1. A (a) is a statement of the second law of thermodynamics; (b) is not true because the internal energy of a substance is composed of not only the kinetic energy partitions; there can also be internal energy in chemical potential energy, nuclear energy, etc.; temperature does not directly measure these energy forms; (c) temperature rises with increasing particle kinetic energy, so (c) is not true.

2. C One calorie is the amount of heat required to increase the temperature of 1 g of water by 1 K. This equals 4.18 J.

3. C *Dulong and Petit's law* (true for normal temperature ranges and simple crystal structures) states that the molar heat capacity of a solid element is a constant about 25 J mol^{-1} K^{-1}. The equivalent environment of the atoms determines the partitions (degrees of freedom) into which kinetic energy (heat) can flow, giving an equivalent temperature increase for a given amount of heat, regardless of particle size. This is analogous to the fact that the molar heat capacity of an ideal gas is 3/2 R, regardless of particle size.
   To quantitatively solve this problem, recognize that the molar heat capacity of a substance is the product of its specific heat and its atomic (molecular) weight. We can generate a conversion factor, which will give us our answer:

$$\left(25\ \frac{\mathrm{J}}{\mathrm{mol\,K}}\right)\left(\frac{\mathrm{K\,g}}{20\ \mathrm{J}}\right) = 125\ \mathrm{g}$$

4. B This should be pretty straight-forward, but realize that (d) is incorrect because the internal energy can change also due to work performed *on* (making ΔU more positive) or *by* (making ΔU more negative) the system.

5. C *Hess' Law* allows us to *model* the combustion of carbon monoxide to carbon monoxide by conceiving of first running the CO formation reaction in reverse (ΔH = +110.5 kJ mol^{-1}) to liberate the elements in their standard states, which we then react to form CO_2 (ΔH = -393.5 kJ mol^{-1}) Note that because the standard enthalpies of the elements in their standard states is zero, the difference in stoichiometrical coefficients in oxygen here is insignificant.

6. B Notice that when our reaction at stoichiometric values has gone to completion, our system will have changed from 3 moles of gas to 2 moles of gas, and our volume will have decreased. Therefore, the environment has performed pressure-volume work on our system, which gains energy thereby. Even though the system will have lost heat due to the exothermic process, it will have gained some internal energy due to this work. Therefore, the loss of internal energy is less than the loss of heat. Below is a statement of the second law of thermodynamics, showing that the change in internal energy is equal to the amount of heat that flows minus the work performed *by* the system:

$$\Delta U = \Delta H - W$$

7. D For all but methane, the elements are listed in their standard states, which by convention are thus assigned a value of zero for standard enthalpy of formation (it takes no energy to form these from their standard states; they are the standard states.) Grounding energy in this way allows us to compute the standard enthalpy of reactions by subtracting the standard enthalpies of reactant species (in stoichiometrical amounts) from the standard enthalpies of product (see #5).

8. D We say that enthalpy is a state function because the enthalpy differences in a transformation between states is not path dependent (depending only on the difference between the initial and final states). Therefore, we can use Hess' law to model complex chemical reactions as simply the breaking reactant bonds, the subsequent reformation of the standard states (see #5 & #7), followed by reaction of the standard states to form product.

9. A Bond energies can give a decent estimate of the enthalpy of a reaction, which we derive by imagining the breaking of the reactant bonds (endothermic) and the subsequent forming of the product bonds from the resulting unbonded atoms (exothermic).