Chapter 5
Thermochemistry

1. The ΔE of a system that releases 14.4 J of heat and does 4.8 J of work on the surroundings is __________J.

   (a). 19.2 J
   (b). 14.4 J
   (c). 4.8 J
   (d).- 19.2 J

   Explanation: The ΔE can be calculated by using the equation $\Delta E = q + w$. Since the system is releasing heat and is doing work on the surroundings both of these numbers will have negative signs. Here the system is *losing* its internal energy. Thus $\Delta E = -14.4 + (-4.8) = -19.2$ J

2. The value of ΔE for a system that performs 23 kJ of work on its surroundings and loses 79 kJ of heat is __________kJ.

   (a). +23 kJ
   (b).-102 kJ
   (c). +102 kJ
   (d). -79 kJ

   Explanation: The ΔE can be calculated by using the equation $\Delta E = q + w$. Since the system is losing heat and is doing work on the surroundings both of these numbers will have negative signs. Here the system is losing its internal energy. Thus $\Delta E = -23 + (-79) = -102$ kJ

3. Calculate the value of ΔE in joules for a system that loses 50 J of heat and has 150 J of work performed on it by the surroundings.

   (a). +50 J
   (b). -100 J
   (c). +100 J
   (d). -200 J

   Explanation: The ΔE can be calculated by using the equation $\Delta E = q + w$. Since the system is releasing heat the numerical value of the heat will be negative. At the same time the surroundings are doing work on the system and the work will have a positive sign. Here the system is *gaining* internal energy. Thus $\Delta E = +150 \text{ J} + (-50 \text{ J}) = +100$ J

4. The change in the internal energy of a system that releases 2,500 J of heat and that does 7,655 J of work on the surroundings is ___________J.

   (a). +10,155 J
   (b). -5,155 J
   (c). -7655 J
   (d). -10,155 J

   Explanation: The ΔE can be calculated by using the equation $\Delta E = q + w$. Since the system is releasing heat and is doing work on the surroundings both of these numbers will have negative signs. Here the system is *losing* its internal energy. Thus $\Delta E = -2500\text{ J} + (-7655\text{ J}) = -10155\text{ J}$

5. The value of ΔH^o for the reaction below is -72 kJ. How many kJ of heat will be released when 1.0 mol of HBr is formed in this reaction?

$$H_2\,(g) + Br_2\,(g) \longrightarrow 2HBr\,(g)$$

   (a). +144 kJ
   (b). +72 kJ
   (c). +36 kJ
   (d). -36 kJ

   Explanation: The ΔH^o for the reaction indicates the amount of heat released by the reaction where 2 moles of HBr (from the balanced equation) are formed. If only one mole of HBr is formed then the reaction should release only 36 kJ (72/2) of heat.

6. The value of ΔH^o for the reaction below is -126 kJ. How many kJ of heat will be released when 2.00 mol of NaOH is formed in the reaction?

$$2Na_2O_2\,(s) + 2H_2O\,(l) \longrightarrow 4NaOH\,(s) + O_2\,(g)$$

   (a). +252 kJ
   (b). +63 kJ
   (c). +3.9 kJ
   (d). +7.8 kJ

   Explanation: The ΔH^o for the reaction indicates the amount of heat released by the reaction where 4 moles of NaOH (from the balanced equation) are formed. If only two moles of NaOH are formed then the reaction should release only 63 kJ (126/2) of heat.

7. The value of ΔH^o for the reaction below is -790 kJ. Calculate the enthalpy change accompanying the complete reaction of 0.75 g of S

$$2S\ (s) + 3O_2\ (g) \longrightarrow 2SO_3\ (g)$$

(a). +23 kJ
(b). -9.2 kJ
(c). -18 kJ
(d). +12 kJ

Explanation: The balanced equation shows a reaction in which 2 moles of S are reacting. This reaction has a value of ΔH^o = -790 kJ. Need to convert the grams of sulfur to moles of sulfur and then calculate the value of ΔH^o for that number of moles of sulfur. Remember to factor in the **2 moles** that are reacting in the equation. Failure to do this will give you an incorrect answer of -18 kJ.

$$0.75\text{ g S} \times \frac{1\text{ mole S}}{32.065\text{ g S}} \times \frac{-790\text{ kJ}}{2\text{ moles S}} = -9.2\text{ kJ}$$

8. The value of ΔH^o for the reaction below is -482 kJ. Calculate the heat (kJ) released to the surroundings when 12.0 g of CO (g) reacts completely.

$$2CO\ (g) + O_2\ (g) \longrightarrow 2CO_2\ (g)$$

(a). -2.89×10^3 kJ
(b). -207 kJ
(c). -103 kJ
(d). -65.7 kJ

Explanation: The balanced equation involves 2 moles of CO reacting which releases 482 kJ. To calculate the amount of heat released by the reaction of 12.0 g of CO, need to convert the grams of CO to moles of CO and then calculate the amount of heat released. Do not forget to factor in the **2 moles** not doing this will give you the incorrect answer of -207 kJ.

$$12.0\text{ g CO} \times \frac{1\text{ mole CO}}{28.01\text{ g}} \times \frac{-482\text{ kJ}}{2\text{ moles CO}} = -103\text{ kJ}$$

9. The value of ΔH^o for the reaction below is -336 kJ. Calculate the heat (kJ) released to the surroundings when 23.0 g of HC1 is formed.

$$CH_4\ (g) + 3C1_2\ (g) \longrightarrow CHC1_3\ (l) + 3HCl\ (g)$$

(a). 177
(b). 2.57×10^3
(c). 70.7
(d). 211

Explanation: The balanced equation involves the formation of 3 moles of HCl, which releases 336 kJ. The grams of HCl need to be converted to moles and then the amount of heat released by this reaction can be calculated. Do not forget to factor in the **3 moles** not doing this will give you the incorrect answer of -211 kJ.

$$23.0\text{ g HCl} \times \frac{1\text{ mole HCl}}{36.458\text{ g}} \times \frac{-336\text{ kJ}}{3\text{ mole HCl}} = -70.7\text{ kJ}$$

10. The value of ΔH^{o} for the reaction below is +128.1 kJ:

$$CH_3OH\ (l) \longrightarrow CO\ (g) + 2H_2\ (g)$$

How many kJ of heat are consumed when 5.10 g of H_2 (g) is formed as shown in the equation?

(a). 653
(b). 62.0
(c). 324
(d). 162

Explanation: The balanced equation involves the formation of 2 moles of H_2, which consumes 128.1 kJ. Covert the grams of H_2 to moles of H_2 and then calculate the amount of heat consumed in this reaction. Do not forget to factor in the **2 moles** of H_2, not doing this will give you the incorrect answer of 324 kJ

$$5.10\text{ g } H_2 \times \frac{1\text{ mole } H_2}{2.016\text{ g}} \times \frac{+128.1\text{ kJ}}{2\text{ moles } H_2} = +162\text{ kJ}$$

11. The value of ΔH^{o} for the reaction below is +128.1 kJ:

$$CH_3OH\ (l) \longrightarrow CO\ (g) + 2H_2\ (g)$$

How many kJ of heat are consumed when 5.10 g of CO (g) is formed as shown in the equation?

(a). 0.182 kJ
(b). 162 kJ
(c). 8.31 kJ
(d). 23.3 kJ

Explanation: The balanced equation involves the formation of 1 mole of CO, which consumes 128.1 kJ. Covert the grams of CO to moles of CO and then calculate the amount of heat consumed in this reaction.

$$5.10\text{ g CO} \times \frac{1\text{ mole CO}}{28.01\text{ g}} \times \frac{+128.1\text{ kJ}}{1\text{ mole CO}} = 23.3\text{ kJ}$$

12. The value of ΔH^o for the reaction below is -1107 kJ:

$$2Ba\ (s) + O_2\ (g) \longrightarrow 2\ BaO\ (s)$$

How many kJ of heat are released when 5.75 g of BaO (s) is produced?

(a). 56.9 kJ
(b). 23.2 kJ
(c). 20.8 kJ
(d). 193 kJ

Explanation: The balanced equation involves the formation of 2 moles of BaO, which releases 1107 kJ of heat. Convert the grams of BaO produced and then calculate the amount of heat released.

$$5.75\text{ g BaO} \times \frac{1\text{ mole BaO}}{153.326\text{ g}} \times \frac{-1107\text{ kJ}}{2\text{ moles BaO}} = -20.8\text{ kJ}$$

13. The molar heat capacity of a compound with the formula C_2H_6SO is 88.0 J/mol-K. The specific heat capacity of this substance is __________ J/g-K.

(a). 88.0
(b). 1.13
(c). -1.13
(d). 6.88 x 10^3

Explanation: This question is based on the definitions of the 2 quantities involved. The molar heat capacity of any substance is the heat capacity of **one mole** of the substance while the specific heat capacity involves **only 1 gram** of the substance.

$$\text{Specific heat capacity} = \frac{\text{Molar heat capacity}}{\text{Molar mass}} = \frac{88.0\ \text{J}/\text{mol K}}{78.134\ \text{g}/\text{mol}} = 1.13\ \text{J}/\text{g K}$$

14. A sample of aluminum metal absorbs 9.86 J of heat, upon which the temperature of the sample increased from 23.2°C to 30.5°C. Since the specific heat capacity of aluminum is 0.90 J/g-K, the mass of the sample is __________g.

(a). 72
(b). 65
(c). 1.50
(d). 8.1

Explanation: This question is based on the formula used to calculate the amount of heat "q" involved. The formula for this calculation is: q = m x c x ΔT, where m is the mass of the substance, c is its specific heat capacity and ΔT is the change in temperature. In this question the ΔT is a positive number as the T_{final} (30.5 + 273.15

= 308.15 K) is higher than the $T_{initial}$.(23.2 + 273.15 = 296.35 K). It is important that the temperatures are converted to Kelvin and not left in the Celsius scale.

$$m = \frac{q}{c \times \ddot{A}T} = \frac{+9.86\ J}{0.90\ J/g \times 7.30\ K} = 1.50\ g$$

15. The specific heat capacity of lead is 0.13 J/g-K. How much heat (in J) is required to raise the temperature of 150.0 g of lead from 22.01°C to 37.01°C?

(a). 2.0 J
(b). 0.13 J
(c). 5.8 x 10^{-4} J
(d). 292.5 J

Explanation: This question is based on the formula used to calculate the amount of heat "q" needed here. The sample would have to absorb heat to raise its temperature.
q = m x c x ΔT
q = (150 g) x (0.13 J/g K) x (15.00K) = 292.5 J

16. Objects can possess energy as __________.

(a). endothermic energy
(b). potential energy
(c). kinetic energy

(a) a only
(b) a and b
(c) a and c
(d) b and c

Explanation: Only potential and kinetic are types of energy while endothermic relates to a process in which a system absorbs energy.

17. The internal energy of a system is always increased by __________.

(a). adding heat to the system
(b). having the system do work on the surroundings
(c). withdrawing heat from the system
(d). adding heat to the system and having the system do work on the surroundings.

Explanation: Addition of heat is the only process listed here that will result in increasing the internal energy of a system.

18. Which one of the following conditions would always result in an increase in the internal energy of a system?

    (a). The system loses heat and does work on the surroundings.
    (b). The system loses heat and has work done on it by the surroundings
    (c). The system gains heat and has work done on it by the surroundings.
    (d). None of the above is correct.

19. When a system ________ ΔE is <u>always</u> negative.

    (a). absorbs heat and does work
    (b). gives off heat and does work
    (c). absorbs heat and has work done on it
    (d). none of the above is <u>always</u> negative

20. Which one of the following is an exothermic process?

    (a). ice melting
    (b). water evaporating
    (c). boiling soup
    (d). condensation of water vapor

    Explanation: All the processes except condensation of water vapor will require that hat is absorbed. In an exothermic process heat is always <u>released</u>.

21. Of the following, which one is a state function?

    (a). H
    (b). q
    (c). w
    (d). All of the above

22. Which of the following is a statement of the first law of thermodynamics?

    (a). $E_k = \frac{1}{2}mv^2$
    (b). A negative ΔH corresponds to an exothermic process
    (c). $\Delta E = q - w$
    (d). Energy lost by the system must be gained by the surroundings.

    Explanation: The first law of thermodynamics is essentially the law of conservation of energy, more specifically regards a system and surroundings.

23. A__________ ΔH corresponds to an ___________ process.

    (a). negative, endothermic
    (b). negative, exothermic
    (c). positive, exothermic
    (d). zero, exothermic

    Explanation: An exothermic system will lose its enthalpy corresponding to a negative ΔH.

24. The internal energy of a system can be increased by __________.

        (a). transferring heat from the surroundings to the system
        (b). transferring heat from the system to the surroundings
        (c). doing work on the system

    (a). a only
    (b). b only
    (c). c only
    (d). a and c

    Explanation: An increase in the internal energy will be possible only of the system gains energy from the surroundings and has work done on it by the surroundings.

25. ΔH for an endothermic process is _______ while ΔH for an exothermic process is __________.

    (a). zero, positive
    (b). zero, negative
    (c). positive, negative
    (d). negative, positive

    Explanation: In an endothermic process the enthalpy of the system will increase as the system will gain energy while in an endothermic process the system will lose energy resulting in a loss in enthalpy.

26. For a given process at constant pressure, ΔH is negative. This means that the process is __________.

    (a). endothermic
    (b). a static process
    (c). exothermic

    Explanation: An exothermic process will lose energy and hence have a negative value of ΔH.

27. Which one of the following statements is true?

    (a). Enthalpy is an intensive property.
    (b). The enthalpy change for a reaction is independent of the state of the reactants and products.
    (c). Enthalpy is a state function.
    (d). H is the value of q measured under conditions of constant volume.

    Explanation: Enthalpy is dependent on the amount of and state of the reactants and products and is measured at constant pressure, making (a), (b) and (d) false.

28. A chemical reaction that absorbs heat from the surroundings is said to be __________ and has a __________ ΔH at constant pressure.

    (a). endothermic, positive
    (b). endothermic, negative
    (c). exothermic, negative
    (d). exothermic, positive

