1. If the same amount of heat is added to 3.00 g samples of different substances, what quantity do you need to tell you the substance that would experience the **smallest** temperature change?
2. Thermodynamics answers what types of questions about chemical systems?
3. Given this thermochemical equation: $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H°_{rxn}$= -394kJ, answer the following: (a) the enthalpy change for the reverse reaction, (b) the enthalpy change for forming 1/2 mole of $CO_2(g)$, and (c) the relationship of $\Delta H°_{rxn}$ to $\Delta H°_f$.
4. A 50.000g sample of water with an initial temperature of 22.0°C reaches a final temperature of 77.5°C. How much heat in J did it take to raise the temperature of the water? The specific heat capacity of water is 4.184 J g^{-1} $°C^{-1}$.
5. A 50.00g sample of aluminum with an initial temperature of 22.0°C reaches a final temperature of 77.5°C. How much heat in J did it take to raise the temperature of the aluminum? The specific heat capacity of aluminum is 0.902 J g^{-1} $°C^{-1}$.
6. A heat of 11600J was applied to raise a 50.000g sample of water (specific heat capacity of water is 4.184 J g^{-1} $°C^{-1}$) from an initial temperature of 22.0°C to a final temperature. What is that final temperature of the water?
7. A heat of 2503J was applied to raise a 50.00g sample of aluminum (specific heat capacity of aluminum is 0.902 J g^{-1} $°C^{-1}$) from an initial temperature of 21.0°C to a final temperature. What is that final temperature of the aluminum?
8. A 27.1 g block of magnesium with an initial temperature of 90.5°C is dropped into a calorimeter containing 100.0g of water at 18.5°C. The specific heat capacity of water is 4.184J/gK. The system reaches an equilibrium temperature of 23.1°C. What is the specific heat capacity of magnesium from these data?
9. A student wishes to determine the calorimeter constant, q*cal* for a coffee cup calorimeter. A 75.60 g sample of iridium (specific heat capacity of 0.1300J/gK) with an initial temperature of 98.60°C is placed into a calorimeter containing 100.0g of water at a temperature of 14.50°C. The system reaches a final temperature of 15.70°C. The specific heat capacity of water is 4.184J/gK. What is q*cal* from these data?
10. The heat capacity of a bomb (constant volume) calorimeter is found to be 7.45 kJ $°C^{-1}$. When a 1.652g sample of sulfur is burned in the calorimeter, the temperature of the calorimeter increases from 23.19°C to 28.63°C. Calculate the ΔH_{rxn} for the following reaction from these data: $S(s) + O_2(g) \rightarrow SO_2(g)$.
11. When 0.400 mol of propane (C_3H_8) gas is burned in excess oxygen, 888 k**J** of heat is evolved. (a) Calculate ΔH_{rxn} for the following reaction: C_3H_8 (g) + 5 O_2 (g) → 3 CO_2 (g) + 4 H_2O (l). (b) Is the reaction endothermic or exothermic? (c) Would you expect the reaction to be product favored?
12. When a reaction, A → B, is written in the reverse direction, B → A, what happens to $\Delta H°_{rxn}$?
13. A cow can burp 280L of methane per day at STP (22.4L=1mol). How much heat could be produced by collecting and completely burning the methane (CH_4) collected in 1 day from 1 cow? $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H°_{rxn}$ = -802.3kJ

14. Given that $\Delta H°_{fusion}$ for ice, $H_2O(s)$, is 6kJ/mol, while $\Delta H°_{fusion}$ for solid ethanol is 5kJ/mol. (a) How much heat is required to melt 1 mol of each substance? (b) Which substance requires more heat to melt? (c) Is the process exothermic or endothermic?
15. Given that for sodium metal, $\Delta H°_{fus}$ is 2.60 kJ/mol, $\Delta H°_{vap}$ is 89.04 kJ/mol, Cp(Na(s)) =1.228J/gK, Cp(Na(l))=1.423J/gK, and that Cp(Na(g))=0.904J/gK, calculate the heat in J required to raise the temperature of 25.0g of sodium metal from a solid at 25°C to a liquid at 98°C. The freezing point of sodium is 98°C. The boiling point of sodium metal is 883°C.
16. Given the following equations:

    $2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$ $\Delta H_{rxn}° = -571.6$ kJ

    $C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$ $\Delta H_{rxn}° = -393.5$ kJ

    $CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2(g) + 2\ H_2O\ (l)$ $\Delta H_{rxn}° = -890.4$ kJ

    Determine the enthalpy change ($\Delta_r H°$) in kJ for the following reaction:

    $C\ (s) + 2\ H_2\ (g) \rightarrow CH_4\ (g)$
17. Given the following thermodynamic data, calculate the enthalpy of the balanced chemical reaction (the one shown below is not balanced) at 25°C?

    $\Delta_f H°$ (kJ/mol): $CH_3CO_2H(\ell)$= −484.0; $CO_2(g)$= −393.5, $H_2O(g)$= −241.8

    $CH_3CO_2H(\ell) + O_2\ (g) \rightarrow CO_2(g) + H_2O(g)$
18. Which of these cations would be predicted to have the most exothermic hydration energy: $Cs^+(g)$, $Sr^{2+}(g)$, $Al^{3+}(g)$
19. Which of these salts would have the greatest lattice energy: CaF_2, $CaCl_2$, KF, KCl
20. Which substance in #19 has the highest vapor pressure?
21. Which substance in #19 has the highest ΔH_{vap}?
22. Which substance as the highest melting point: HF, O_2, CO, CH_4
23. Which substance in #22 has the lowest boiling point?
24. Which pure substance has the weakest intermolecular forces? N_2, H_2O, H_2, HCl, NaCl, CH_4
25. Rank the compounds by in #24 by increasing vapor pressure (lowest vapor pressure to highest)
26. Arrange these compounds by increasing molecular polarity (lowest molecular polarity to highest): HCl, HI, HF
27. Arrange these compounds by increasing boiling point (lowest to highest): CO_2, NH_3, O_2
28. The slope of the line obtained by plotting the natural logarithm of the vapor pressure of a gas against the reciprocal of its Kelvin temperature is?
29. What does ΔH_{vap} tell us about intermolecular forces?
30. Which of the substances in #22 and #24 can hydrogen bond?
31. Which of the substances in #22 and #24 have London forces (induced-dipole induced-dipole force) as the predominant intermolecular force?
32. Which of the substances in #22 and #24 have the dipole-dipole force as the predominant intermolecular force?
33. Which of the substances in #22 and #24 have the ion-ion force as the predominant intermolecular force?
34. Salts that are 2:1 or 1:2 cation:anion could adopt which cubic crystal structure?

35. This property explains why some insects can walk on water
36. This property explains why water can get to the tops of tall trees
37. A body-centered cubic unit cell contains how many atoms in the unit cell
38. Where are the atoms located for a simple cubic unit cell?
39. What fraction of the atoms in 38 are located inside the unit cell

1) the substance with the largest specific heat capacity. 2) thermodynamics can tell us if a chemical reaction would be expected to form products if given enough time and energy. 3) (a) +394kJ, (b) -197kJ, (c) the reaction as written is $\Delta H°_f$ for $CO_2(g)$. 4) 11600J, 5) 2503J, 6) 77.5°C - see problem 5, this is a good way to practice, 7) 76.4°C - see problem 6, 8) 1.051J/g °C, 9) Using q(Ir) + q(water) + q*cal* = 0, q*cal* = 313J. 10) -787kJ, q*rxn*+q*bomb*=0, q*rxn*+ 7.45kJ/°C*(28.63 °C - 23.19°C)=0, 11) (a) -2220kJ, (b) exothermic, (c) yes because product-favored reactions tend to be exothermic. 12) $\Delta H^o_{reaction}$ (B→A) = $-\Delta H^o_{reaction}$ (A→B) - the sign is the negative of what it was before. 13) 10MJ, 14) (a) 6kJ for water and 5kJ for ethanol, (b) water takes more heat to melt, (c) melting is endothermic, q>0. 15) 5070J, 16) -74.9kJ, 17) -786.7kJ, 18) Al^{3+}, 19) CaF_2, 20) KCl, 21) CaF_2, 22) HF, 23) CH_4, 24) H_2, 25) NaCl, H_2O, HCl, N_2, CH_4, H_2, 26) HI, HCl, HF 27) O_2, CO_2, NH_3, 28) $-\Delta H_{vap}/R$, 29) A higher ΔH_{vap} indicates stronger intermolecular forces, 30) HF, H_2O, 31) H_2, N_2, O_2, CH_4, 32) HCl, 33) NaCl, 34) face-centered cubic, 35) surface tension, 36) capillary action, 37) 2, 38) the corners of the cube, 39) 1/8th