Enthalpy Wrap-Up Problems – Answer Key

1. Carbon is found in several forms. Graphite and diamond are common forms known for millennia; other forms, such as buckyballs, were discovered only in the last few decades. Combustion of graphite in the presence of oxygen forms carbon dioxide as a product

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

releasing 394 kJ of energy per mole of C atoms. Combustion of buckyballs, C_{60}, however

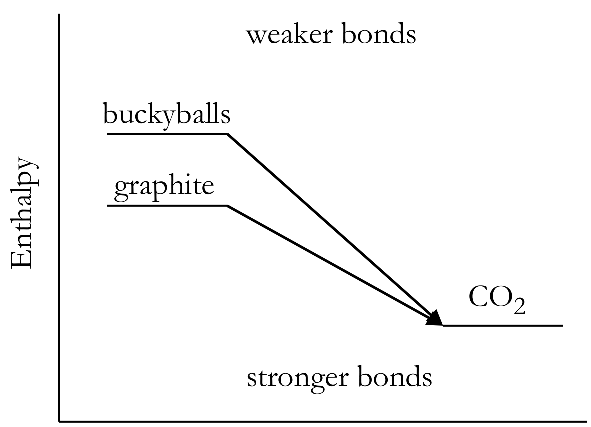
$$C_{60}(s) + 60O_2(g) \rightarrow 60CO_2(g)$$

releases approximately 26,100 kJ of energy per mole of C_{60}. Which of these forms of carbon has the stronger average carbon-carbon bond? Assume that for graphite and buckyballs, the number of bonds per carbon atom is identical. Clearly explain your reasoning.

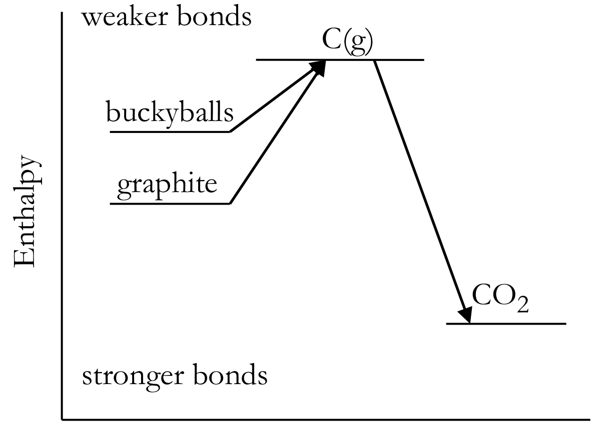
Answer. To compare the two we first convert the energy released during the combustion of buckyballs from kJ/mol C_{60} to kJ/mol C; thus

$$\frac{26100\text{ kJ}}{\text{mol C}_{60}} \times \frac{1\text{ mol C}_{60}}{60\text{ mol C}} = \frac{435\text{ kJ}}{\text{mol C}}$$

Converting a mole of buckyballs to CO_2 releases more energy per mole of carbon than converting a mole of graphite to CO_2. We know that a reaction is exothermic (energy released) when we move from reactants with weaker bonds to products with stronger bonds. Because both reactions have the same final state (CO_2) and because buckyballs release more energy per mole of C, this form of carbon must have the weaker bonds (see figure below). Graphite, therefore, has the stronger bonds.



An alternative way to reach the same conclusion is to imagine reactions converting both graphite and buckyballs to gas-phase carbon atoms and then allowing the atoms to react with O_2 to form CO_2; the energy diagram now looks like this:



The energy needed to convert buckyballs to C*(g)* is less than the energy needed to convert graphite to C*(g)*; thus, the carbon-carbon bonds in buckyballs are weaker than those in graphite.

2. In an experiment to determine the enthalpy of dissolution for solid NaOH

$$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

you dissolve 2.0 g of solid NaOH in 100.0 mL of water and observes a temperature increase. In a second experiment you add 4.0 g of NaOH to 200.0 mL of water. Do you expect the change in temperature for the second experiment to be: (a) the same as that of the first experiment; (b) approximately twice that of the first experiment; (c) approximately four times that of the first experiment; (d) approximately one-half of that for the first experiment; or (e) approximately one-fourth of that for the first experiment?

Explain your reasoning.

Answer. If the amount of heat released when 2.0 g of NaOH dissolves is x, then

$$q_{\text{rxn1}} = x$$

The amount of heat released when dissolving 4.0 g of NaOH is twice as great, or $2x$, and $q_{\text{rxn2}} = 2x$. In each case the amount of heat absorbed by the solution, q_{soln}, is equal in magnitude but opposite in sign from q_{rxn}; thus, assuming a density of 1.00 g/mL, we find that

$$q_{\text{soln1}} = (100.0\text{ g}) \times (4.184\text{ J/g}\bullet°\text{C}) \times \Delta T_1 = -x$$

$$q_{\text{soln2}} = (200.0\text{ g})\times(4.184\text{ J/g}\bullet°\text{C})\times\Delta T_2 = -2x$$

Solving each equation for ΔT shows that ΔT_1 and ΔT_2 is the same for both reactions.