Name: ______________________________

Heat – Unit Practice Problems

Heat and Temperature

1. Pretend you are doing a scientific study on Earth.
   a. Name three things in the system you are studying.
   b. What is the surroundings?
   c. Explain if Earth is an open system, a closed system or an isolated system.
2. Distinguish between kinetic and potential energy in the following examples: two separated magnets; an avalanche of snow; temperature, books on library shelves; a mountain stream; a stock car race; separation of charge in a battery, chemical bonds.
3. A piece of steel is heated up, then submerged in an isolated system filled with water. The water heats up as the steel cools. Explain how this models the conservation of energy.
4. How are the molecules different in a solid, a liquid and a gas?
5. How does heat depend on temperature?

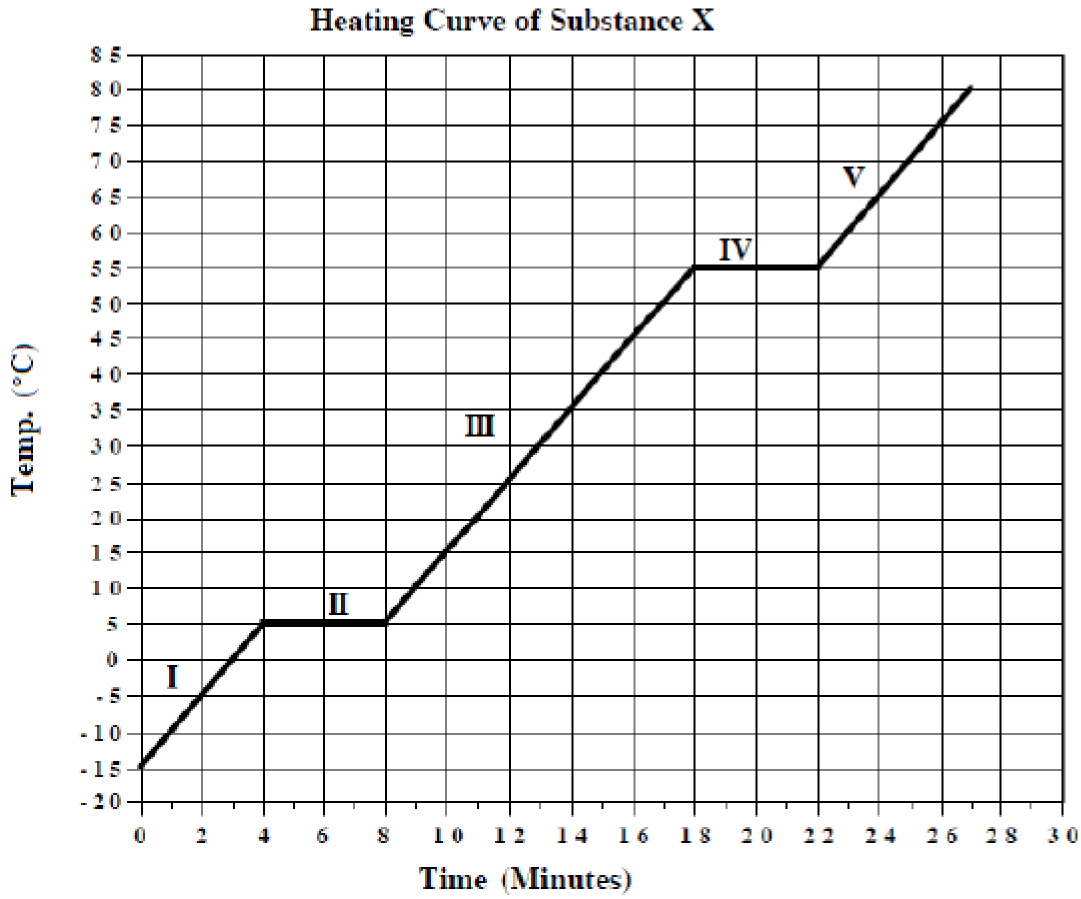
Calorimetry and Specific Heat Capacity

6. Could another liquid be used as effectively as water in a calorimeter? Why or why not?
7. You have two lawn chairs, one made of iron and one made of aluminum. Both are painted the same colour. On a sunny day, which would be hotter to sit in, assuming they have both been in the sun for the same amount of time? Why?
8. Why is the climate more constant (not as hot in the summer and not as cold in the winter) for places near the ocean?
9. The temperature of 335 g of water changed from 24.5°C to 26.4°C. How much heat did this sample absorb?
10. How much heat, in kilojoules, has to be removed from 225 g of ethanol to lower its temperature from 25.0°C to 10.0°C?
11. Heat capacity is the amount of energy needed to heat up an object by one degree (mass does not matter). A calorimeter has a heat capacity of 1265 J/°C. A reaction causes the temperature of the calorimeter to change from 22.34°C to 25.12°C. How much energy does the calorimeter absorb?
12. What is the specific heat of silicon if it takes 192 J to raise the temperature of 45.0 g of Si by 6.0°C?
13. Assuming that Coca Cola has the same specific heat as water, calculate the amount of heat in kJ transferred when one can (about 355 g) is cooled from 25°C to 3°C.
14. An insulated cup contains 75.0 g of water at 24.00°C. A 26.00 g sample of metal at 82.25°C is added. The final temperature of the water and metal is 28.34°C. What is the specific heat of the metal?
15. A 2.88 g piece of silver at 522.0°C is dropped into a calorimeter of water. The initial water temperature is 15.2°C. The final temperature of both the water and silver is 18.1°C. What volume of water is in the calorimeter?
16. 52.8 mL of ethanol with an initial temperature of 62.4°C is poured into 109.2 mL of water with an initial temperature of 19.9°C. What is the final temperature of the liquid mixture?

Phase Changes

17. What is happening to the molecules in a substance when the state is changing?
18. What is happening to the molecules in a substance when the temperature is changing?

19. How can you identify the melting and boiling points on a heating curve?
20. Use the following heating curve to answer the questions:



a. In what part of the curve is the substance a solid?
b. In what part of the curve is the substance a gas?
c. In what part of the curve is the substance both a solid and a liquid?
d. What is the melting point of the substance?
e. What is the boiling point of the substance?
f. In what part(s) of the curve would kinetic energy be increasing?
g. In what part(s) of the curve would potential energy be changing?
h. At what temperature are the molecules moving the slowest?
i. At what temperature would the molecules be furthest apart?

21. Your friend asks you to boil a pot of water, and to leave it to boil for five minutes so it will be "really hot". What is wrong with this statement?
22. 28. Sketch a cooling curve for ethanol as it is cooled from 90°C to -120°. Ethanol condenses at 78°C and freezes at -114°C.

Latent Heat

23. A substance has a heat of fusion of 100 J/g and a heat of vaporization of 1000 J/g. Which takes more energy: melting or vaporization? Explain why, using what you know about the molecules in solids, liquids and gases.
24. How much energy does it take to boil 100.0 mL of ethanol? (ρ = 0.789 g/mL, L_v = 838.3 J/g)
25. How much energy does it take to melt 1.0 g of solid gold? (L_f = 67 J/g)

26. How much energy is required to convert 250.0 g of ice at 0°C to steam at 100°C? (L_f = 333.7 J/g, L_v = 2257 J/kg, c = 4.184 J/g·°C)
27. If you are melting copper (L_f = 207 J/g) and iron (209 J/g) if all energy input is constant, which will melt fastest?
28. You want to boil 250.0 mL of water at 20°C in your 900 W microwave. (L_v = 2257 J/g, ρ = 1.00 g/mL, c = 4.184 J/g·°C, boiling point = 100°C)
    a. How much energy will this take?
    b. How long will it take on full power?
29. You want to melt 454 g of butter in your 700 W microwave. It starts at room temperature, 20°C. How much time do you need to microwave it on full power? (L_f = 53 J/g, boiling point = 32°C, c = 1.04 J/g·°C)
30. You put a 1.0 kg ham that is 20°C in your freezer, which is set at -4°C. Ham freezes at -1.7°C. How much energy is released as the ham freezes? The specific heat for ham above freezing is 2.72 J/g·°C, specific heat below freezing is 1.55 J/g·°C and L_f = 187 J/g.

Thermochemical Equations and Enthalpy

31. What information needs to be in a thermochemical equation?
32. The complete combustion of acetic acid (CH_3COOH) releases 871 kJ of heat per mole of acid. Write a thermochemical equation for the reaction.
33. Exactly 332 kJ of heat is required for the decomposition of aluminum oxide into its elements. Write a thermochemical equation for the reaction.
34. When magnesium metal combines with oxygen from the air, 1204 kJ of heat energy is released.
    a. Write a thermochemical equation for the reaction.
    b. Calculate the amount of heat transferred if 1.3 g of magnesium reacts.
    c. What mass of magnesium oxide is produced during an enthalpy change of -96 kJ?
35. The decomposition of baking soda is represented by the following thermochemical equation:

$$2\ NaHCO_3(s) + 129\ kJ \rightarrow Na_2CO_3(s) + H_2O\ (l) + CO_2(g)$$

    a. What is the value for ΔH?
    b. How much heat is needed to decompose 10.0 g of baking soda?

Heat of Solution

36. When solid sodium hydroxide is dissolved into water, forming aqueous sodium and hydroxide ions, 445 kJ/mol of energy is released.
    a. Write the thermochemical equation for this process. (Hint: water is not included in the equation.)
    b. How much heat, in kJ, is released when 30 g of sodium hydroxide is dissolved?
37. 10.2 g of potassium chlorate ($KClO_3$) is dissolved in 65.0 mL of water.
    a. Write the thermochemical equation for this process.
    b. The amount of heat produced or absorbed by the reaction
    c. The temperature change of the water
38. You dissolve 0.55 g of potassium hydroxide in 75.0 mL of water. What will the temperature change of the solution be? (Assume the mass of the solid is negligible.)
39. You want to make a hot pack using calcium chloride (ΔH = -82.9 kJ/mol). It needs to have 250.0 mL of water and should go from room temperature (20°C) to 40°C. What mass of calcium chloride should you use? (Assume the mass of solid is negligible in the mass of solution.)

40. A student uses an aluminum calorimeter to determine the enthalpy of solution for solid ammonium nitrate. The student assumes that the calorimeter neither gains nor loses heat during the experiment, that the density and specific heat capacity of the final solution are the same as water and that the final mass of the solution is 150.00 g. The data were collected and recorded as follows:

Mass of aluminum calorimeter	25.45 g
Mass of aluminum calorimeter and contents	175.45 g
Mass of ammonium nitrate	1.68 g
Initial temperature of calorimeter and contents	22.30°C
Final temperature of calorimeter and contents	20.98°C

a. What is the experimental value of ΔH_{sol} for ammonium nitrate?
b. What are some sources of error in this experiment?