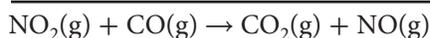
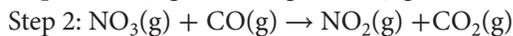


$$\frac{0.050 \frac{\text{mol}}{\text{L}\cdot\text{s}}}{\text{rate of consumption of H}_2\text{O}} = \frac{4}{6}$$

$$\begin{aligned} \text{rate at which H}_2\text{O(g) reacts} &= \frac{6}{4} \times 0.050 \text{ mol/L}\cdot\text{s} \\ &= 7.5 \times 10^{-2} \text{ mol/L}\cdot\text{s} \end{aligned}$$

- 18.** One possible mechanism is shown below.



- 19.** rate of change in $\text{AgNO}_3(\text{aq})$

$$= \frac{\Delta C}{\Delta t} = \frac{0.42 \text{ mol/L} - 0.28 \text{ mol/L}}{60.0 \text{ s} - 30.0 \text{ s}}$$

$$= 4.667 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

The relative rate of change of a reactant or product is in proportion to the mole ratios in the balanced equation.

From the balanced equation, mole ratio $\text{AgNO}_3:\text{Cl}^-$ is 2:2 (or 1:1). Therefore, the rate of reaction of chloride is equal to the rate of reaction of silver nitrate.

$$\text{rate of reaction of Cl}^- = 4.667 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

$$= 4.7 \times 10^{-3} \text{ mol/L}\cdot\text{s}$$

- 20.** The activation energy and the enthalpy change remain the same when the concentrations of the reactants are altered. Having a greater number of particles collide at a higher concentration does not affect the mechanism of the reaction or the nature of the bonds in the reactants. The reaction will proceed more quickly and more thermal energy will be generated more quickly, but the value of ΔH_r for the reaction as written in a thermochemical equation will not change.
- 21.** It is a second-order reaction; for second-order reactions, the rate is proportional to the square of the reactant.
- 22.** The number of effective collisions must be doubled to double the rate of reaction. This can be accomplished by increasing the temperature and/or, if the reactants are gases, by increasing the pressure.
- 23. a.** Because there are three “peaks” representing three transition states, there are three steps involved in this reaction.
- b.** A: the reactants
 B: transition state of Step 1
 C: the products formed in Step 1, intermediates, and any unreacted reactants
 D: the transition state of Step 2
 E: the products formed in Step 2, intermediates, and any unreacted reactants

F: the transition state for Step 3

G: the products

H: the activation energy for the reverse of Step 1

I: the enthalpy change for Step 1

J: the activation energy for Step 2

K: the activation energy for Step 3

L: the enthalpy change for the overall reaction

- c.** A net quantity of energy is added because the products are higher in energy than the reactants, which means the overall reaction is endothermic.
- 24.** Chemists use very short laser pulses to monitor chemical reactions. The durations of these pulses are in the range of femtoseconds. The first pulse provides the activation energy for the reaction. Chemists analyze the absorbance of the second pulse, and use the data to determine the absorbance characteristics of any activated complexes or chemical intermediates. This data allows chemists to identify the chemical species that are formed during a reaction and to provide evidence to support or refute a proposed reaction mechanism.
- 25.** An increase in temperature will increase the rate of a reaction. As temperature increases, more particles will collide with enough energy to react. Student diagrams should resemble Figure 6.15 in the student textbook.

Answers to Unit 3 Review Questions (Student textbook pages 403-9)

1. a
2. d
3. c
4. a
5. c
6. b
7. e
8. d
9. b
10. e

- 11.** Since both water samples are at the same temperature, the average kinetic energy of the particles will be the same. Thermal energy is the sum of the kinetic energies of all the particles in a sample. Since there are more particles in the bathtub of water, the thermal energy will be greater in the bathtub sample.

- 12. a.** Thermal energy is removed from the air (the system).
- b.** The enthalpy change, ΔH_{cond} , is negative.
From Table 5.3, $\Delta H_{\text{vap}} = +40.7 \text{ kJ/mol}$.
Therefore $\Delta H_{\text{cond}} = -40.7 \text{ kJ}$.
- c.** $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\ell) + 40.7 \text{ kJ}$
- 13.** Activation energy is the minimum collision energy between two particles that is required for a reaction to take place between the particles. For a reaction to occur, chemical bonds must be broken. The activation energy is primarily used for this process. The stronger the bonds, the greater the activation energy.
- 14.** The relative amount of energy required to break bonds in the reactants compared to the energy given off when new bonds form in the products determines whether the reaction is exothermic or endothermic. There is no relationship between this and the activation energy.
- 15. a.** three steps
- b.** Step 3
- c.** C to E
- d.** B, C, F
- 16. a.** To determine the average rate of reaction, one of the following can be determined before and after a period of time in the reaction: (i) the change in pH since $\text{HCl}(\text{aq})$ is an acid, (ii) the change in mass of $\text{LiCO}_3(\text{s})$, and (iii) the change in volume of $\text{CO}_2(\text{g})$.
- b.** To determine the instantaneous rate, one of these quantities must be measured over time. A graph of this quantity vs. time must be plotted and the slope of the tangent to the graph at some point in time must be measured.
- 17.** The kinetic molecular theory of gases describes the particles in constant motion, colliding with one another and with the walls of the container in elastic collisions. An increase in temperature adds kinetic energy so the particles move faster and collide more often. More of the collisions will be between particles having energy equal to or exceeding the activation energy and therefore more collisions will be effective. This will lead to an increase in the rate of the reaction.
- 18.**

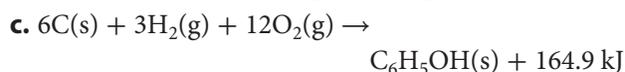
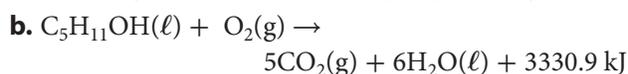
catalyst	product	process
Fe, K_2O , Al_2O_3	NH_3	Haber-Bosch
V_2O_5	H_2SO_4	Contact
Pt/Ir	HNO_3	Ostwald
xylanase	white paper	–
amylase	starch	–

- 19.** A catalyst provides a site at which an alternate pathway of lower activation energy can occur. More of the colliding particles will have a kinetic energy equal to or exceeding the activation energy barrier. There will be more collisions that are effective and the rate of reaction will increase.
- 20. a.** An elementary step is a step in a series of simple reactions that represent the progress of a chemical reaction. The sum of the elementary steps gives the overall reaction.
- b.** The rate determining step is the slowest of the elementary steps in the reaction mechanism.
- c.** The rate determining step will have the highest activation energy of all the steps in the reaction mechanism.
- d.** The reactants in the rate determining step will be part of the rate law.
- 21.** The units for k will indicate the overall order of the reaction. The magnitude of k will give an indication of the rate of reaction. The larger the value of k , the faster the rate of reaction. Knowing the value of k when the concentration of the reactants is given allows you to predict the rate of reaction.
- 22.** The temperature change can be used but there are practical difficulties to control in order for results to be considered valid. The solution will absorb the heat change as the reaction proceeds. It will be necessary to have the solution stirred so that the heat change is uniform. It will be necessary to ensure that the volume of the solution is always the same for each trial so that any changes in temperature are solely because of the reaction. The reading of a thermometer must be done quickly for each reading.
- 23.** No, it cannot be stated with certainty. Reaction rates must be tested experimentally to determine exact relationships concerning rates. If the reaction is exothermic, once some particles react, the energy given off will give more particles sufficient kinetic energy to have effective collisions and the reaction will continue. For an endothermic reaction few particles are likely to combine initially and little energy will be available, or at least less energy than for the exothermic reaction, to give other particles kinetic energy to keep the reaction going. It could be expected that the average rate of the exothermic reaction will be faster over a given time period, but it cannot be stated with certainty.

24. In solution the molecules or ions are closer together and forces of attraction exist between the particles, which are not present in gases. Therefore, the 10°C increase in temperature would not increase the kinetic energy of particles in solution as much as gas particles. Therefore, it would not result in an increase in reaction rate to the same extent.

25. The diagram correctly shows the path of the reaction without a catalyst having a lower activation energy. The diagram is incorrect because the potential energy of the product must be the same with and without a catalyst.

$$\begin{aligned} \mathbf{26.} \quad \Delta H_r^\circ &= [\Sigma(n\Delta H_f^\circ \text{ products})] - [\Sigma(n\Delta H_f^\circ \text{ reactants})] \\ &= [(3 \text{ mol})(0.0 \text{ kJ/mol})] - [(2 \text{ mol})(+142.7 \text{ kJ/mol})] \\ &= -285.4 \text{ kJ} \end{aligned}$$



$$\mathbf{28.} \quad n(\text{NH}_3) = \frac{6.46 \cancel{\text{g}}}{17.04 \frac{\cancel{\text{g}}}{\text{mol}}} = 0.3791 \text{ mol}$$

$$\Delta H_{\text{fre}} = -5.66 \text{ kJ/mol}$$

$$\text{heat given off} = n\Delta H_{\text{fre}} = 0.3791 \cancel{\text{mol}} \times -5.66 \frac{\text{kJ}}{\cancel{\text{mol}}} = -2.14 \text{ kJ}$$



b. $M(\text{CH}_4) = 16.05 \text{ g/mol}$

$$n(\text{CH}_4) = \frac{100.0 \cancel{\text{g}}}{16.05 \frac{\cancel{\text{g}}}{\text{mol}}} = 6.230 \text{ mol}$$

$$Q = n\Delta H_{\text{vap}} = 6.230 \cancel{\text{mol}} \times 8.19 \frac{\text{kJ}}{\cancel{\text{mol}}} = 51.0 \text{ kJ}$$

30. For NH_3 : -226 kJ/mol ; for O_2 : -181 kJ/mol ;
for NO : -226 kJ/mol ; for H_2O : -151 kJ/mol

$$\mathbf{31.} \quad Q = mc\Delta T = 324 \text{ g} \times 2.8 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times \Delta T$$

$$\Delta T = \frac{2560 \cancel{\text{J}}}{324 \cancel{\text{g}} \times 2.80 \frac{\text{J}}{\cancel{\text{g}} \cdot ^\circ\text{C}}} = 2.82^\circ\text{C}$$

$$T_f - T_i = 2.82^\circ\text{C}$$

$$T_f - 5.0^\circ\text{C} = 2.82^\circ\text{C}$$

$$T_f = 7.8^\circ\text{C}$$

32. a. $\frac{0.26 \text{ mol}}{355.4 \text{ kJ}} = \frac{1 \text{ mol}}{x}$

$$x = 1366.9 \text{ kJ/mol}$$

The molar heat of combustion is -1367 kJ/mol

b. ethanol

33. a. This thermochemical equation shows the decomposition of 1 mol of ammonium hydrogen sulfite, $\text{NH}_4\text{HSO}_3(\text{s})$, into its elements. It is the reverse of the formation equation for this compound.

b. $\frac{\Delta H_1}{\Delta H_2} = \frac{n_1\text{S}(\text{s})}{n_2\text{S}(\text{s})}$

$$\frac{768.6 \text{ kJ}}{2690.1 \text{ kJ}} = \frac{1 \text{ mol S}(\text{s})}{x \text{ mol S}(\text{s})}$$

$$x = 3.500 \text{ mol}$$

$$\text{mass S}(\text{s}) = nM = 3.500 \cancel{\text{mol}} \times 32.07 \frac{\text{g}}{\cancel{\text{mol}}} = 112.2 \text{ g}$$

34. By using gasoline, there is $47.30 \text{ MJ/kg} - 44.80 \text{ MJ/kg} = 2.50 \text{ MJ/kg}$ more energy is available to heat the water.

$$2.50 \cancel{\text{MJ}} \times 1 \times 10^3 \frac{\text{kJ}}{\cancel{\text{MJ}}} = 2.5 \times 10^3 \text{ kJ}$$

$$Q = n\Delta H_{\text{vap}}$$

$$n = \frac{Q}{\Delta H_{\text{vap}}} = \frac{2.5 \times 10^3 \cancel{\text{kJ}}}{40.7 \frac{\cancel{\text{kJ}}}{\text{mol}}} = 61.425 \text{ mol}$$

$$m(\text{H}_2\text{O}) = 61.425 \cancel{\text{mol}} \times 18.02 \frac{\text{g}}{\cancel{\text{mol}}} = 1.107 \times 10^3 \text{ g} = 1.11 \text{ kg}$$

35. a, b, d, e, and f

36. a. $Q = mc\Delta T = 500.0 \text{ g} \times 4.19 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times 12.3^\circ\text{C} = 25\,768 \text{ J} = 25.8 \text{ kJ}$

b. $n(\text{CH}_3\text{OH}) = \frac{m}{M} = \frac{1.84 \cancel{\text{g}}}{32.05 \frac{\cancel{\text{g}}}{\text{mol}}} = 0.05741 \text{ mol}$

$$\text{Heat given off during combustion} = n\Delta H_c^\circ$$

$$= 0.05741 \cancel{\text{mol}} \times -726.1 \frac{\text{kJ}}{\cancel{\text{mol}}} = -41.686 \text{ kJ}$$

$$= -41.69 \text{ kJ}$$

c. Energy input (the energy given off during combustion) = -41.69 kJ

Energy output (the thermal energy absorbed by the water) = 25.8 kJ

d. Efficiency = $\frac{\text{energy output}}{\text{energy input}} \times 100\%$

$$= \frac{25.768 \text{ kJ}}{41.686 \text{ kJ}} \times 100\% = 61.8\%$$

37. $\Delta H_r^\circ = \Sigma(n\Delta H_f^\circ \text{ products}) - \Sigma(n\Delta H_f^\circ \text{ reactants})$

$$-1189.2 \text{ kJ} = [(2 \text{ mol})(\Delta H_f^\circ \text{CO}_2(\text{g})) + (3 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\ell))] - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_2\text{H}_6\text{O}_2(\ell)) + (\frac{5}{2} \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g}))]$$

$$-1189.2 \text{ kJ} = [(2 \text{ mol})(-393.5 \text{ kJ/mol}) + (1 \text{ mol})(-285.8 \text{ kJ/mol})] - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_2\text{H}_6\text{O}_2(\ell)) + (\frac{5}{2} \text{ mol})(0 \text{ kJ/mol})]$$

$$-1189.2 \text{ kJ} = (-1644.4 \text{ kJ/mol}) - (\Delta H_f^\circ \text{C}_2\text{H}_6\text{O}_2(\ell))$$

$$\Delta H_f^\circ \text{C}_2\text{H}_6\text{O}_2(\ell) = -455.2 \text{ kJ}$$

The heat of formation of ethylene glycol is -455.2 kJ/mol

38. energy given off when 1.00 kg of natural gas burns

$$= 54.0 \frac{\text{kJ}}{\text{g}} \times 1000 \frac{\text{g}}{\text{kg}}$$

$$= 5.40 \times 10^4 \frac{\text{kJ}}{\text{kg}}$$

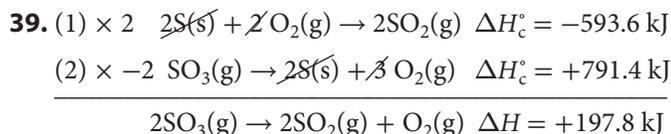
$$= 5.40 \times 10^4 \frac{\text{kJ}}{\text{kg}} \times 0.00100 \frac{\text{MJ}}{\text{kJ}} = 54.0 \frac{\text{MJ}}{\text{kg}}$$

Energy given off when 1.00 kg of natural gas burns is 54.0 MJ.

particulate emission level = $3.35 \frac{\text{kg}}{\text{MJ}}$

mass of particulate matter emitted

$$= 3.35 \frac{\text{kg}}{\text{MJ}} \times 54.0 \text{ MJ} = 1.809 \times 10^2 \text{ kg} = 181 \text{ kg}$$



40. a. $M(\text{C}_{12}\text{H}_{22}\text{O}_{11}) = 342.34 \text{ g/mol}$

$$n(\text{C}_{12}\text{H}_{22}\text{O}_{11}) = \frac{m}{M} = \frac{3.00 \text{ g}}{342.34 \frac{\text{g}}{\text{mol}}} = 0.008763 \text{ mol}$$

Heat given off during combustion

$$= n\Delta H_c^\circ = 0.008763 \text{ mol} \times -5650 \frac{\text{kJ}}{\text{mol}}$$

$$= -49.512 \text{ kJ} = -49.5 \text{ kJ}$$

b. The calorimeter will absorb the heat given off by the combustion reaction, or $Q = +49.5 \text{ kJ}$.

c. Heat capacity of bomb calorimeter = $1284 \frac{\text{J}}{^\circ\text{C}}$

$$\Delta T = \frac{49512 \text{ J}}{1284 \frac{\text{J}}{^\circ\text{C}}} = 38.6 \text{ }^\circ\text{C}$$

d. rate of burning = $\frac{n}{\Delta t}$

$$= \frac{0.008763 \text{ mol}}{38.0 \text{ s}} = 2.31 \times 10^{-4} \frac{\text{mol}}{\text{s}}$$

41. a. mol ratio $\text{Ag}^+(\text{aq}):\text{AgNO}_3(\text{aq}) = 1:1$

$$\Delta C(\text{AgNO}_3) = c_f - c_i = 0.185 \frac{\text{mol}}{\text{L}} - 0.200 \frac{\text{mol}}{\text{L}}$$

$$= 0.015 \frac{\text{mol}}{\text{L}}$$

$$\Delta \text{Ag}^+(\text{aq}) = \frac{\Delta C}{\Delta t} = \frac{0.015 \frac{\text{mol}}{\text{L}}}{25.0 \text{ min} \times \frac{60 \text{ s}}{\text{min}}}$$

$$= 1.0 \times 10^{-5} \frac{\text{mol}}{\text{L}\cdot\text{s}}$$

b. mol ratio $\text{Ag}^+(\text{aq}):\text{AgNO}_3(\text{aq}) = 1:1$
 mol ratio $\text{Fe}^{2+}(\text{aq}):\text{Fe}(\text{NO}_3)_2(\text{aq}) = 1:1$
 The relative rate of change of a reactant or product is in proportion to their mole ratio in the balanced equation.

From the balanced equation, the mole ratio $\text{Ag}^+(\text{aq}):\text{Fe}^{2+}(\text{aq})$ is 2:1

$$\frac{\text{rate of formation of Fe}^{2+}}{\text{rate of consumption of Ag}^+} = \frac{n \text{ Fe}^{2+}}{n \text{ Ag}^+} = \frac{1}{2}$$

$$\frac{\text{rate of formation of Fe}^{2+}}{1.0 \times 10^{-5} \frac{\text{mol}}{\text{L}\cdot\text{s}} \text{ Ag}^+} = \frac{1}{2}$$

$$\text{Rate of formation of Fe}^{2+}(\text{aq}) = \frac{1}{2} \times 1.0 \times 10^{-5} \frac{\text{mol}}{\text{L}\cdot\text{s}}$$

$$= 5.0 \times 10^{-6} \text{ mol/L}\cdot\text{s}$$

42. instantaneous rates are $0.0050 \text{ mol/L}\cdot\text{s}$ ($t = 0 \text{ s}$) and $0.00020 \text{ mol/L}\cdot\text{s}$ ($t = 30.0 \text{ s}$); average rate is $0.00075 \text{ mol/L}\cdot\text{s}$

43. For reaction 1: $E_{a(\text{rev})} = E_{a(\text{fwd})} + \Delta H$
 $= 25 \text{ kJ} + 50 \text{ kJ} = 75 \text{ kJ}$

For reaction 2: $E_{a(\text{rev})} = E_{a(\text{fwd})} - \Delta H$
 $25 \text{ kJ} = E_{a(\text{fwd})} - 50 \text{ kJ}$

$$E_{a(\text{fwd})} = 75 \text{ kJ}$$

$E_{a(\text{fwd})}$ and $E_{a(\text{rev})}$ are both 75 kJ

44. $M(\text{BrO}_3^-) = 127.9 \text{ g/mol}$

$$1.28 \frac{\text{g}}{\text{L}\cdot\text{s}} = \frac{1.28 \text{ g}}{127.9 \frac{\text{g}}{\text{mol}} \text{ L}\cdot\text{s}} = \frac{0.0100 \text{ mol}}{\text{L}\cdot\text{s}}$$

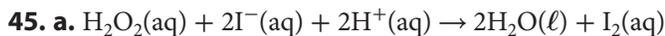
$$\frac{\text{rate of consumption BrO}_3^-(\text{aq})}{\text{rate of formation Br}_2(\ell)} = \frac{n \text{ BrO}_3^-(\text{aq})}{n \text{ Br}_2(\ell)}$$

$$\frac{0.0100 \frac{\text{mol}}{\text{L}\cdot\text{s}}}{\text{rate of formation Br}_2(\ell)} = \frac{1}{3}$$

$$\text{rate of formation of Br}_2(\ell) = 3 \times 0.0100 \frac{\text{mol}}{\text{L}\cdot\text{s}}$$

$$= 0.0300 \frac{\text{mol}}{\text{L}\cdot\text{s}}$$

$$= 3.00 \times 10^{-2} \text{ mol/L}\cdot\text{s}$$

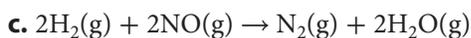


b. (i) Change in pH. As $\text{H}^+(\text{aq})$ is consumed the pH will rise.

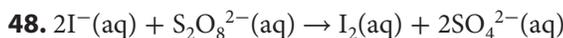
(ii) Both reactants are clear and colourless in aqueous solution. As $\text{I}_2(\text{aq})$ forms the solution will gradually turn reddish-brown as $\text{I}_2(\text{s})$ dissolves in the alcohol. This colour change could be used to monitor the rate of reaction with the use of a spectrophotometer.

46. a. No, a catalyst is not present. If a catalyst was present it would be shown as a reactant in one step and as a product in a later step.

b. $\text{N}(\text{g})$ and $\text{N}_2\text{O}(\text{g})$



47. $\frac{\text{mol}^{\frac{3}{2}}}{\text{L}^{\frac{3}{2}}\cdot\text{s}}$



49. The relative rate of change of a reactant or product is in proportion to their mole ratio in the balanced equation.

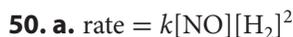
The mole ratio $\text{H}_2(\text{g}):\text{PH}_3(\text{g})$ is $\frac{3}{2}:1$. For every $\frac{3}{2}$ mol of $\text{H}_2(\text{g})$ produced, 1 mol of $\text{PH}_3(\text{g})$ decomposes.

$$\frac{\text{rate of decomposition of } \text{PH}_3(\text{g})}{\text{rate of formation of } \text{H}_2(\text{g})} = \frac{n \text{ PH}_3(\text{g})}{n \text{ H}_2(\text{g})} = \frac{2}{3}$$

$$\text{rate of decomposition of } \text{PH}_3(\text{g}) = \text{rate of formation } \text{H}_2(\text{g}) \times \frac{2}{3}$$

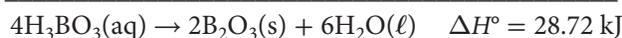
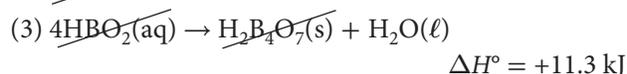
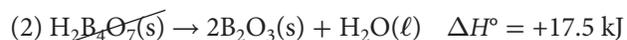
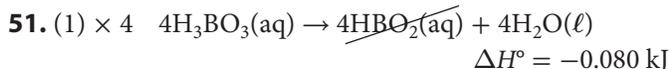
$$= 4.8 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{s}} \times \frac{2}{3}$$

$$= 3.2 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{s}}$$

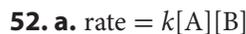
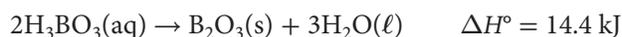


b. $\frac{\text{rate}_1}{\text{rate}_2} = \frac{k[\text{NO}(\text{g})][\text{H}_2(\text{g})]^2}{k[2\text{NO}(\text{g})][2\text{H}_2(\text{g})]^2} = \frac{1}{8}$

When both occur together the rate will increase by a factor of 8 times.

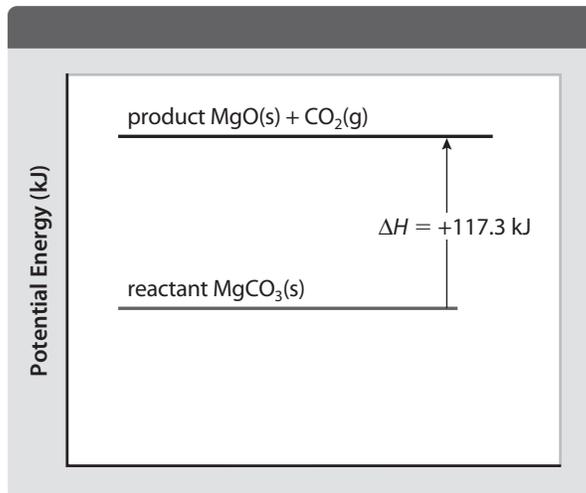


divide equation and enthalpy by 2



b. 2

53.



54. Energy is the ability to do work or produce heat. Potential energy is stored energy. Kinetic energy is energy due to movement of particles. Compare these energies before and after lighting the gas. When unlit, the natural gas by itself is not doing work or producing heat. Its energy is stored in the chemical bonds of methane gas. When lit, bonds are broken and new bonds are formed, releasing thermal energy that heats the surroundings. The thermal energy increases movement of particles, which results in an increase in kinetic energy.

55. The symbol for a change in value is the Greek letter, lower case delta, Δ . By convention, this is understood to mean the final state minus the initial state for whatever variable is described.

ΔH_{fre} represents the difference in total energy between the liquid and solid.

$$\Delta H_{\text{fre}} = \text{energy of solid (final state)} - \text{energy of liquid (initial state)}$$

Since energy must be taken away from the liquid for freezing to occur, the energy of the solid is a smaller quantity than the energy of the liquid. ΔH_{fre} will be a negative quantity.

56. Answer should include the following: ΔH_r , ΔH_f , ΔH_{comb} , ΔH_{vap} , ΔH_{cond} , ΔH_{fre} , ΔH_{melt} , ΔH_{neut} , ΔH_{soln} .

ΔH_r is the enthalpy change of a reaction.

ΔH_f is the enthalpy change of a reaction that is the formation of a compound from its elements.

ΔH_{comb} is the enthalpy change of a combustion reaction.

ΔH_{vap} , ΔH_{cond} , ΔH_{fre} , and ΔH_{melt} represent the change in enthalpy that is associated with phase changes.

ΔH_{neut} is the enthalpy change of a neutralization reaction.

ΔH_{soln} is the enthalpy change associated with a solute dissolving in a solvent.

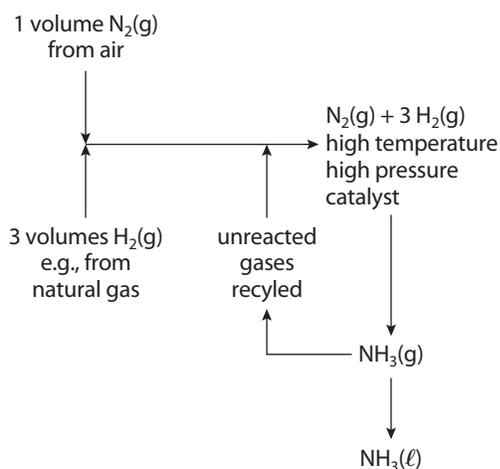
57. Student reports should be one-page in length, with a clear outline of the different processes. The following information could be included:

Catalytic cracking breaks complex hydrocarbons into simpler molecules in order to increase the quality and quantity of lighter, more desirable products such as kerosene, gasoline, and heating oil. Use of a zeolite catalyst (aluminosilicates) increases the yield of products under much less severe operating conditions than in thermal cracking.

Catalytic reforming is the reuniting of smaller hydrocarbons to make larger ones using a platinum and platinum-rhenium mix for a catalyst. For example, low weight naphtha fraction can be combined into aromatics, which are used in blending gasoline.

Catalytic isomerization rearranges molecules into structural forms of the same molar mass but different molecular structure. For example, low octane number saturated hydrocarbons are converted to high octane number hydrocarbons. Catalysts are specific for a reaction but can include Pt/Al₂O₃, MoO₃/TiO₂, and Re(CO) compounds.

58. A schematic similar to the following should be provided; more detail can be included.

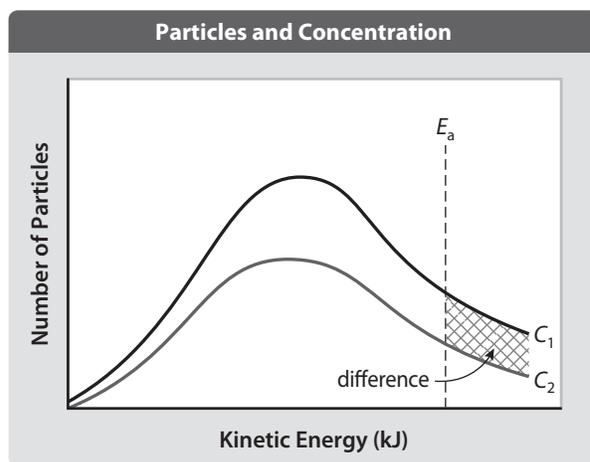


59. Possible examples include: catalytic converters, fuel cells, formation of formaldehyde from methanol

60. List the two requirements that must be met in order for a reaction to occur when particles collide, namely correct orientation and sufficient kinetic energy to equal or exceed the activation energy for the reaction. For the reaction to occur as shown in the balanced equation, 2 particles of X and 3 particles of Y must collide at one instant and meet the two requirements. Demonstrate models of the two types of reactants. This possibility has a very low probability of happening. There is a need for a reaction mechanism made up of bimolecular or two-particle collisions that will result in the equivalent overall reaction.

61. Graphic organizers should provide the definition of each term and indicate how they relate to rates of reaction and, when applicable, each other.

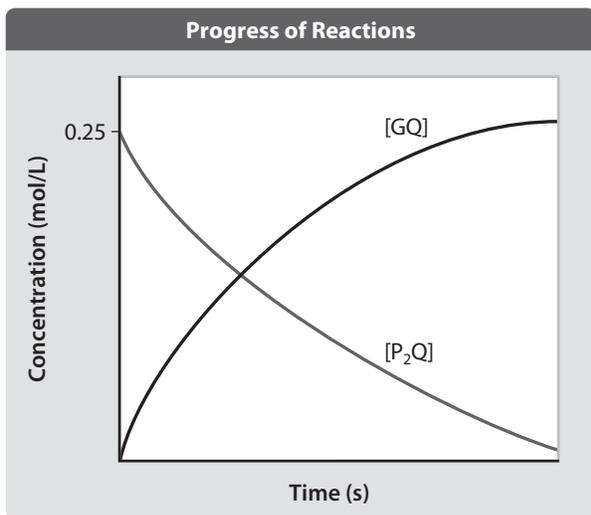
62.



In the graphic, concentration $c_1 >$ concentration c_2 .

The total number of collisions and the total number of collisions with energy greater than the activation energy are both greater at concentration c_1 . The shaded area represents the difference in the number of particles at concentration c_1 having energy greater than the activation energy, E_a . This increase in the number of collisions with energy greater than E_a results in an increase in the reaction rate.

63. Since all of P₂Q(aq) is used up, and in the balanced equation, the mole ratio P₂Q(aq):GQ(aq) is 1:1, the concentration of GQ(aq) at the completion of the reaction is 0.25 mol/L. G(s) and P(s) are not shown on this graph because they are solids and their amounts are not expressed in concentration units.



- 64.** The layer of $\text{Al}_2\text{O}_3(\text{s})$ is very thin and very little heat is actually given off per unit area of aluminum. In addition, the heat is quickly conducted away by the metal.
- 65. a.** Each metal has a different specific heat capacity. This means that it takes a different quantity of heat to change its temperature by 1°C .

b. metal 1

heat lost by metal = heat gained by water

$$mc\Delta T = mc\Delta T$$

$$12.5 \text{ g} \times c \times (150.0^\circ\text{C} - 21.5^\circ\text{C}) = 100.0 \text{ g} \times 4.19 \frac{\text{J}}{\text{g}^\circ\text{C}} \times (21.5^\circ\text{C} - 20.0^\circ\text{C})$$

$$c = 0.391 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

metal 2

heat lost by metal = heat gained by water

$$mc\Delta T = mc\Delta T$$

$$12.5 \text{ g} \times c \times (150.0^\circ\text{C} - 23.4^\circ\text{C}) = 100.0 \text{ g} \times 4.19 \frac{\text{J}}{\text{g}^\circ\text{C}} \times (23.4^\circ\text{C} - 20.0^\circ\text{C})$$

$$c = 0.900 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

metal 3

heat lost by metal = heat gained by water

$$mc\Delta T = mc\Delta T$$

$$12.5 \text{ g} \times c \times (150.0^\circ\text{C} - 21.7^\circ\text{C}) = 100.0 \text{ g} \times 4.19 \frac{\text{J}}{\text{g}^\circ\text{C}} \times (21.7^\circ\text{C} - 20.0^\circ\text{C})$$

$$c = 0.444 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

Referring to Table 5.1, metal 1 is Cu, metal 2 is Al, and metal 3 is Fe.

- 66.** *Sample answers:* cold temperature of a refrigeration unit to reduce the rate of food spoilage; application of a paint to a surface to prevent corrosion; vacuum packaging of food to exclude air and water to reduce the rate of spoilage; package with an inactive gas such

as nitrogen to keep food out of contact with air and prevent drying out and spoilage; a solution of hydrogen peroxide will include an inhibitor to reduce the rate of decomposition of H_2O_2 .

- 67. a.** heat gained by surroundings

$$= (mc\Delta T)_{\text{water}} + (mc\Delta T)_{\text{iron}}$$

$$= 882 \text{ g} \times 4.19 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 85.0^\circ\text{C} + 1208 \text{ g} \times 0.449 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 85.0^\circ\text{C}$$

$$= 314\,124 \text{ J} + 46\,103 \text{ J}$$

$$= 360\,227 \text{ J} = 3.602 \times 10^5 \text{ J}$$

The transfer of heat from the burner to the pot of water is 45% efficient.

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\%$$

$$45.0\% = \frac{-3.602 \times 10^5 \text{ J}}{\text{Energy input}} \times 100\%$$

$$\text{Energy input} = 8.0044 \times 10^5 \text{ kJ}$$

Heat given off from burning butane = ΔH

$$= -8.0044 \times 10^5 \text{ kJ}$$

$$\Delta H_c \text{ butane} = -2877.6 \text{ kJ/mol}$$

$$n = \frac{\Delta H}{\Delta H_c} = \frac{-8.0044 \times 10^5 \text{ kJ}}{-2877.6 \frac{\text{kJ}}{\text{mol}}} = 278.16 \text{ mol}$$

$$\text{mass of } \text{C}_4\text{H}_{10} = n \times M = 278.16 \text{ mol} \times 58.14 \frac{\text{g}}{\text{mol}} = 1.6172 \times 10^4 \text{ g or } 16.2 \text{ kg}$$

- b.** Rate of burning butane = $\frac{\Delta Q}{\Delta t}$

$$= \frac{278.16 \text{ mol}}{75.0 \text{ min}} = 3.71 \frac{\text{mol}}{\text{min}}$$

- 68. a.** Incomplete combustion occurs when the fuel is first lit and burns with an orange flame. Complete combustion occurs when there is more oxygen present and the flame burns blue.

- b.** $\Delta H_r^\circ = \Sigma(n\Delta H_f^\circ \text{ products}) - \Sigma(n\Delta H_f^\circ \text{ reactants})$

$$\Delta H_{\text{comb}}^\circ = [(2 \text{ mol})(\Delta H_f^\circ \text{ CO}_2(\text{g})) + (1 \text{ mol})(\Delta H_f^\circ \text{ H}_2\text{O}(\text{g}))] - [(1 \text{ mol})(\Delta H_f^\circ \text{ C}_2\text{H}_2(\text{g})) + (\frac{5}{2} \text{ mol})(\Delta H_f^\circ \text{ O}_2(\text{g}))]$$

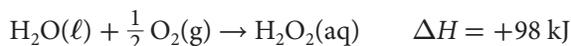
$$= [(2 \text{ mol})(-393.5 \text{ kJ/mol}) + (1 \text{ mol})(-241.8 \text{ kJ/mol})] - [(1 \text{ mol})(227.4 \text{ kJ/mol}) + (\frac{5}{2} \text{ mol})(0 \text{ kJ/mol})]$$

$$= (-1028.8 \text{ kJ/mol}) - (227.4 \text{ kJ/mol})$$

$$= -1256.2 \text{ kJ/mol}$$

The enthalpy of combustion for complete combustion of acetylene is -1256.2 kJ/mol .

d. The reverse reaction will be



A heat of formation equation shows the elements as reactants and one mole of compound as product. This is not a heat of formation equation because the reactants include a compound, $\text{H}_2\text{O}(\ell)$.

74. The rate of reaction is directly proportional to the concentration of reactant X. The rate is independent of the concentration of Y. If the volume of the container is doubled, the concentration of all reactants will be halved. Only the change in the concentration of X will affect the rate. Since the concentration of X is half the original concentration, the rate will be reduced to half the original rate.

Answers to Unit 3 Self-Assessment Questions

(Student textbook pages 410-11)

- e
- d
- c
- b
- a
- d
- b
- d
- b
- c
- This is a formation reaction for the compound copper(II) chloride from its elements. The compound $\text{CuCl}_2(\text{s})$ forms from its elements in an exothermic reaction. Measured at the standard state of 25°C and 100 kPa pressure, 1 mol of $\text{Cu}(\text{s})$ and 1 mol of $\text{Cl}_2(\text{g})$ react to release 220.1 kJ of energy and form $\text{CuCl}_2(\text{s})$.

12. (1) $6\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\ell) \quad \Delta H_f^\circ = +49.1 \text{ kJ/mol}$
 (2) $6\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\text{g}) \quad \Delta H_f^\circ = +82.91 \text{ kJ/mol}$
 (1) $\times -1 \quad \text{C}_6\text{H}_6(\ell) \rightarrow 6\text{C}(\text{s}) + 3\text{H}_2(\text{g})$
 $\Delta H_f^\circ = -49.1 \text{ kJ/mol}$

$$\text{C}_6\text{H}_6(\ell) \rightarrow \text{C}_6\text{H}_6(\text{g}) \quad \Delta H_{\text{vap}}^\circ = +33.8 \text{ kJ/mol}$$

13. The system is isolated so that any thermal energy exchanged with the surroundings outside the calorimeter is small and can be ignored. The thermal energy that is exchanged with the calorimeter (cup, lid, stirring rod) is small enough to be ignored. The density and specific heat capacity of any solutions are the same as for water.

14. H_f° reactions show the elements as reactants in the correct mole ratio as they appear in the product, 1 mol of compound. In the equation as written, one of the reactants is a compound, $\text{CO}(\text{g})$. The correct equation for the heat of formation of $\text{CO}_2(\text{g})$ is $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$.

$$\begin{aligned} 15. n(\text{H}_2\text{O}) &= \frac{9.83 \times 10^{13} \frac{\text{g}}{\text{mol}}}{18.02 \frac{\text{g}}{\text{mol}}} = 5.4550 \times 10^{12} \text{ mol} \\ &= (5.4550 \times 10^{12} \text{ mol}) \times \left(-40.7 \frac{\text{kJ}}{\text{mol}}\right) \\ &= -2.22 \times 10^{14} \text{ kJ} \end{aligned}$$

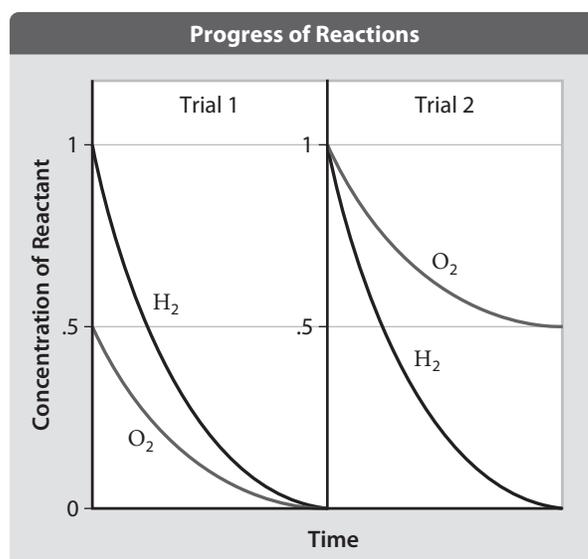
$$\begin{aligned} 16. \Delta H_r^\circ &= [\Sigma(n\Delta H_f^\circ \text{ products})] - [\Sigma(n\Delta H_f^\circ \text{ reactants})] \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{CO}_2(\text{g})) + (2 \text{ mol})(\Delta H_f^\circ \text{SO}_2(\text{g}))] - \\ &\quad [(1 \text{ mol})(\Delta H_f^\circ \text{CS}_2(\ell)) + (3 \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g}))] \\ &= [(1 \text{ mol})(-393.5 \text{ kJ/mol}) + (2 \text{ mol})(-269.8 \text{ kJ/mol})] - \\ &\quad [(1 \text{ mol})(+89 \text{ kJ/mol}) + (3 \text{ mol})(0 \text{ kJ/mol})] \\ &= -1022.1 \text{ kJ} \end{aligned}$$

This is the heat given off when 2 mol of $\text{SO}_2(\text{g})$ is produced.

The heat of reaction per mol $\text{SO}_2(\text{g})$ is -511.0 kJ/mol .

$$\begin{aligned} 17. \text{heat gained by ethanol} &= mc\Delta T \\ &= 250.0 \text{ g} \cdot 2.44 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times 30.0 ^\circ\text{C} \\ &= 1.830 \times 10^4 \text{ J} = 18.30 \text{ kJ} \\ \text{Efficiency} &= \frac{\text{Energy output}}{\text{Energy input}} \times 100\% \\ &= \frac{18.30 \text{ kJ}}{27.3 \text{ kJ}} \times 100\% = 67.0\% \end{aligned}$$

18. From the balanced equation, the gases react as $\text{H}_2(\text{g}):\text{O}_2(\text{g}) = 2:1$. In the first trial, there will be no reactant gases remaining after the reaction. In the second trial, the $\text{O}_2(\text{g})$ is in excess and half of the original amount of this gas will remain.



- 19.** Since the reaction occurs only when the heat is supplied, the reaction must be endothermic. Thermal energy must be supplied to the particles of $\text{HgO}(s)$ that equals or exceeds the activation energy barrier for the reaction.
- 20.** Every reaction will have an activation energy. The activation energy for these reactions is not zero. The collision necessary to cause a reaction will occur between two ions that are opposite in charge. The attractions that must be overcome in aqueous solution are solvent-solvent and solvent-solute attractions, both of which will be small in magnitude. The thermal energy in the solutions at room temperature is sufficient to supply this small amount of energy.
- 21.** The steps associated with using Hess's law and enthalpies of formation is given as an example. Flowcharts should include the following key components.
- Write the balanced chemical equation.
for example:
$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$$
 - Use the following equation and standard molar enthalpies of formation data from a reference source.
$$\Delta H_r^\circ = \Sigma(n\Delta H_f^\circ \text{ products}) - \Sigma(n\Delta H_f^\circ \text{ reactants})$$

$$\Delta H_{\text{comb}}^\circ = [(1 \text{ mol})(\Delta H_f^\circ \text{CO}_2(g)) + (2 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(g))] - [(1 \text{ mol})(\Delta H_f^\circ \text{CH}_4(g)) + (2 \text{ mol})(\Delta H_f^\circ \text{O}_2(g))]$$

$$= [(1 \text{ mol})(-393.5 \text{ kJ/mol}) + (2 \text{ mol})(-241.8 \text{ kJ/mol})] - [(1 \text{ mol})(-74.6 \text{ kJ/mol}) + (2 \text{ mol})(0 \text{ kJ/mol})]$$

$$= [(-393.5 \text{ kJ}) - (483.6 \text{ kJ/mol})] - (-74.6)$$

$$= -802.5 \text{ kJ}$$
- 22. a.** $2\text{NO}_2(g) + \text{Br}_2(g) \rightarrow 2\text{NO}_2\text{Br}(g)$
b. step 2
c. 1
- 23.** The reactants in the rate determining step will be present in the rate law. The rate determining step is the slow step in the mechanism. In the proposed mechanism, the rate determining step shows an intermediate in the reaction mechanism. An intermediate is produced in one step and used up in a subsequent step. An intermediate will not be a component in a rate law. The mechanism is not consistent with the rate law.
- 24.** Enthalpy changes associated with nuclear changes are much greater than those of physical and chemical changes. An example of a nuclear change is the 2×10^{10} kJ of energy released when 1 mol of uranium-235 undergoes nuclear fission. Enthalpy changes of chemical changes are greater than enthalpy changes of physical changes. An example is the enthalpy of combustion of methane, which is -890 kJ. An example of enthalpy change associated with physical changes is the enthalpy of melting for water, which is approximately 6 kJ/mol.
- 25.** The projected use of energy resources by 2015 is 40%. These projections include 15% for wind energy and up to 5% for solar energy. Advantages of wind energy include low operating costs and no chemical pollutants. Disadvantages include noise pollution and potential danger to birds near the wind farms. Advantages to solar energy are that the source is free and unlimited. However, the technology is still expensive. The majority of the remaining renewable energy is hydroelectric power. Advantages of hydroelectric power include no combustion emissions and high efficiency. Disadvantages include flooding of land and damage to habitats in the region that the dam is built.