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Chapter 10: Basic Review Worksheet

1. How is the concept of *energy* defined?

2. As a ball rolls down a hill _______ energy decreases, ________ energy increases, and _______ energy remains the same.

3. What is meant by the term *state function*? Provide an example.

4. What does *temperature* measure?

5. Explain what is meant by the terms *exothermic* and *endothermic*.

6. Define the terms *system* and *surroundings*.

7. What are some common *units* of energy and how are these units defined?

8. For each of the following amounts of energy, perform the indicated conversion of units.
   a. 459 J to calories
   b. 55.31 kJ to joules
   c. 84.1 kJ to kilocalories

9. Calculate ΔE for each of the following cases.
   a. $q = 34$ J, $w = -22$ J
   b. $q = -28$ J, $w = -23$ cal
   c. $q = -15$ J, $w = +12$ J

10. For an endothermic process, q is reported with a _________ sign, and for an exothermic process, q is reported with a ________ sign.

11. What is meant by the *specific heat capacity* of a material?

12. Calculate the *mass* (in grams) of each of the following substances that could be warmed over the indicated temperature range by application of exactly 1.0 kJ of energy.
    a. water, from 15°C to 42°C
    b. iron, from 25°C to 125°C
    c. carbon, from -10°C to 47°C

13. What is meant by the term *enthalpy*?

14. Why does the quality of energy decrease when we use it?

15. What is the driving force for every exothermic process?

16. Consider the combustion of propane

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) \qquad \Delta H = -2221\ \text{kJ}$$

a. Is the combustion of propane endothermic or exothermic?
b. Is energy as heat released into the surroundings or absorbed by the system?
c. How much energy as heat is released when 2 mol of propane are burned in excess oxygen?

17. Given the following data

$4CuO(s) \rightarrow 2Cu_2O(s) + O_2(g)$ $\Delta H = 288$ kJ
$Cu_2O(s) \rightarrow Cu(s) + CuO(g)$ $\Delta H = 11$ kJ

Determine ΔH for the reaction $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

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Chapter 10: Review Worksheet

1. There is no heat in an insulated system at a constant temperature of 400.°C. Explain.

2. Explain how Figure 10.1 shows the law of conservation of energy.

3. For each of the following amounts of energy, perform the indicated conversion of units.
   - b. 7031 cal to kilojoules
   - d. 78.3 kcal to kilojoules
   - e. 4541 cal to kilocalories

4. Energy is a state function. Why isn't heat a state function?

5. At the end of Section 10.3, an exothermic reaction is explained. Write an analogous explanation of an endothermic reaction.

6. You start a fire in a fireplace by striking a match and lighting crumpled paper under some logs. Explain all of the energy transitions in this scenario using the terms *exothermic*, *endothermic*, *system*, *surroundings*, *potential energy*, and *kinetic energy* in the discussion.

7. How is specific heat capacity used to calculate the energy change when a substance is heated?

8. Calculate the *mass* (in grams) of each of the following substances that could be warmed over the indicated temperature range by application of exactly 1.0 kJ of energy.
   - d. gold, from 56°F to 74°F
   - e. silver, from 289 K to 385 K
   - f. aluminum, from -10°C to 85.0°F

9. Calculate ΔE for each of the following cases.
   - a. A system releases 23 J of heat while 12 J of work is done on it.
   - b. 14 J of work is done on a system and 5.0 calories of heat are released.
   - c. A system absorbs 87 J of heat and performs 32 J of work.

10. Consider the combustion of propane

$$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2(g) + 4H_2O(l) \qquad \Delta H = -2221\text{ kJ}$$

   - a. How much energy as heat is released when 100.0 g of propane is burned in excess oxygen?
   - b. How much energy as heat is released when 100.0 g of carbon dioxide is produced?

11. Use the following data to calculate the value of ΔH for the reaction

$$I_2(s) + Cl_2(g) \rightarrow 2ICl(g)$$

	ΔH (kJ/mol)
$Cl_2(g) \rightarrow 2Cl(g)$	242.3
$I_2(g) \rightarrow 2I(g)$	151.0
$ICl(g) \rightarrow I(g) + Cl(g)$	211.3
$I_2(s) \rightarrow I_2(g)$	62.8

12. Is enthalpy a state function? Explain your answer.

13. Explain how we can have an “energy crisis” given that the law of conservation of energy is true.

14. What acts as a driving force for every endothermic process that occurs?

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Chapter 10: Challenge Review Worksheet

1. According to the text, increasing entropy is a driving force, and steam has higher entropy than liquid water. So why doesn't all of the liquid water on earth spontaneously turn to steam?
2. You heat a 5.00 g sample of iron in a boiling water bath. You then quickly and carefully place the iron into 125.0 mL of water at 25.0°C in a coffee cup calorimeter. Determine the final temperature of the water and explain all assumptions.
3. Explain why the first law of thermodynamics cannot be used to explain why a ball cannot spontaneously roll up a hill, but why the second law can be used.

For aluminum: $2\,Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

Molar masses: Al, 26.98 g; Cl_2 70.90 g; $AlCl_3$, 133.3 g

$$25.0\text{ gAl} \times \frac{1\text{ mol}}{26.98\text{ g}} = 0.927\text{ mol Al}$$

$$50.0\text{ g } Cl_2 \times \frac{1\text{ mol}}{70.90\text{ g}} = 0.705\text{ mol } Cl_2$$

Cl_2 is the limiting reactant

$$0.705\text{ mol } Cl_2 \times \frac{2\text{ mol } AlCl_3}{3\text{ mol } Cl_2} \times \frac{133.3\text{ g}}{1\text{ mol}} = 62.7\text{ g } AlCl_3$$

4. (31.2 g KCl / 47.6 g KCl) x 100% = 65.6% yield

5. The relevant reaction is $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

$$1.48\text{ AgCl} \times \frac{1\text{ mol AgCl}}{143.35\text{ g AgCl}} \times \frac{1\text{ mol } Ag^+}{1\text{ mol AgCl}} = 0.0103\text{ mol } Ag^+$$

The 0.0103 mol Ag^+ comes from $AgNO_3$, thus we started with 0.0103 mol $AgNO_3$ in the mixture.

$$0.0103\text{ mol } AgNO_3 \times \frac{169.91\text{g}}{1\text{ mol } AgNO_3} = 1.75\text{g } AgNO_3$$

$\frac{1.75\text{g } AgNO_3}{5.00\text{g mixture}} \times 100\% = 35.0\%$. The mixture is 35.0% $AgNO_3$

Chapter 10: Basic Review Worksheet

1. Scientists define energy as "the capacity to do work or to produce heat". As with "matter", energy is such a fundamental concept that it is hard to define.
2. potential; kinetic; total
3. A state function is a property that is independent of pathway. Energy and elevation are examples of state functions. Heat and work are not state functions.
4. Temperature is a measure of the average kinetic energy of the particles of a system.
5. An exothermic reaction or process is one in which energy as heat is released to the surroundings; an endothermic reaction or process is one the system absorbs energy as heat from the surroundings.

6. The system is the part of the universe on which we focus attention; the surroundings are the rest of the universe.

7. Although the SI unit of energy is the joule, until relatively recently, energies were more commonly given in terms of the calorie: one calorie is defined to be the amount of heat required to raise the temperature of one gram of water by one Celsius degree. The calorie is a "working" definition, and we can more easily appreciate this amount of energy. In terms of the SI unit, 1 calorie = 4.184 joule, so it takes 4.184 J to raise the temperature of one gram of water by one Celsius degree.

8. a. $459\ \text{J} \times \dfrac{1\ \text{cal}}{4.184\ \text{J}} = 109.7 = 110.\ \text{cal}$
   b. $55.31\ \text{kJ} = 55{,}310\ \text{J} = 5.531 \times 10^4\ \text{J}$
   c. $84.1\ \text{kJ} \times \dfrac{1\ \text{kcal}}{4.184\ \text{kJ}} = 20.1\ \text{kcal}$

9. $\Delta E = q + w$
   a. 34 J + (-22 J) = 12 J
   b. (-28 J) + (-23 cal) = (-28 J) + (-96 J) = -124 J
   c. (-15 J) + 12 J = -3 J

10. positive; negative

11. The specific heat capacity of a substance is the amount of energy required to raise the temperature of one gram of a substance by one Celsius degree. Therefore, the specific heat capacity of water must be 1.000 cal/g°C or 4.184 J/g°C.

12. a. temperature change = 27°C
    mass = 1.0×10^3 J/(4.184 J/g°C) (27°C) = 8.9 g
    b. temperature change = 100.°C
    mass = 1.0×10^3 J/(0.45 J/g°C) (100.°C) = 22 g
    c. temperature change = 57°C
    mass = 1.0×10^3 J/(0.71 J/g°C) (57°C) = 25 g

13. Enthalpy (H) has a mathematical definition: H = E + PV. However, at constant pressure, it turns out that the change in enthalpy is equal to the energy flow as heat.

14. The quality of energy is decreasing because of entropy as a driving force. For example, the amount of energy before and after a car consumes a gallon of gasoline is the same. However, the energy is more "spread out" once it has been burned, and is less able to perform useful work. Thus, the quantity is constant, but the quality is always decreasing.

15. The driving force is entropy, or more specifically, "energy spread". That is, for an exothermic reaction, energy is spread from more concentrated (in the system) to less concentrated (in the surroundings).

16. a. Since ΔH is given with a negative sign, the reaction is exothermic.
    b. Energy as heat is released into the surroundings.
    c. 2(-2221 J) = -4442 J. Thus, 4442 J of energy as heat is released when 2 mol of propane is burned in excess oxygen.

17. Reverse the two given reactions, multiply the second by 2, and add:

$$\begin{array}{ll} 2Cu_2O(s) + O_2(g) \rightarrow 4CuO(s) & \Delta H = -288\ kJ \\ \underline{2Cu(s) + 2CuO(g) \rightarrow 2Cu_2O(s)} & \underline{\Delta H = \ -22\ kJ} \\ 2Cu(s) + O_2(g) \rightarrow 2CuO(s) & \Delta H = -310.\ kJ \end{array}$$

Chapter 10: Review Worksheet

1. Heat is defined as a flow of energy due to a temperature difference. In a insulated system at constant temperature there is no temperature difference, thus no energy flow as heat. This can confuse students because 400°C seems "hot" and this equates to thinking of the system as "containing heat". However, heat is not a substance. If we were to stick our hands in such a system, there would be a difference in temperature between the system and our hand, and there would be energy flow as heat.

2. In Figure 10.1a, ball A has greater potential energy than ball B. As ball A rolls down the hill, the potential energy decreases, but its kinetic energy increases. When ball A hits ball B, the kinetic energy of A is transferred to ball B. The potential energy of ball B is increased. Assuming that all of the energy was transferred from ball A to ball B (no frictional heating for example) the potential energy of ball B in Figure 10.1b is the same as the potential energy of ball A in Figure 10.1a.

3. a. $7{,}031\ \text{cal} \times \dfrac{4.184\ \text{J}}{1\ \text{cal}} \times \dfrac{1\ \text{kJ}}{1000\ \text{J}} = 29.42\ \text{kJ}$
   b. $78.3\ \text{kcal} \times \dfrac{4.184\ \text{J}}{1\ \text{cal}} = 327.6 = 328\ \text{kJ}$
   c. 4,541 cal = 4.541 kcal

4. Heat is only one way that energy is transferred. For example, if a ball rolls down a hill, the amount of heat involved literally depends on the pathway (for example, if the pathway is rough, there is more friction, thus more heat involved). However, the work the ball can do once it rolls down the hill is dependent on the pathway as well. These factors are opposite of one another and thus energy is conserved even though heat and work vary.

5. In any endothermic reaction, some of the thermal energy (random kinetic energy) via heat is converted to potential energy stored in the chemical bonds.

6. Answers will vary. However, crumpling paper and striking a match are endothermic (although the burning of a match is overall an exothermic process). The paper and logs are the system, while the room makes up the surroundings. Potential energy is stored in the match, the paper, and the logs. Kinetic energy is released as thermal energy via heat.

7. To see how specific heat capacities may be used to calculate the energy change for a process, consider this example: "How much energy is required to warm 25.0 g of water from 15.1°C to

35.2°C?" The specific heat capacity of water is 1.000 cal/g°C. This is the quantity of energy required to raise the temperature of only one gram of water by only one Gelsius degree. In this example, we are raising the temperature by (35.2 – 15.1) = 20.1°C. So, using the specific heat capacity as a conversion factor, we can say

$$\text{energy required} = \left(\frac{1.00\ \text{cal}}{\text{g}\ ^\circ\text{C}}\right) \times (25.0\ \text{g}) \times (20.1^\circ\text{C}) = 503\ \text{cal}$$

Notice how the units of g and °C cancel, leaving the answer in energy units only. This sort of calculation of energy change can be done for any substance, using the substance's own specific heat capacity.

8. a. 56°F = 13.3°C; 75 °F = 23.9°C
   temperature change = 24 -13 = 11°C
   mass = 1.0×10^3 J/(0.13 J/g°C) (11°C) = 7.3×10^2 g
   b. temperature change = 385K –289K = 96K =96°C
   mass = 1.0×10^3 J/[(0.24 J/g°C) (96°C)] = 43 g
   c. 85.0°F = 29.4°C
   temperature change = 29.4°C – (-10.0°C) = 39.4°C
   mass = 1.0×10^3 J/[(0.89 J/g°C) (39.4°C)] = 29 g

9. E = q + w
   a. (-23 J) + 12 J = -11 J
   b. 14 J + (-5 cal) = 14 J + (-21 J) = -7 J
   c. 87 J + (-32 J) = 55 J

10. a. $100.0\text{g } C_3H_8 \times \frac{1\ \text{mol}\ C_3H_8}{44.094\text{g}\ C_3H_8} \times \frac{-2221\ \text{kJ}}{1\ \text{mol}\ C_3H_8} = -5037\ \text{kJ}$

    b. $100.0\text{g } CO_2 \times \frac{1\ \text{mol}\ CO_2}{44.01\text{g}\ CO_2} \times \frac{-2221\ \text{kJ}}{3\ \text{mol}\ CO_2} = -1682\ \text{kJ}$

11.

	ΔH (kJ/mol)
$I_2(s) \rightarrow I_2(g)$	62.8
$I_2(g) \rightarrow 2I(g)$	151.0
$Cl_2(g) \rightarrow 2Cl(g)$	242.3
$2I(g) + 2Cl(g) \rightarrow 2ICl(g)$	-2(211.3)
$I_2(s) + Cl_2(g) \rightarrow 2ICl(g)$	33.5

12. Yes, enthalpy is a state function. The definition of enthalpy (H) is H = E + PV. Since E, P, and V are all state functions, H is a state function as well.

13. The energy crisis has to do with the fact that the quality of the energy is decreasing (due to entropy, or more specifically "energy spread"). The amount of energy is constant, but the amount of useful work that can be done with this energy is decreasing.

14. One driving force is "energy spread", which occurs in exothermic processes as energy goes from more to less concentrated. However, if an endothermic process occurs, this must be because of "matter spread" (for example, when a gas is produced from a solid, matter is "spread").

Chapter 10: Challenge Review Worksheet

1. We have to consider both "matter spread" and "energy spread" as driving forces. While water turning to steam is favored due to matter spread, it is disfavored by energy spread. This is because energy must flow via heat into water in order to make the conversion to steam. This is one example of many processes in which the two driving forces are counter to each other, therefore making spontaneity temperature dependent.

2. We will assume the following:
    a. All of the heat lost by the iron is gained by the water.
    b. The iron has an initial temperature of 100.0°C.
    c. The density of water is 1.00 g 1mL over the temperature.
    d. The iron and water achieve the same final temperature.

    To make both signs positive, we will define the change in temperature of the iron as 100.0°C - t_f and the change in temperature of the water as t_f - 25.0°C.

    heat lost by iron = heat gained by water
    (5.00) (0.45 J/g°C) (100.0 - t_f) = (125.0g) (4.184J/g°C) (t_f - 25.0)
    225 - 2.25t_f = 523 t_f – 13075
    13300 = 525.25 t_f
    25.3°C = t_f

3. The first law of thermodynamics tells us that energy is conserved. Thus, when a ball rolls down a hill, potential energy is converted to kinetic energy due to the motion of the ball, and to frictional heating of the ball and hill. There is nothing in the first law, however, that states that the energy that has been "spread" cannot "re-concentrate" to the ball and cause the ball to roll back up the hill. The second law of thermodynamics tells us that entropy (or disorder) is constantly increasing. Thus, energy spread as the ball rolls down the hill is a driving force, and energy will not spontaneously "re-concentrate".

Chapter 11: Basic Review Worksheet

1. Electromagnetic radiation represents the propagation of energy through space in the form of waves.

2. A representative wave is depicted in Figure 11.3 in the text.

3. An atom is said to be in its ground state when it is in its lowest possible energy state. When an atom possesses more energy than in its ground state, the atom is said to be in an excited state.

4. Photons are discrete quantities of radiation. Atoms do not gain or emit radiation randomly, but rather do so only in discrete bundles of radiation called photons.