lose energy?

The ice gains energy from the surrounding air and increases its energy (endothermic).

The air loses energy to the ice and decreases its energy (exothermic)

3. 100 grams of hot water at 80 ^{o}C is combined with 100 grams of cool water at 20 ^{o}C. What is the final temperature of the combined water?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mc(T_f - T_i)$ (100 g)(4.18)(x- 80 ^{o}C) = - (100 g)(4.18)(x- 20 ^{o}C) T_f = 50 ^{o}C

4. 100 grams of hot water at 80 ^{o}C is combined with 50 grams of cool water at 20 ^{o}C. What is the final temperature of the combined water?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mc(T_f - T_i)$ (100 g)(4.18)(x- 80 ^{o}C) = - (50 g)(4.18)(x- 20 ^{o}C) T_f = 60 ^{o}C

5. A 25 gram piece of hot metal at 97 ^{o}C is plunged into a 35 gram cup of cool water at 19 ^{o}C. The metal gives up its heat to the water until they are both at 22 ^{o}C. What is the specific heat capacity of the metal?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mc(T_f - T_i)$ (25 g)(x)(22- 97 ^{o}C) = - (35 g)(4.18)(22- 19 ^{o}C) C = 0.234 J/g^{o}C

6. A 75 gram piece of hot metal at 100 ^{o}C is dropped into a 50 gram cup of cool water at 22 ^{o}C. The final temperature of the system is 30 ^{o}C. What is the specific heat capacity of the metal?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mc(T_f - T_i)$ (75 g)(x)(30- 100 ^{o}C) = - (50 g)(4.18)(30- 22 ^{o}C) C = 0.318 J/g^{o}C

7. A 120 gram piece of copper at 75 ^{o}C is put into a 20 gram sample of water at 10 ^{o}C. What is the final temperature of the system after the copper releases all its extra heat?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mc(T_f - T_i)$ (120 g)(0.385)(x- 75 ^{o}C) = - (20 g)(4.18)(x- 10 ^{o}C) T_f = 33.1 ^{o}C

8. 100 g of water at 22 ^{o}C was combined with 20 grams of ice. The final temperature of the system was 6 ^{o}C. Calculate the heat of fusion of ice based upon this data.

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mL_f$ (100 g)(4.18)(6- 22 ^{o}C) = - (20 g)(x) L_f = 334.4 J/g

9. 200 g of water at 22 ^{o}C was combined with 15 grams of ice. Knowing that the heat of fusion of ice is 333 J/g, calculate the final temperature of the system.

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mL_f$ (200 g)(4.18)(x- 22 ^{o}C) = - (15 g)(333 J/g) T_f = 16 ^{o}C

10. Suppose you heat a 50 g piece of silver to 99.8 ^{o}C and then drop it onto ice. When the metal's temperature has dropped to 0 ^{o}C, it is found that 3.54 g of ice has melted. What is the specific heat capacity of silver?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mL_f$ (50 g)(x)(0- 99.8 ^{o}C) = - (3.54 g)(333 J/g) c = 0.236 J/g^{o}C

11. A 9.36 g piece of Pt is heated to 98.6 ^{o}C and then dropped onto a block of ice. When the temperature of the metal has dropped to 0 ^{o}C, it is found that 0.37 g of ice melted. What is the specific heat capacity of Pt?

$Q_{lost} = - Q_{gain}$ $mc(T_f - T_i) = - mL_f$ (9.36 g)(x)(0- 98.6 ^{o}C) = - (0.37 g)(333 J/g) T_f = 0.134 J/g^{o}C

12. Describe each of the reactions below as either endothermic or exothermic. Some are tricky, be careful. Decide whether heat is entering or leaving the system in question:

A. Exploding fireworks	endothermic	**[exothermic]**
B. Melting snow	**[endothermic]**	exothermic
C. Lava cooling	endothermic	**[exothermic]**
D. Paper burning	endothermic	**[exothermic]**
E. An ice cube freezing	endothermic	**[exothermic]**
F. Water evaporating	**[endothermic]**	exothermic

Enthalpy Problems

1. For the equation: $Mg + \frac{1}{2} CO_2 \rightarrow MgO + \frac{1}{2} C$

   If 16.7 kJ of heat is given off per 1.0 gram of magnesium reacted, calculate the ΔH_{rxn}

   $\Delta H^o_{rxn} = Q_{rxn}$/moles ΔH^o_{rxn} = -16.7 kJ/0.0411 moles = -406 kJ/mole

2. For the equation: $(NH_4)_2Cr_2O_7 \rightarrow N_2 + 4\ H_2O + Cr_2O_3$

   If 3 grams of $(NH_4)_2Cr_2O_7$ gives off 3.57 kJ of energy, calculate the ΔH_{rxn}

   $\Delta H^o_{rxn} = Q_{rxn}$/moles ΔH^o_{rxn} = -3.57 kJ/0.0119 moles = -300 kJ/mole

3. For the equation: $Na + H_2O \rightarrow NaOH + \frac{1}{2} H_2$

   If 5 grams of sodium is placed into 100 grams of water at 20 ^oC and the final temperature of the system reaches 27 ^oC, calculate the ΔH_{rxn}

   $\Delta H^o_{rxn} = Q_{rxn}$/moles $Q_{rxn} = - Q_{water}$ = - (100 g)(4.18 J/g^oC)(7 ^oC) = -2926 J

   ΔH^o_{rxn} = -2926 J/0.217 moles = -13.5 kJ/mole

4. For the equation: $Ca + 2\ HCl \rightarrow CaCl_2 + H_2$

   If 2 grams of calcium is placed into 75 mL of 1 M HCl at 18 ^oC and the final temperature of the system reaches 23 ^oC, calculate the ΔH_{rxn}. (Assume the density of the solution is 1 g/mL and the specific heat capacity of the HCl = 4.18 J/g^oC)

   $\Delta H^o_{rxn} = Q_{rxn}$/moles $Q_{rxn} = - Q_{water}$ = - (75 g)(4.18 J/g^oC)(5 ^oC) = -1567.5 J

   ΔH^o_{rxn} = -1567.5 J/0.0499 moles = -31.4 kJ/mole

5. For the equation: $NaNO_3$ (s) $\rightarrow$ $NaNO_3$ (aq)

   If 20 grams of $NaNO_3$ were placed into 200 grams of water at 22 ^oC, and the temperature dropped to 12 ^oC, what is the ΔH_{rxn}?

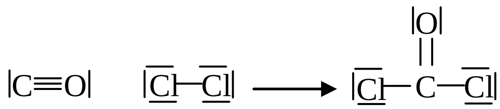
   $\Delta H^o_{rxn} = Q_{rxn}$/moles $Q_{rxn} = - Q_{water}$ = - (200 g)(4.18 J/g^oC)(-10 ^oC) = + 8360 J

   ΔH^o_{rxn} = + 8360 J /0.235 moles = +35.6 kJ/mole

Bond Energies

1. Draw the Lewis Dot structures and determine the ΔH_{rxn} for:

$$CO + Cl_2 \rightarrow Cl_2CO$$

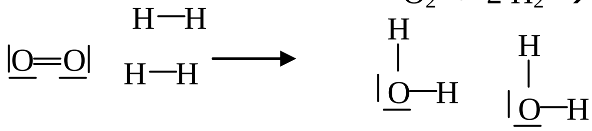


$\Delta H_{rxn} = \Sigma\ Bonds_{react} - \Sigma\ Bonds_{prod}$

$\Delta H_{rxn} = (1046 + 242) - (2*339 + 745) = -135$ kJ

2. Draw the Lewis Dot structures for and determine the ΔH_{rxn} for:

$$O_2 + 2\ H_2 \rightarrow 2\ H_2O$$

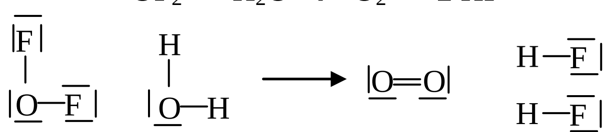


$\Delta H_{rxn} = \Sigma\ Bonds_{react} - \Sigma\ Bonds_{prod}$

$\Delta H_{rxn} = (498 + 2*436) - (4*463) = -482$ kJ

3. Draw the Lewis Dot structures for and determine the energy for the O-F bond for:

$$OF_2 + H_2O \rightarrow O_2 + 2\ HF \qquad \Delta H_{rxn} = -318\ \text{kJ}$$



$\Delta H_{rxn} = \Sigma\ Bonds_{react} - \Sigma\ Bonds_{prod}$

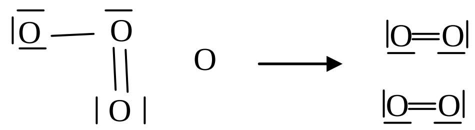
$-318 = (2X + 2*463) - (498 + 2*565)$

X = 192 kJ

4. For the reaction:

$$O_3 + O \rightarrow 2\,O_2 \qquad \Delta H_{rxn} = -394$$

a. Draw the Lewis dot structures for all species



b. Calculate the O-O bond energy in O_3

$\Delta H_{rxn} = \Sigma\ Bonds_{react} - \Sigma\ Bonds_{prod}$

$-394 = (2X) - (2*498)$

X = 301 kJ

c. What is the bond order of the O-O bond in O_3 (remember it is a resonance structure)?

Bond order = 1.5

d. Compare the answer you got in part b to the O-O bond energy and the O=O bond energy you can look up. Does your answer make sense? Why or why not?

O-O bond = 146 bond order = 1

O=O bond = 498 bond order = 2

O---O bond in O_3 bond order 1.5 has energy of 301 kJ so that makes sense as it is between the 146 and 498 or the bond orders 1 and 2

Hess' Law Problems

1. Given the following equations:

$2\,H_2 + O_2 \rightarrow 2\,H_2O$	ΔH_{rxn}= -572 kJ	$H_2O \rightarrow H_2 + \frac{1}{2}\,O_2$	ΔH_{rxn}= +286 kJ
$N_2O_5 + H_2O \rightarrow 2\,HNO_3$	ΔH_{rxn}= -74 kJ	$2\,HNO_3 \rightarrow N_2O_5 + H_2O$	ΔH_{rxn}= +74 kJ
$\frac{1}{2}\,N_2 + 3/2\,O_2 + \frac{1}{2}\,H_2 \rightarrow HNO_3$	ΔH_{rxn}= -174 kJ	$N_2 + 3\,O_2 + H_2 \rightarrow 2\,HNO_3$	H_{rxn}= 2(-174 kJ)

Calculate ΔH_{rxn} for:

$N_2 + 5/2\,O_2 \rightarrow N_2O_5$ ΔH_{rxn}= + 12 kJ

2. Given the following equations:

$N_2 + 3\,H_2 \rightarrow 2\,NH_3$	ΔH_{rxn}= -92 kJ	$\frac{1}{2}\,N_2 + 3/2\,H_2 \rightarrow NH_3$	ΔH_{rxn}= -46 kJ
$4\,NH_3 + 5\,O_2 \rightarrow 4\,NO + 6\,H_2O$	ΔH_{rxn}= -906 kJ	$NH_3 + 5/4\,O_2 \rightarrow NO + 3/2\,H_2O$	ΔH_{rxn}= -226 kJ
$H_2 + \frac{1}{2}\,O_2 \rightarrow H_2O$	ΔH_{rxn}= -242 kJ	$3/2\,H_2O \rightarrow 3/2\,H_2 + \frac{3}{4}\,O_2$	ΔH_{rxn}= +363 kJ

Calculate the ΔH_{rxn} for:

$\frac{1}{2}\,N_2 + \frac{1}{2}\,O_2 \rightarrow NO$ ΔH_{rxn}= +91 kJ

3. Given the following equations:

$Sr + \frac{1}{2}\,O_2 \rightarrow SrO$	ΔH_{rxn}= -592 kJ	$Sr + \frac{1}{2}\,O_2 \rightarrow SrO$	ΔH_{rxn}= -592 kJ
$SrO + CO_2 \rightarrow SrCO_3$	ΔH_{rxn}= -234 kJ	$SrO + CO_2 \rightarrow SrCO_3$	ΔH_{rxn}= -234 kJ
$C + O_2 \rightarrow CO_2$	ΔH_{rxn}= -394	$C + O_2 \rightarrow CO_2$	ΔH_{rxn}= -394

Calculate the ΔH_{rxn} for:

$Sr + C + 3/2\,O_2 \rightarrow SrCO_3$ ΔH_{rxn}= -1220 kJ

4. Given the following equations:

$C + 2\,H_2 \rightarrow CH_4$	ΔH_{rxn}= -75 kJ	$CH_4 \rightarrow C + 2\,H_2$	ΔH_{rxn}= +75 kJ
$C + 2\,Cl_2 \rightarrow CCl_4$	ΔH_{rxn}= -96 kJ	$C + 2\,Cl_2 \rightarrow CCl_4$	ΔH_{rxn}= -96 kJ
$H_2 + Cl_2 \rightarrow 2\,HCl$	ΔH_{rxn}= -92 kJ	$2\,H_2 + 2\,Cl_2 \rightarrow 4\,HCl$	ΔH_{rxn}= 2(-92) kJ

Calculate the ΔH_{rxn} for:

$CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl$ ΔH_{rxn}= -205 kJ

5. Given the following equations:

$C_5H_{12} + 8\,O_2 \rightarrow 5\,CO_2 + 6\,H_2O$	ΔH_{rxn}= -3506 kJ	$5\,CO_2 + 6\,H_2O \rightarrow C_5H_{12} + 8\,O_2$	ΔH_{rxn}= +3506 kJ
$C + O_2 \rightarrow CO_2$	ΔH_{rxn}= -394 kJ	$5\,C + 5\,O_2 \rightarrow 5\,CO_2$	ΔH_{rxn}= 5(-394) kJ
$2\,H_2 + O_2 \rightarrow 2\,H_2O$	ΔH_{rxn}= -484 kJ	$6\,H_2 + 3\,O_2 \rightarrow 6\,H_2O$	ΔH_{rxn}= 3(-484) kJ

Calculate the ΔH_{rxn} for:

$5\,C + 6\,H_2 \rightarrow C_5H_{12}$ ΔH_{rxn}= +84 kJ

ΔH_f^o

1. For each of the substances below, write a balanced equation showing the formation of 1 mole of the compound from its elements in their standard states. Look up the value for ΔH_f^o in the Appendix in the back of a book or online.
   a) Al_2O_3 — $2\ Al\ (s) + 3/2\ O_2\ (g) \rightarrow Al_2O_3\ (s)$ $\quad \Delta H_f^o = -1676$ kJ/mole
   b) $Mg(OH)_2$ — $Mg\ (s) + H_2\ (g) + O_2\ (g) \rightarrow Mg(OH)_2\ (s)$ $\quad \Delta H_f^o = -925$ kJ/mole
   c) $C_{12}H_{22}O_{11}$ — $12\ C\ (s) + 11\ H_2\ (g) + 11/2\ O_2\ (g) \rightarrow C_{12}H_{22}O_{11}\ (s)$ $\quad \Delta H_f^o = -2226$ kJ/mole
   d) $NaHCO_3$ — $Na\ (s) + ½\ H_2\ (g) + C\ (s) + 3/2\ O_2\ (g) \rightarrow NaHCO_3\ (s)$ $\quad \Delta H_f^o = -951$ kJ/mole

2. For each of the reactions below, calculate ΔH^o_{rxn}. Look up the values for ΔH_f^o in the Appendix in the back of a book or online.
   a) $SO_2\ (g) + ½\ O_2\ (g) \rightarrow SO_3\ (g)$ $\quad \Delta H^o_{rxn} = (-396) - (-297) = -99$ kJ
   b) $4\ NH_3\ (g) + 5\ O_2\ (g) \rightarrow 4\ NO\ (g) + 6\ H_2O\ (g)$ $\quad \Delta H^o_{rxn} = [4(91) + 6(-242)] - 4(-46) = -904$ kJ
   c) $NH_4NO_3\ (s) \rightarrow N_2O\ (g) + 2\ H_2O\ (g)$ $\quad \Delta H^o_{rxn} = [(82) + 2(-242)] - (-366) = -36$ kJ

3. Nitroglycerin is a powerful explosive that explodes by the following equation while giving off 4200 kJ of heat.
   $C_3H_5(NO_3)_3 \rightarrow 3/2\ N_2\ (g) + ¼\ O_2\ (g) + 3\ CO_2\ (g) + 5/2\ H_2O\ (g)$

   Calculate the ΔH_f^o of nitroglycerin. $-4200\text{ kJ} = [3(-394) + 5/2(-242)] - x \quad x = 2413$ kJ

4. Oxygen difluoride reacts with water vapor to produce 318 kJ of heat by the following equation:
   $OF_2\ (g) + H_2O\ (g) \rightarrow 2\ HF\ (g) + O_2\ (g)$

   Calculate the ΔH_f^o of OF_2 $\quad -318\text{ kJ} = [2(-273)] - [x + (-242)] \quad x = 14$ kj

5. Large scale H_2 can be made by the following 3 steps:
   Step 1: $SO_2\ (g) + 2\ H_2O\ (g) + Br_2\ (g) \rightarrow H_2SO_4\ (l) + 2\ HBr\ (g)$ $\quad \Delta H^o_{rxn} = [(-814) + 2(-36)] - [(-297) + 2(-242) + 31] = -136$ kJ
   Step 2: $H_2SO_4\ (l) \rightarrow H_2O\ (g) + SO_2\ (g) + ½\ O_2\ (g)$ $\quad \Delta H^o_{rxn} = [(-242) + (-297)] - [(-814)] = +275$ kJ
   Step 3: $2HBr\ (g) \rightarrow H_2\ (g) + Br_2\ (g)$ $\quad \Delta H^o_{rxn} = [31] - [2(-36)] = +103$ kJ
   Total: $H_2O\ (g) \rightarrow H_2\ (g) + ½\ O_2\ (g)$ $\quad \Delta H^o_{rxn} = -136 + 275 + 103 = +242$ kJ

   Calculate the ΔH^o_{rxn} of each step. Calculate the ΔH^o_{rxn} of the overall reaction. Is the overall reaction endo or exothermic?
   Endothermic

Bomb Calorimetry

1. You burn 0.300 g of C in an excess of O_2 in a bomb calorimeter. The temperature of the calorimeter which contains 775 grams of water increases from 25 to 27.38 ^oC. The heat capacity of the bomb is 893 J/K. What is ΔH_{rxn} per mole of carbon?

$Q_{rxn} = - [q_{water} + q_{bomb}] = - [mc\Delta T_{water} + k\Delta T_{bomb}]$

$Q_{rxn} = -[(775\ g)(4.18\ J/gC)(2.38\ ^oC) + (893\ J/K)(2.38\ K)] = -9.835\ kJ$

0.30 g C *(1 mole/12 g) = 0.025 moles C

ΔH_{rxn} = q/mole = -9.835 kJ/0.025 mole = -393 kJ/mole

2. You burn 1.50 g of benzoic acid ($C_6H_5CO_2H$) in a bomb calorimeter which has a constant of 789 J/K. The calorimeter is filled with 775 g of water and the temperature increases from 22.50 to 31.69 ^oC. What is ΔH_{rxn} per mole of benzoic acid?

$Q_{rxn} = - [q_{water} + q_{bomb}] = - [mc\Delta T_{water} + k\Delta T_{bomb}]$

$Q_{rxn} = -[(775\ g)(4.18\ J/gC)(9.19\ ^oC) + (789\ J/K)(9.19\ K)] = -37.0\ kJ$

1.50 g $C_6H_5CO_2H$ *(1 mole/122 g) = 0.0123 moles $C_6H_5CO_2H$

ΔH_{rxn} = q/mole = -37 kJ/0.0123 mole = -3008 kJ/mole

Entropy Problems

1. Without doing any calculations, determine the sign of each of the entropy changes below:
   a. $Mg (s) + Cl_2 (g) \rightarrow MgCl_2 (s)$ — ΔS = negative
   b. $N_2 (g) + 3 H_2 (g) \rightarrow 2 NH_3 (g)$ — ΔS = negative
   c. $2 H_2O (l) \rightarrow 2 H_2 (g) + O_2 (g)$ — ΔS = positive
   d. $2 KClO_3 (s) \rightarrow 2 KCl (s) + 3 O_2 (g)$ — ΔS = positive

2. Calculate the ΔS^o of each of the above equations and compare the value to your predictions from #1.
   a. $Mg (s) + Cl_2 (g) \rightarrow MgCl_2 (s)$ — $\Delta S = (90) – [(33) + (223)] = -166$ J/K
   b. $N_2 (g) + 3 H_2 (g) \rightarrow 2 NH_3 (g)$ — $\Delta S = [2(193)] – [(153) + 3(131)] = -160$ J/K
   c. $2 H_2O (l) \rightarrow 2 H_2 (g) + O_2 (g)$ — $\Delta S = [2(131) + 205] – [2(70)] = + 327$ J/K
   d. $2 KClO_3 (s) \rightarrow 2 KCl (s) + 3 O_2 (g)$ — $\Delta S = [2(83) + 3(205)] – [2(143)] = + 495$ J/K

3. Summarize each of the laws of Thermodynamics:
   a. 1st Law of Thermodynamics: $\Delta E = q + w$
   b. 2nd Law of Thermodynamics: $\Delta S_{universe} > 0$ for a spontaneous rxn
   c. 3rd Law of Thermodynamics: ΔS of a perfect crystal at 0 Kelvin is 0

4. List three things the entropy value of a substance depends upon:
   a. Temperature
   b. State of matter
   c. Molecular complexity

5. Define entropy. A measure of the energy randomization or energy dispersal in a system (disorder)

Spontaneity and Gibb's Free Energy

1. Describe how the sign of ΔG determines the spontaneity of a reaction:
   a. $\Delta G > 0$ Non-spontaneous (positive)
   b. $\Delta G < 0$ Spontaneous (negative)
   c. $\Delta G = 0$ At equilibrium

2. Given the following thermodynamic values, determine how temperature effects the spontaneity of the reaction. Is the reaction enthalpy driven, entropy driven, both, or neither.

a.	$\Delta H = + 327$ kJ/mol	$\Delta S = + 127$ J/molK	enthalpy driven	**[entropy driven]**	both	neither
b.	$\Delta H = - 187$ kJ/mol	$\Delta S = - 298$ J/molK	**[enthalpy driven]**	entropy driven	both	neither
c.	$\Delta H = + 194$ kJ/mol	$\Delta S = - 4$ J/molK	enthalpy driven	entropy driven	both	**[neither]**
a.	$\Delta H = - 397$ kJ/mol	$\Delta S = + 653$ J/molK	enthalpy driven	entropy driven	**[both]**	neither

3. For each reaction below, determine the temperature at which the reaction will be spontaneous.
   a. $\Delta H = + 245$ kJ/mol $\Delta S = + 48$ J/molK $T = \Delta H/\Delta S = 245/0.048$ $T > 5104$ K spontaneous
   b. $\Delta H = + 187$ kJ/mol $\Delta S = + 365$ J/molK $T = \Delta H/\Delta S = 187/0.365$ $T > 512$ K spontaneous
   c. $\Delta H = - 456$ kJ/mol $\Delta S = - 38$ J/molK $T = \Delta H/\Delta S = -456/-0.038$ $T < 12000$ K spontaneous
   d. $\Delta H = - 547$ kJ/mol $\Delta S = - 97$ J/molK $T = \Delta H/\Delta S = -547/-0.097$ $T < 5639$ K spontaneous

4. Define Gibb's Free Energy The energy available to do work in a thermodynamic system

5. Using the values of ΔG^o_f, calculate the ΔG^o_{rxn} of each of the equations below. Is each reaction spontaneous at standard conditions?
   a. N_2O_4 (g) → 2 NO_2 (g) $\Delta G = [2(51)] – [100] = 2$ kJ Non-spontaneous
   b. NH_4Cl (s) → HCl (g) + NH_3 (g) $\Delta G = [(-95) + (-16)] – [-203] = 92$ kJ Non-spontan
   c. 3 H_2 (g) + Fe_2O_3 (s) → 2 Fe (s) + 3 H_2O (g) $\Delta G = [3(-229)] – [-742] = 55$ kJ Non-spontaneous
   d. N_2 (g) + 3 H_2(g) → 2 NH_3 (g) $\Delta G = [2(-16)] – [0] = -32$ kJ Spontaneous

Energy & Voltage

1. Write the equation that relates ΔG^o and E^o? $\Delta G^o = -nFE^o$
2. What is n? the number of moles of electrons transferred
3. What is F? Faraday's constant, 96500 J/voltmole
4. For each of the following equations, what must n be?
   a. Cu^{+2} + Mg → Mg^{+2} + Cu n = 2
   b. 2 Ag^{+1} + Sn → Sn^{+2} + 2 Ag n = 2
   c. 2 Al + 3 Pb^{+2} → 3 Pb + 2 Al^{+3} n = 6
5. What is the ΔG^o of the following batteries?
   a. Cu^{+2} + Mg → Mg^{+2} + Cu
   E^o = 2.71 V ΔG^o = -(2 moles)(96500 J/Vmol)(2.71 V) = -523 kJ
   b. 2 Ag^{+1} + Sn → Sn^{+2} + 2 Ag
   E^o = 0.94 V ΔG^o = -(2 moles)(96500 J/Vmol)(0.94 V) = -181 kJ
   c. 2 Al + 3 Pb^{+2} → 3 Pb + 2 Al^{+3}
   E^o = 1.53 V ΔG^o = -(6 moles)(96500 J/Vmol)(1.53 V) = -886 kJ
6. What sign must each of the following be for a battery?
   a. E^o positive
   b. ΔG^o negative

Combination Thermo Problems

1. For the reaction: 2 NO (g) + O_2 (g) → 2 NO_2 (g) ΔH_{rxn}= -114.1 kJ ΔS_{rxn}= -146.5 J/K
   a. How much heat is released when 73.1 g of NO is converted to NO_2?
   73.1 g NO *(1 mole/30 g) = 2.44 moles NO *(-114.1 kJ/2 mole NO) = -139 kJ

   b. Indicate whether the value of ΔG_{rxn}would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.
   $\Delta G = \Delta H - T\Delta S$ since ΔS is negative, as T increased, it will make ΔG more positive as you subtract a negative value making it more positive
   c. The value of S^o = 210.8 J/Kmol for NO. The value of S^o = 240.1 J/Kmol for NO_2. Calculate the value of S^o for O_2
   $\Delta S = (2*S_{NO2}) - (2*S_{NO} + S_{O2})$
   $-146.5 = (2*240.1) - (2*210.8 + X)$
   X = 205 J/K
   d. The bond energy of the N-O bond in NO is 607 kJ/mol. The bond energy of the O-O bond in O_2 is 495 kJ/mol. Calculate the bond energy of the N-O bond in NO_2.

$\Delta H_{rxn} = \Sigma\ Bonds_{react} - \Sigma\ Bonds_{prod}$

$-114.1 = (2*607 + 495) - 2(2*X)$

X = 456 kJ/mole

2. For the substance propane, C_3H_8:

a. Write a balanced equation for the complete combustion of propane with oxygen to yield carbon dioxide and water.

$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

b. Calculate the volume of air at 30 ^oC and 1.0 atm that is needed to burn completely 10.0 grams of propane. Assume that air is 21% O_2.

10 g C_3H_8 *(1 mole/44 g) = 0.227 moles C_3H_8 *(5 mole O_2/1 mole C_3H_8) = 1.136 mole O_2

V = nRT/P = (1.136 mole)(0.0821)(303 K)/(1 atm) = 28.3 L of O_2 needed

Since air is only 21% O_2 you need 134.6 L air

c. The heat of combustion of propane is -2220.1 kJ/mole. Knowing that ΔH_f = -285.3 kJ/mole for H_2O and ΔH_f= -393.5 kJ/mole for CO_2, calculate the ΔH_f for propane.

$\Delta H_{rxn} = \Sigma\Delta H_f^o$ prod – $\Sigma\Delta H_f^o$ react

ΔH_{rxn} = (3*CO_2 + 4*H_2O) – X

-2220.1 = (3*(-393.5) + 4*(-285.3)) –X

-2220.1 = -1180.5 -1141.2 –X

X = -101.6 kJ/mole

d. Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kg of water (C_{water}= 4.18 J/gK), calculate the increase in temperature of the water.

0.682 moles C_3H_8 * (-2220.1 kJ/mole) = -1514 kJ or 1514000 J given to water

$q_{water} = mc\Delta T$ 1514000 J = (8000 g)(4.18 J/gK)ΔT ΔT = 45.3 ^oC

3. For the reaction: $2\ Fe\ (s) + 3/2\ O_2\ (g) \rightarrow Fe_2O_3\ (s)$ ΔH_f= -824 kJ/mol

A 75.0 g sample of Fe is mixed with 11.5 L of O_2 at 2.66 atm and 298 K.

a. Calculate the number of moles of:

i) Fe 1.34 moles

ii) O_2 1.25 moles

b. Identify the limiting reagent. Justify your answer.

Try to use all 1.34 moles of Fe:

1.34 moles Fe *(3/2 moles O_2/2 moles Fe) = 1.005 moles O_2 needed which we have enough of so Fe is limiting reagent

c. Calculate the number of moles of Fe_2O_3 produced if the reaction goes to completion.

1.34 moles Fe *(1 mole Fe_2O_3/2 mole Fe) = 0.67 moles Fe_2O_3

d. If ΔG_f= -740 kJ/mole for Fe_2O_3:

i) Calculate the ΔS_f for Fe_2O_3

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

-740,000 J = -824000 J – (298K)ΔS

ΔS = -282 J/K

ii) Which is more responsible for the spontaneity of the reaction, the enthalpy or the entropy? Justify your answer.

Enthalpy (ΔH) tends to be negative for spontaneous reactions

Entropy (ΔS) tends to be positive for spontaneous reactions

Since the enthalpy is negative and the entropy is negative, the reaction is enthalpy-driven only

e. Knowing that: $2\ FeO\ (s) + ½\ O_2(g) \rightarrow Fe_2O_3\ (s)$ ΔH_{rxn}= -280 kJ

Calculate ΔH_f of FeO (s)

$\Delta H_{rxn} = \Sigma \Delta H_f^o$ prod – $\Sigma \Delta H_f^o$ react

ΔH_{rxn} = (Fe_2O_3) –2*(FeO)

-280 = (-824) –2X

2X = -544

X = -272 kJ/mole

4. Consider pentane (C_5H_{12} MW = 72.15 g/mole)

a. Write the balanced equation for the combustion of pentane to yield CO_2 and H_2O

$C_5H_{12} + 8\ O_2 \rightarrow 5\ CO_2 + 6\ H_2O$

b. What volume of CO_2 at 25 °C and 785 mm Hg will result from combustion of 2.5 g of pentane?

2.5 g*(1 mole/72.15 g)= 0.03465 moles C_5H_{12}*(5 mole CO_2/1 mole C_5H_{12}) = 0.173 mole CO_2

V = nRT/P = (0.173 mole CO_2)(0.0821)(298 K)/(1.033 atm) = 4.103 L

c. Combustion of 5 g of pentane releases 243 kJ of heat. Calcualte $\Delta H_{combustion}$ pentane.

ΔH = q/mole

5 g C_5H_{12}*(1 mole/72.15 g) = 0.0693 moles

ΔH = (-243 kJ/0.0693 mole) = -3506 kJ/mole

d. Under identical conditions, an unknown gas effuses twice as fast as pentane. Calculate the molar mass of this gas.

Rate X/Rate C_5H_{12} = $(72.15/X)^{1/2}$

2/1 = $(72.15/X)^{1/2}$

X = 18 g/mole

e. Draw three structural isomers of pentane.