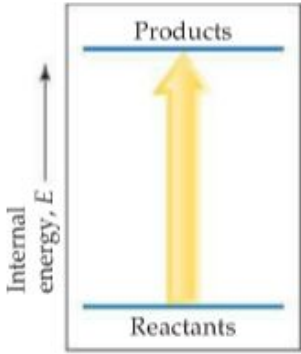
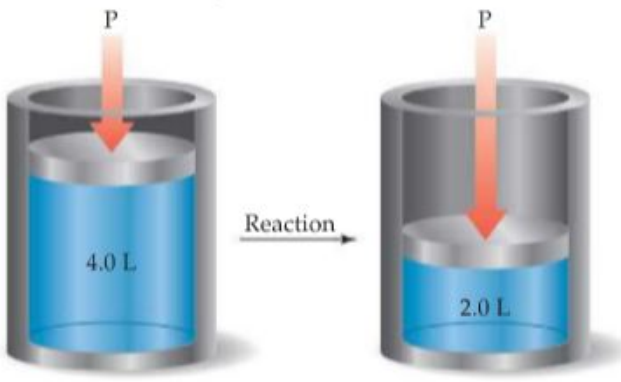
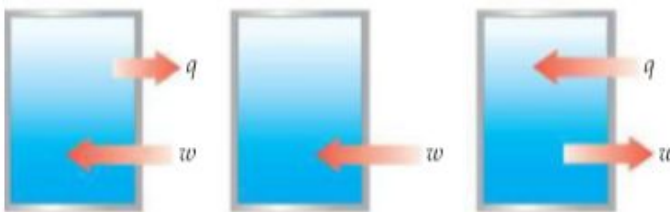


# AP Chemistry Unit Problem Sets.

## Problem Set 1: Heat, Work, Enthalpy, and Internal Energy

1	2
<p>5.2 Consider the accompanying energy diagram.</p> <p>(a) Does this diagram represent an increase or decrease in the internal energy of the system? (b) What sign is given to <math>\Delta E</math> for this process? (c) If there is no work associated with the process, is it exothermic or endothermic?</p> 	<p>5.19 (a) What is work? (b) How do we determine the amount of work done, given the force associated with the work? (c) What is heat? (d) Under what conditions is heat transferred from one object to another?</p>
3	4
<p>5.5 In the cylinder diagrammed below, a chemical process occurs at constant temperature and pressure. (a) Is the sign of <math>w</math> indicated by this change positive or negative? (b) If the process is endothermic, does the internal energy of the system within the cylinder increase or decrease during the change and is <math>\Delta E</math> positive or negative?</p> 	<p>5.3 The contents of the closed box in each of the following illustrations represent a system, and the arrows show the changes to the system during some process. The lengths of the arrows represent the relative magnitudes of <math>q</math> and <math>w</math></p> <p>(a) Which of these processes is endothermic? (b) For which of these processes, if any, is <math>\Delta E &lt; 0</math>? (c) For which process, if any, is there a net gain in internal energy?</p> 
5	6
<p>4.3. Are the following processes exothermic or endothermic?</p> <ol style="list-style-type: none"> <li>When solid KBr is dissolved in water, the solution gets colder.</li> <li>Natural gas (<math>\text{CH}_4</math>) is burned in a furnace.</li> <li>When concentrated <math>\text{H}_2\text{SO}_4</math> is added to water, the solution gets very hot.</li> <li>Water is boiled in a teakettle.</li> <li>the combustion of gasoline in a car engine</li> <li>water condensing on a cold pipe</li> <li><math>\text{CO}_2(s) \longrightarrow \text{CO}_2(g)</math></li> <li><math>\text{F}_2(g) \longrightarrow 2\text{F}(g)</math></li> </ol>	<p>4. Liquid water turns to ice. Is this process endothermic or exothermic? Explain what is occurring using the terms <i>system</i>, <i>surroundings</i>, <i>heat</i>, <i>potential energy</i>, and <i>kinetic energy</i> in the discussion.</p>

7	8
<p><b>31.</b> Calculate <math>\Delta E</math> for each of the following.</p> <p>a. <math>q = -47 \text{ kJ}</math>, <math>w = +88 \text{ kJ}</math></p> <p>b. <math>q = +82 \text{ kJ}</math>, <math>w = -47 \text{ kJ}</math></p> <p>c. <math>q = +47 \text{ kJ}</math>, <math>w = 0</math></p> <p>d. In which of these cases do the surroundings do work on the system?</p>	<p><b>33.</b> If the internal energy of a thermodynamic system is increased by 300. J while 75 J of expansion work is done, how much heat was transferred and in which direction, to or from the system?</p>
9	10
<p><b>35.</b> A sample of an ideal gas at 15.0 atm and 10.0 L is allowed to expand against a constant external pressure of 2.00 atm at a constant temperature. Calculate the work in units of kJ for the gas expansion. (<i>Hint:</i> Boyle's law applies.)</p>	<p><b>34.</b> Calculate the internal energy change for each of the following.</p> <p>a. One hundred (100.) joules of work is required to compress a gas. At the same time, the gas releases 23 J of heat.</p> <p>b. A piston is compressed from a volume of 8.30 L to 2.80 L against a constant pressure of 1.90 atm. In the process, there is a heat gain by the system of 350. J.</p> <p>c. A piston expands against 1.00 atm of pressure from 11.2 L to 29.1 L. In the process, 1037 J of heat is absorbed.</p>

## Problem Set 2: Heat Capacity and Calorimetry

1	2																								
<p>5.48 Two solid objects, A and B, are placed in boiling water and allowed to come to temperature there. Each is then lifted out and placed in separate beakers containing 1000 g water at 10.0 °C. Object A increases the water temperature by 3.50 °C; B increases the water temperature by 2.60 °C. (a) Which object has the larger heat capacity? (b) What can you say about the specific heats of A and B?</p>	<p>5.51 The specific heat of iron metal is 0.450 J/g·K. How many J of heat are necessary to raise the temperature of a 1.05-kg block of iron from 25.0 °C to 88.5 °C?</p>																								
3	4																								
<p>5.50 (a) Which substance in Table 5.2 requires the smallest amount of energy to increase the temperature of 50.0 g of that substance by 10 K? (b) Calculate the energy needed for this temperature change.</p>	<p>5.2. The specific heat capacity of silver is 0.24 J/°C · g.</p> <p>a. Calculate the energy required to raise the temperature of 150.0 g Ag from 273 K to 298 K.</p> <p>b. Calculate the energy required to raise the temperature of 1.0 mole of Ag by 1.0°C (called the <i>molar heat capacity</i> of silver).</p> <p>c. It takes 1.25 kJ of energy to heat a sample of pure silver from 12.0°C to 15.2°C. Calculate the mass of the sample of silver.</p>																								
<p><b>TABLE 5.2</b> ■ Specific Heats of Some Substances at 298 K</p> <table border="1" style="width: 100%; border-collapse: collapse;"> <thead> <tr> <th colspan="2" style="text-align: center;">Elements</th> <th colspan="2" style="text-align: center;">Compounds</th> </tr> <tr> <th style="text-align: left;">Substance</th> <th style="text-align: center;">Specific Heat (J/g·K)</th> <th style="text-align: left;">Substance</th> <th style="text-align: center;">Specific Heat (J/g·K)</th> </tr> </thead> <tbody> <tr> <td>N<sub>2</sub>(g)</td> <td style="text-align: center;">1.04</td> <td>H<sub>2</sub>O(l)</td> <td style="text-align: center;">4.18</td> </tr> <tr> <td>Al(s)</td> <td style="text-align: center;">0.90</td> <td>CH<sub>4</sub>(g)</td> <td style="text-align: center;">2.20</td> </tr> <tr> <td>Fe(s)</td> <td style="text-align: center;">0.45</td> <td>CO<sub>2</sub>(g)</td> <td style="text-align: center;">0.84</td> </tr> <tr> <td>Hg(l)</td> <td style="text-align: center;">0.14</td> <td>CaCO<sub>3</sub>(s)</td> <td style="text-align: center;">0.82</td> </tr> </tbody> </table>		Elements		Compounds		Substance	Specific Heat (J/g·K)	Substance	Specific Heat (J/g·K)	N <sub>2</sub> (g)	1.04	H <sub>2</sub> O(l)	4.18	Al(s)	0.90	CH <sub>4</sub> (g)	2.20	Fe(s)	0.45	CO <sub>2</sub> (g)	0.84	Hg(l)	0.14	CaCO <sub>3</sub> (s)	0.82
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<p>5.4. It takes 585 J of energy to raise the temperature of 125.6 g mercury from 20.0°C to 53.5°C. Calculate the specific heat capacity and the molar heat capacity of mercury.</p>	<p>5.5. A 30.0-g sample of water at 280. K is mixed with 50.0 g water at 330. K. Calculate the final temperature of the mixture assuming no heat loss to the surroundings.</p>																								
7	8																								
<p>5.9. A 150.0-g sample of a metal at 75.0°C is added to 150.0 g H<sub>2</sub>O at 15.0°C. The temperature of the water rises to 18.3°C. Calculate the specific heat capacity of the metal, assuming that all the heat lost by the metal is gained by the water.</p>	<p>6.0. A 110.-g sample of copper (specific heat capacity = 0.20 J/°C · g) is heated to 82.4°C and then placed in a container of water at 22.3°C. The final temperature of the water and copper is 24.9°C. What is the mass of the water in the container, assuming that all the heat lost by the copper is gained by the water?</p>																								
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<p>5.54 (a) When a 3.88-g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee-cup calorimeter (Figure 5.17), the temperature drops from 23.0 °C to 18.4 °C. Calculate <math>\Delta H</math> (in kJ/mol NH<sub>4</sub>NO<sub>3</sub>) for the solution process</p> $\text{NH}_4\text{NO}_3(s) \longrightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq)$ <p>Assume that the specific heat of the solution is the same as that of pure water. (b) Is this process endothermic or exothermic?</p>	<p>6.5. Consider the dissolution of CaCl<sub>2</sub>:</p> $\text{CaCl}_2(s) \longrightarrow \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq) \quad \Delta H = -81.5 \text{ kJ}$ <p>An 11.0-g sample of CaCl<sub>2</sub> is dissolved in 125 g water, with both substances at 25.0°C. Calculate the final temperature of the solution assuming no heat loss to the surroundings and assuming the solution has a specific heat capacity of 4.18 J/°C · g.</p>																								
11	12																								
<p>5.56 A 1.800-g sample of phenol (C<sub>6</sub>H<sub>5</sub>OH) was burned in a bomb calorimeter whose total heat capacity is 11.66 kJ/°C. The temperature of the calorimeter plus contents increased from 21.36 °C to 26.37 °C. (a) Write a balanced chemical equation for the bomb calorimeter reaction. (b) What is the heat of combustion per gram of phenol? Per mole of phenol?</p>	<p>6.8. The combustion of 0.1584 g benzoic acid increases the temperature of a bomb calorimeter by 2.54°C. Calculate the heat capacity of this calorimeter. (The energy released by combustion of benzoic acid is 26.42 kJ/g.) A 0.2130-g sample of vanillin (C<sub>8</sub>H<sub>8</sub>O<sub>3</sub>) is then burned in the same calorimeter, and the temperature increases by 3.25°C. What is the energy of combustion per gram of vanillin? Per mole of vanillin?</p>																								



### Problem Set 3: Heat and Phase Change; Heating and Cooling Curves

1	2																				
<p><b>98.</b> The molar heat of fusion of benzene (<math>C_6H_6</math>) is 9.92 kJ/mol. Its molar heat of vaporization is 30.7 kJ/mol. Calculate the heat required to melt 8.25 g benzene at its normal melting point. Calculate the heat required to vaporize 8.25 g benzene at its normal boiling point. Why is the heat of vaporization more than three times the heat of fusion?</p>	<p><b>95.</b> A substance, X, has the following properties:</p> <table border="1" style="margin-left: auto; margin-right: auto;"> <thead> <tr> <th colspan="4" style="text-align: center;">Specific Heat Capacities</th> </tr> </thead> <tbody> <tr> <td><math>\Delta H_{\text{vap}}</math></td> <td>20. kJ/mol</td> <td>C(s)</td> <td>3.0 J/g · °C</td> </tr> <tr> <td><math>\Delta H_{\text{fus}}</math></td> <td>5.0 kJ/mol</td> <td>C(l)</td> <td>2.5 J/g · °C</td> </tr> <tr> <td>bp</td> <td>75°C</td> <td>C(g)</td> <td>1.0 J/g · °C</td> </tr> <tr> <td>mp</td> <td>-15°C</td> <td></td> <td></td> </tr> </tbody> </table> <p>Sketch a heating curve for this substance starting at -50.0°C and ending at 100°C.</p>	Specific Heat Capacities				$\Delta H_{\text{vap}}$	20. kJ/mol	C(s)	3.0 J/g · °C	$\Delta H_{\text{fus}}$	5.0 kJ/mol	C(l)	2.5 J/g · °C	bp	75°C	C(g)	1.0 J/g · °C	mp	-15°C		
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<p><b>99.</b> What quantity of energy does it take to convert 0.500 kg ice at -20.°C to steam at 250.°C? Specific heat capacities: ice, 2.03 J/g · °C; liquid, 4.2 J/g · °C; steam, 2.0 J/g · °C; <math>\Delta H_{\text{vap}} = 40.7</math> kJ/mol; <math>\Delta H_{\text{fus}} = 6.02</math> kJ/mol.</p>	<p><b>100.</b> Consider a 75.0-g sample of <math>H_2O(g)</math> at 125°C. What phase or phases are present when 215 kJ of energy is removed from this sample?</p>																				
5	6																				
<p><b>11.40</b> The fluorocarbon compound <math>C_2Cl_3F_3</math> has a normal boiling point of 47.6 °C. The specific heats of <math>C_2Cl_3F_3(l)</math> and <math>C_2Cl_3F_3(g)</math> are 0.91 J/g-K and 0.67 J/g-K, respectively. The heat of vaporization for the compound is 27.49 kJ/mol. Calculate the heat required to convert 50.0 g of <math>C_2Cl_3F_3</math> from a liquid at 10.00 °C to a gas at 85.00 °C.</p>	<p><b>11.37</b> For many years drinking water has been cooled in hot climates by evaporating it from the surfaces of canvas bags or porous clay pots. How many grams of water can be cooled from 35 °C to 20 °C by the evaporation of 60 g of water? (The heat of vaporization of water in this temperature range is 2.4 kJ/g. The specific heat of water is 4.18 J/g-K.)</p>																				
7	8																				
<p><b>101.</b> An ice cube tray contains enough water at 22.0°C to make 18 ice cubes that each has a mass of 30.0 g. The tray is placed in a freezer that uses <math>CF_2Cl_2</math> as a refrigerant. The heat of vaporization of <math>CF_2Cl_2</math> is 158 J/g. What mass of <math>CF_2Cl_2</math> must be vaporized in the refrigeration cycle to convert all the water at 22.0°C to ice at -5.0°C? The heat capacities for <math>H_2O(s)</math> and <math>H_2O(l)</math> are 2.03 J/g · °C and 4.18 J/g · °C, respectively, and the enthalpy of fusion for ice is 6.02 kJ/mol.</p>	<p><b>102.</b> A 0.250-g chunk of sodium metal is cautiously dropped into a mixture of 50.0 g water and 50.0 g ice, both at 0°C. The reaction is</p> $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g) \quad \Delta H = -368 \text{ kJ}$ <p>Assuming no heat loss to the surroundings, will the ice melt? Assuming the final mixture has a specific heat capacity of 4.18 J/g · °C, calculate the final temperature. The enthalpy of fusion for ice is 6.02 kJ/mol.</p>																				

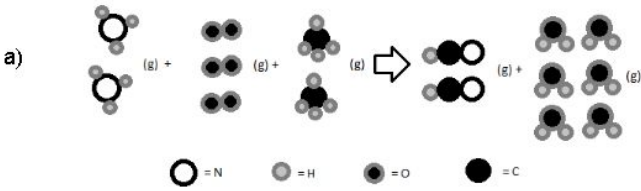
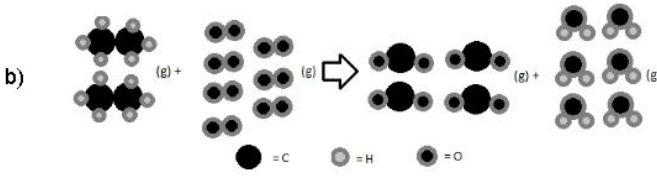
## Problem Set 4: Thermochemical Equations and Heat Stoichiometry

1	2
<p>46. Consider the following reaction:</p> $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -572 \text{ kJ}$ <p>a. How much heat is evolved for the production of 1.00 mole of <math>\text{H}_2\text{O}(\text{l})</math>?</p> <p>b. How much heat is evolved when 4.03 g hydrogen are reacted with excess oxygen?</p> <p>c. How much heat is evolved when 186 g oxygen are reacted with excess hydrogen?</p> <p>d. The total volume of hydrogen gas needed to fill the <i>Hindenburg</i> was <math>2.0 \times 10^8</math> L at 1.0 atm and <math>25^\circ\text{C}</math>. How much heat was evolved when the <i>Hindenburg</i> exploded, assuming all of the hydrogen reacted?</p>	<p>45. The overall reaction in a commercial heat pack can be represented as</p> $4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \longrightarrow 2\text{Fe}_2\text{O}_3(\text{s}) \quad \Delta H = -1652 \text{ kJ}$ <p>a. How much heat is released when 4.00 moles of iron are reacted with excess <math>\text{O}_2</math>?</p> <p>b. How much heat is released when 1.00 mole of <math>\text{Fe}_2\text{O}_3</math> is produced?</p> <p>c. How much heat is released when 1.00 g iron is reacted with excess <math>\text{O}_2</math>?</p> <p>d. How much heat is released when 10.0 g Fe and 2.00 g <math>\text{O}_2</math> are reacted?</p>
3	4
<p>5.42 Consider the following reaction:</p> $\text{CH}_3\text{OH}(\text{g}) \longrightarrow \text{CO}(\text{g}) + 2 \text{H}_2(\text{g}) \quad \Delta H = +90.7 \text{ kJ}$ <p>(a) Is heat absorbed or released in the course of this reaction? (b) Calculate the amount of heat transferred when 45.0 g of <math>\text{CH}_3\text{OH}(\text{g})</math> is decomposed by this reaction at constant pressure. (c) For a given sample of <math>\text{CH}_3\text{OH}</math>, the enthalpy change on reaction is 25.8 kJ. How many grams of hydrogen gas are produced? What is the value of <math>\Delta H</math> for the reverse of the previous reaction? (d) How many kilojoules of heat are released when 50.9 g of <math>\text{CO}(\text{g})</math> reacts completely with <math>\text{H}_2(\text{g})</math> to form <math>\text{CH}_3\text{OH}(\text{g})</math> at constant pressure?</p>	<p>66. Consider the reaction</p> $2\text{HCl}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \longrightarrow \text{BaCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -118 \text{ kJ}$ <p>Calculate the heat when 100.0 mL of 0.500 M HCl is mixed with 300.0 mL of 0.100 M <math>\text{Ba}(\text{OH})_2</math>. Assuming that the temperature of both solutions was initially <math>25.0^\circ\text{C}</math> and that the final mixture has a mass of 400.0 g and a specific heat capacity of <math>4.18 \text{ J}^\circ\text{C}^{-1} \cdot \text{g}</math>, calculate the final temperature of the mixture.</p>
5	6
<p>Meals-ready-to-eat (MREs) are military meals that can be heated on a flameless heater. The heat is produced by the following reaction:</p> $\text{Mg}(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \rightarrow \text{Mg}(\text{OH})_2(\text{s}) + \text{H}_2(\text{g}) \quad \Delta H = -355 \text{ kJ/mol}$ <p>Calculate the number of grams of Mg needed for this reaction to release enough energy to increase the temperature of 75 mL of water from <math>21^\circ\text{C}</math> to <math>79^\circ\text{C}</math>.</p>	<p>133. The preparation of <math>\text{NO}_2(\text{g})</math> from <math>\text{N}_2(\text{g})</math> and <math>\text{O}_2(\text{g})</math> is an endothermic reaction:</p> $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow \text{NO}_2(\text{g}) \text{ (unbalanced)}$ <p>The enthalpy change of reaction for the balanced equation (with lowest whole-number coefficients) is <math>\Delta H = 67.7 \text{ kJ}</math>. If <math>2.50 \times 10^2</math> mL <math>\text{N}_2(\text{g})</math> at <math>100.^\circ\text{C}</math> and 3.50 atm and <math>4.50 \times 10^2</math> mL <math>\text{O}_2(\text{g})</math> at <math>100.^\circ\text{C}</math> and 3.50 atm are mixed, what amount of heat is necessary to synthesize the maximum yield of <math>\text{NO}_2(\text{g})</math>?</p>

## Problem Set 5: Hess's Law

1	2
<p>Calculate the <math>\Delta H_{\text{rxn}}</math> of  <math>\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})</math>                      based on the following reactions:                      A) <math>\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) \quad \Delta H = +180 \text{ kJ}</math>                      B) <math>2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \quad \Delta H = -112 \text{ kJ}</math></p>	<p>Calculate the <math>\Delta H_{\text{rxn}}</math> of  <math>\text{C}_{(\text{graphite})} \rightarrow \text{C}_{(\text{diamond})}</math>                      based on the following reactions:                      A) <math>\text{C}_{(\text{graphite})} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \quad \Delta H = -394 \text{ kJ}</math>                      B) <math>\text{C}_{(\text{diamond})} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \quad \Delta H = -396 \text{ kJ}</math></p>
3	4
<p><b>5.61</b> Calculate the enthalpy change for the reaction  <math>\text{P}_4\text{O}_6(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s})</math>                      given the following enthalpies of reaction:  <math>\text{P}_4(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_6(\text{s}) \quad \Delta H = -1640.1 \text{ kJ}</math>  <math>\text{P}_4(\text{s}) + 5\text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s}) \quad \Delta H = -2940.1 \text{ kJ}</math></p>	<p>Calculate the <math>\Delta H_{\text{rxn}}</math> of  <math>2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})</math>                      based on the following reactions:                      A) <math>2\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g}) \quad \Delta H = -84.7 \text{ kJ}</math>                      B) <math>\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \quad \Delta H = -393.5 \text{ kJ}</math>                      C) <math>\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) \quad \Delta H = -242 \text{ kJ}</math></p>
5	6
<p><b>5.62</b> From the enthalpies of reaction  <math>2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \quad \Delta H = -483.6 \text{ kJ}</math>  <math>3\text{O}_2(\text{g}) \rightarrow 2\text{O}_3(\text{g}) \quad \Delta H = +284.6 \text{ kJ}</math>                      calculate the heat of the reaction  <math>3\text{H}_2(\text{g}) + \text{O}_3(\text{g}) \rightarrow 3\text{H}_2\text{O}(\text{g})</math></p>	<p><b>73.</b> Given the following data  <math>2\text{O}_3(\text{g}) \rightarrow 3\text{O}_2(\text{g}) \quad \Delta H = -427 \text{ kJ}</math>  <math>\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g}) \quad \Delta H = 495 \text{ kJ}</math>  <math>\text{NO}(\text{g}) + \text{O}_3(\text{g}) \rightarrow \text{NO}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = -199 \text{ kJ}</math>                      calculate <math>\Delta H</math> for the reaction  <math>\text{NO}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{NO}_2(\text{g})</math></p>
7	8
<p><b>72.</b> Given the following data  <math>2\text{ClF}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{Cl}_2\text{O}(\text{g}) + \text{F}_2\text{O}(\text{g}) \quad \Delta H = 167.4 \text{ kJ}</math>  <math>2\text{ClF}_3(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{Cl}_2\text{O}(\text{g}) + 3\text{F}_2\text{O}(\text{g}) \quad \Delta H = 341.4 \text{ kJ}</math>  <math>2\text{F}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{F}_2\text{O}(\text{g}) \quad \Delta H = -43.4 \text{ kJ}</math>                      calculate <math>\Delta H</math> for the reaction  <math>\text{ClF}(\text{g}) + \text{F}_2(\text{g}) \rightarrow \text{ClF}_3(\text{g})</math></p>	<p><b>5.64</b> Given the data  <math>\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) \quad \Delta H = +180.7 \text{ kJ}</math>  <math>2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \quad \Delta H = -113.1 \text{ kJ}</math>  <math>2\text{N}_2\text{O}(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = -163.2 \text{ kJ}</math>                      use Hess's law to calculate <math>\Delta H</math> for the reaction  <math>\text{N}_2\text{O}(\text{g}) + \text{NO}_2(\text{g}) \rightarrow 3\text{NO}(\text{g})</math></p>
9	10
<p><b>110.</b> Given the following data  <math>\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g}) \quad \Delta H^\circ = -23 \text{ kJ}</math>  <math>3\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow 2\text{Fe}_3\text{O}_4(\text{s}) + \text{CO}_2(\text{g}) \quad \Delta H^\circ = -39 \text{ kJ}</math>  <math>\text{Fe}_3\text{O}_4(\text{s}) + \text{CO}(\text{g}) \rightarrow 3\text{FeO}(\text{s}) + \text{CO}_2(\text{g}) \quad \Delta H^\circ = 18 \text{ kJ}</math>                      calculate <math>\Delta H^\circ</math> for the reaction  <math>\text{FeO}(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Fe}(\text{s}) + \text{CO}_2(\text{g})</math></p>	<p><b>5.63</b> From the enthalpies of reaction  <math>\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightarrow 2\text{HF}(\text{g}) \quad \Delta H = -537 \text{ kJ}</math>  <math>\text{C}(\text{s}) + 2\text{F}_2(\text{g}) \rightarrow \text{CF}_4(\text{g}) \quad \Delta H = -680 \text{ kJ}</math>  <math>2\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g}) \quad \Delta H = +52.3 \text{ kJ}</math>                      calculate <math>\Delta H</math> for the reaction of ethylene with <math>\text{F}_2</math>:  <math>\text{C}_2\text{H}_4(\text{g}) + 6\text{F}_2(\text{g}) \rightarrow 2\text{CF}_4(\text{g}) + 4\text{HF}(\text{g})</math></p>
11	12
<p><b>74.</b> Calculate <math>\Delta H</math> for the reaction  <math>\text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})</math>                      given the following data:  <math>2\text{NH}_3(\text{g}) + 3\text{N}_2\text{O}(\text{g}) \rightarrow 4\text{N}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) \quad \Delta H = -1010. \text{ kJ}</math>  <math>\text{N}_2\text{O}(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{l}) + \text{H}_2\text{O}(\text{l}) \quad \Delta H = -317 \text{ kJ}</math>  <math>2\text{NH}_3(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{l}) + \text{H}_2\text{O}(\text{l}) \quad \Delta H = -143 \text{ kJ}</math>  <math>\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \quad \Delta H = -286 \text{ kJ}</math></p>	<p><b>76.</b> Given the following data  <math>\text{P}_4(\text{s}) + 6\text{Cl}_2(\text{g}) \rightarrow 4\text{PCl}_3(\text{g}) \quad \Delta H = -1225.6 \text{ kJ}</math>  <math>\text{P}_4(\text{s}) + 5\text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s}) \quad \Delta H = -2967.3 \text{ kJ}</math>  <math>\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{PCl}_5(\text{g}) \quad \Delta H = -84.2 \text{ kJ}</math>  <math>\text{PCl}_3(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{Cl}_3\text{PO}(\text{g}) \quad \Delta H = -285.7 \text{ kJ}</math>                      calculate <math>\Delta H</math> for the reaction  <math>\text{P}_4\text{O}_{10}(\text{s}) + 6\text{PCl}_5(\text{g}) \rightarrow 10\text{Cl}_3\text{PO}(\text{g})</math></p>

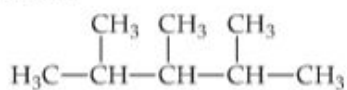
## Problem Set 6: Enthalpy of Formation

1	2
<p>For each of the following compounds, write a balanced thermochemical equation depicting the formation of one mole of the compound from its elements in their standard states and use the reference table to obtain the value of <math>\Delta H_f^\circ</math>.</p> <p>a) <math>\text{NO}_2(\text{g})</math>            b) <math>\text{SO}_3(\text{g})</math>            c) <math>\text{NaBr}(\text{s})</math>            d) <math>\text{Pb}(\text{NO}_3)_2(\text{s})</math></p>	<p>123. Which of the following substances have an enthalpy of formation equal to zero?</p> <p>a. <math>\text{Cl}_2(\text{g})</math>            b. <math>\text{H}_2(\text{g})</math>            c. <math>\text{N}_2(\text{l})</math>            d. <math>\text{Cl}(\text{g})</math></p>
3	4
<p>The <math>\Delta H_f^\circ</math> values of the two allotropes of oxygen, <math>\text{O}_2</math> and <math>\text{O}_3</math>, are 0 and 142.2 kJ/mol, respectively, at 25°C. Which is the more stable form at this temperature? Explain.</p>	<p>5.73 Complete combustion of 1 mol of acetone (<math>\text{C}_3\text{H}_6\text{O}</math>) liberates 1790 kJ:</p> $\text{C}_3\text{H}_6\text{O}(\text{l}) + 4 \text{O}_2(\text{g}) \longrightarrow 3 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\text{l})$ <p style="text-align: right;"><math>\Delta H^\circ = -1790 \text{ kJ}</math></p> <p>(a) Using this information together with data from Appendix C, calculate the enthalpy of formation of acetone.</p> <p>(b) Calculate the kilograms of heat released per gram of acetone when it reacts with oxygen.</p>
5	6
<p>5.71 Using values from Appendix C, calculate the standard enthalpy change for each of the following reactions:</p> <p>(a) <math>2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{SO}_3(\text{g})</math>            (b) <math>\text{Mg}(\text{OH})_2(\text{s}) \longrightarrow \text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l})</math>            (c) <math>\text{N}_2\text{O}_4(\text{g}) + 4 \text{H}_2(\text{g}) \longrightarrow \text{N}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{g})</math>            (d) <math>\text{SiCl}_4(\text{l}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{SiO}_2(\text{s}) + 4 \text{HCl}(\text{g})</math></p>	<p>Use the values of <math>\Delta H_f^\circ</math> in the appendix to calculate <math>\Delta H^\circ</math> for the following reactions:</p> <p>a) </p> <p>b) </p>



7

**5.75** Gasoline is composed primarily of hydrocarbons, including many with eight carbon atoms, called *octanes*. One of the cleanest-burning octanes is a compound called 2,3,4-trimethylpentane, which has the following structural formula:



The complete combustion of one mole of this compound to  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{g})$  leads to  $\Delta H^\circ = -5064.9 \text{ kJ/mol}$ . (a) Write a balanced equation for the combustion of 1 mol of  $\text{C}_8\text{H}_{18}(\text{l})$ . (b) Write a balanced equation for the formation of  $\text{C}_8\text{H}_{18}(\text{l})$  from its elements. (c) By using the information in this problem and data in Table 5.3, calculate  $\Delta H_f^\circ$  for 2,3,4-trimethylpentane.

9

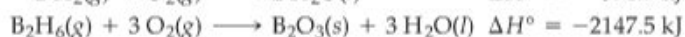
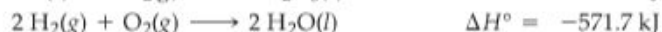
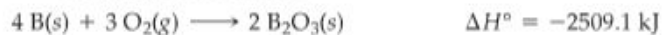
**81.** The Ostwald process for the commercial production of nitric acid from ammonia and oxygen involves the following steps:



- Use the values of  $\Delta H_f^\circ$  in Appendix 4 to calculate the value of  $\Delta H^\circ$  for each of the preceding reactions.
- Write the overall equation for the production of nitric acid by the Ostwald process by combining the preceding equations. (Water is also a product.) Is the overall reaction exothermic or endothermic?

8

**5.103 (a)** Calculate the standard enthalpy of formation of gaseous diborane ( $\text{B}_2\text{H}_6$ ) using the following thermochemical information:



10

**89.** Ethanol ( $\text{C}_2\text{H}_5\text{OH}$ ) has been proposed as an alternative fuel. Calculate the standard enthalpy of combustion per gram of liquid ethanol **using appropriate data from the appendix.**



**Problem Set 7: Additional Exercises**

1	2
<p><b>113.</b> A sample of nickel is heated to 99.8°C and placed in a coffee-cup calorimeter containing 150.0 g water at 23.5°C. After the metal cools, the final temperature of metal and water mixture is 25.0°C. If the specific heat capacity of nickel is 0.444 J/°C · g, what mass of nickel was originally heated? Assume no heat loss to the surroundings.</p>	<p><b>104.</b> When 1.00 L of 2.00 M Na<sub>2</sub>SO<sub>4</sub> solution at 30.0°C is added to 2.00 L of 0.750 M Ba(NO<sub>3</sub>)<sub>2</sub> solution at 30.0°C in a calorimeter, a white solid (BaSO<sub>4</sub>) forms. The temperature of the mixture increases to 42.0°C. Assuming that the specific heat capacity of the solution is 6.37 J/°C · g and that the density of the final solution is 2.00 g/mL, calculate the enthalpy change per mole of BaSO<sub>4</sub> formed.</p>
3	4
<p><math>12 \text{ CO}_2(\text{g}) + 11 \text{ H}_2\text{O}(\text{l}) \longrightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11} + 12 \text{ O}_2(\text{g})</math> <math>\Delta H = 5645 \text{ kJ}</math></p> <p><b>[5.111]</b> It is estimated that the net amount of carbon dioxide fixed by photosynthesis on the landmass of Earth is <math>5.5 \times 10^{16}</math> g/yr of CO<sub>2</sub>. Assume that all this carbon is converted into glucose. <b>(a)</b> Calculate the energy stored by photosynthesis on land per year in kJ. <b>(b)</b> Calculate the average rate of conversion of solar energy into plant energy in MW (1W = 1 J/s). A large nuclear power plant produces about 10<sup>3</sup> MW. The energy of how many such nuclear power plants is equivalent to the solar energy conversion?</p>	<p><b>5.100</b> How many grams of methane [CH<sub>4</sub>(g)] must be combusted to heat 1.00 kg of water from 25.0 °C to 90.0 °C, assuming H<sub>2</sub>O(l) as a product and 100% efficiency in heat transfer?</p>
5	6
7	8
9	10
11	12