***Chemistry***

**5: Thermochemistry**

**5.3: Enthalpy**

39. Explain how the heat measured in Example 5.5 differs from the enthalpy change for the exothermic reaction described by the following equation:



Solution

The enthalpy change of the indicated reaction is for exactly 1 mol HCL and 1 mol NaOH; the heat in the example is produced by 0.0500 mol HCl and 0.0500 mol NaOH.

41. Calculate the enthalpy of solution (Δ*H* for the dissolution) per mole of NH4NO3 under the conditions described in Example 5.6.

Solution

The molar mass of NH4NO3 is 80.0423 g/mol. From the example, 1000 J is required to dissolve 3.21 g of NH4NO3. One mole under the same conditions would require

.

(The heat of solution is positive because the process is endothermic.)

43. Calculate the enthalpy of solution (Δ*H* for the dissolution) per mole of CaCl2.

Solution

The molar mass of CaCl­2­ is 40.078 + 2(35.4572) = 110.992 g/mol. In Exercise 25, 3.0 g of CaCl2 dissolved in water releases a heat of 2.2 kJ. Therefore,

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45. How much heat is produced by burning 4.00 moles of acetylene under standard state conditions?

Solution

The heat of combustion is –1301.1 as given in Table 5.2.

Heat released = 4.00 mol × (–1301.1 kJ/mol) = 5204.4 kJ

47. How many moles of isooctane must be burned to produce 100 kJ of heat under standard state conditions?

Solution

The value of Δ*H*comb = –5461 kJ/mol. To produce 100 kJ requires:

.

49. When 2.50 g of methane burns in oxygen, 125 kJ of heat is produced. What is the enthalpy of combustion per mole of methane under these conditions?

Solution

The molar mass of CH4 is 16.04 g/mol.

Find the mole of CH4 present: 

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51. A sample of 0.562 g of carbon is burned in oxygen in a bomb calorimeter, producing carbon dioxide. Assume both the reactants and products are under standard state conditions, and that the heat released is directly proportional to the enthalpy of combustion of graphite. The temperature of the calorimeter increases from 26.74 °C to 27.93 °C. What is the heat capacity of the calorimeter and its contents?

Solution

The heat produced is found from the enthalpy of combustion:.

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This is the heat released.

*q =* heat capacity × ∆*T*

18.4 kJ = heat capacity × (27.93 ºC – 26.74 ºC) = specific heat × 1.19 ºC

specific heat = 15.5 kJ/ºC

53. Homes may be heated by pumping hot water through radiators. What mass of water will provide the same amount of heat when cooled from 95.0 to 35.0 °C, as the heat provided when 100 g of steam is cooled from 110 °C to 100 °C.

Solution

*q = mc∆T*. The amount of heat transferred from steam is:

*q*steam = 100 g × 1.864 J/(g °C) × (110 – 100) °C = 1864 J

*q*water = *m*water× 4.184 J/(g °C) × (95 – 35) °C = *m*water× 251.04 J/g = *q*steam = 1864 J

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55. Does the standard enthalpy of formation of H2O(*g*) differ from Δ*H*° for the reaction?

Solution

Yes. The standard enthalpy of formation can be determined for anything, including H2O(*g*), and water does not have to be liquid. In this case, it’s the gas-phase water that is the substance for which the heat of formation is to be found. However, the heat of this reaction is defined for two moles of H2O(*g*), thus the heat of formation is half of the heat of the reaction.

57. How many kilojoules of heat will be released when exactly 1 mole of manganese, Mn, is burned to form Mn3O4(*s*) at standard state conditions?

Solution

This process requires 3 mol of Mn. For 1 mol, .

59. The following sequence of reactions occurs in the commercial production of aqueous nitric acid:







Determine the total energy change for the production of one mole of aqueous nitric acid by this process.

Solution

Enough material must be produced in each stage to proceed with the next. So  of reaction 1 produces the NO required in reaction 2. But  units of reaction 2 are required to provide enough NO2 for reaction 3. Up to this stage, the heat produced is:



to have the material to proceed with reaction 3. Therefore, from beginning to end,

−850 + (−139) = −989 kJ are released. The question asks for the enthalpy change for 1 mole.

Therefore, division of the last answer by 2 gives −495 kJ/mol.

61. From the molar heats of formation in Appendix G, determine how much heat is required to evaporate one mole of water: 

Solution

The heat of formation of H2O(*l*) is –285.83 kJ/mol. The heat of formation of H2O(*g*) is –241.82 kJ/mol. For the reaction:

,

the enthalpy of the reaction is the difference of heats of formation of the reactant and the product; thus, it is –241.82 – (–285.83) = 44.01 kJ/mol.

63. Calculate  for the process



from the following information:



Solution

Add the two equations and their heat together.



65. Calculate Δ*H* for the process 

from the following information:



Solution

Reverse the direction of both equations and add the new equations and enthalpies.



67. Calculate the standard molar enthalpy of formation of NO(*g*) from the following data:



Solution

Hess’s law can be applied to the two equations by reversing the direction of the second equation. The first equation is a formation reaction and is so indicated by writing  .

Adding the equations yields:.

This is the heat of formation of 2 mol of NO. For 1 mol,

.

69. Using the data in Appendix G, calculate the standard enthalpy change for each of the following reactions:

(a) 

(b) 

(c) 

(d) 

Solution

(a)

;

(b) ;

(c) ;

(d) 

71. The decomposition of hydrogen peroxide, H2O2, has been used to provide thrust in the control jets of various space vehicles. Using the data in Appendix G, determine how much heat is produced by the decomposition of exactly 1 mole of H2O2 under standard conditions.



Solution



The value relates to the decomposition of 2 mol of hydrogen peroxide. For 1 mol,.

73. Calculate the enthalpy of combustion of butane, C4H10(*g*) for the formation of H2O(*g*) and CO2(*g*). The enthalpy of formation of butane is −126 kJ/mol.

Solution



75. The white pigment TiO2 is prepared by the reaction of titanium tetrachloride, TiCl4, with water vapor in the gas phase:.

How much heat is evolved in the production of exactly 1 mole of TiO2(*s*) under standard state conditions?

Solution



Thus, 67.1 kJ of heat is evolved.

77. In the early days of automobiles, illumination at night was provided by burning acetylene, C2H2. Though no longer used as auto headlamps, acetylene is still used as a source of light by some cave explorers. The acetylene is (was) prepared in the lamp by the reaction of water with calcium carbide, CaC2:

.

Calculate the standard enthalpy of the reaction. The  of CaC2 is –15.14 kcal/mol.

Solution

convert –15.14 kcal to kJ:

–15.14 kcal × 4.184 kJ/kcal = 63.35 kJ/mol

79. The enthalpy of combustion of hard coal averages −35 kJ/g, that of gasoline, −1.28 × 105 kJ/gal. How many kilograms of hard coal provide the same amount of heat as is available from 1.0 gallon of gasoline? Assume that the density of gasoline is 0.692 g/mL (the same as the density of isooctane).

Solution

the amount of heat produced by burning of 1.0 gallon of gasoline is:

*q* = 1.0 gallon × (–1.28 × 105 kJ/gal) = –1.28 × 105 kJ

Mass × (–35 kJ/g) = –1.28 × 105 kJ

Mass = 3657 g or 3.7 kg

81. Among the substances that react with oxygen and that have been considered as potential rocket fuels are diborane [B2H6, produces B2O3(*s*) and H2O(*g*)], methane [CH4, produces CO2(*g*) and H2O(*g*)], and hydrazine [N2H4, produces N2(*g*) and H2O(*g*)]. On the basis of the heat released by 1.00 g of each substance in its reaction with oxygen, which of these compounds offers the best possibility as a rocket fuel? The of B2H6(*g*), CH4(*g*), and N2H4(*l*) may be found in Appendix G.

Solution

Write the balanced equation for each reaction.



Calculate the heat released per mole; then, calculate the heat per gram.



Calculate the heat per mole released per gram.

For B2H6: 

For CH4: 

For N2H4: 

On the assumption that the best rocket fuel is the one that gives off the most heat, B2H6 is the prime candidate. Other things must be considered, however. For example, the moles of gaseous product formed are related to the specific impulse of the fuel, toxicity of products, cost, and ability to contain original fuel (stability and corrosiveness).

83. Ethylene, C2H2, a byproduct from the fractional distillation of petroleum, is fourth among the 50 chemical compounds produced commercially in the largest quantities. About 80% of synthetic ethanol is manufactured from ethylene by its reaction with water in the presence of a suitable catalyst.

Using the data in the table in Appendix G, calculate Δ*H*° for the reaction.

Solution



85. Propane, C3H8, is a hydrocarbon that is commonly used as a fuel.

(a) Write a balanced equation for the complete combustion of propane gas.

(b) Calculate the volume of air at 25 °C and 1.00 atmosphere that is needed to completely combust 25.0 grams of propane. Assume that air is 21.0 percent O2 by volume. (Hint: we will see how to do this calculation in a later chapter on gases—for now use the information that 1.00 L of air at 25 °C and 1.00 atm contains 0.275 g of O2 per liter.)

(c) The heat of combustion of propane is –2219.2 kJ/mol. Calculate the heat of formation, , of propane given that  of H2O(*l*) = –285.8kJ/mol and  of CO2(*g*) = –393.5 kJ/mol.

(d) Assuming that all of the heat released in burning 25.0 grams of propane is transferred to 4.00 kilograms of water, calculate the increase in temperature of the water.

Solution

(a) 

(b) Determine the number of moles of O2 required and from that the number of grams. Then use the density to find the volume.

;

; (c) ; (d) 

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