

Mr. Standring's Chemistry 30 Solutions Package 2013

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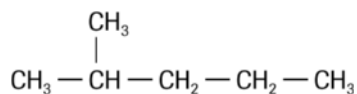
Pages are in order, but many are missing. Just look for the page #'s at the bottom left or right corner.

Page #	Topic	Text Pages
298-321	Hydrocarbons (Alkanes etc.) the Petroleum Industry	350-409
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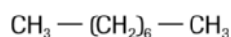
Solutions for Naming Aliphatic and Aromatic Hydrocarbons, Extra Exercises

1. For each of the following IUPAC names, draw a structural diagram.

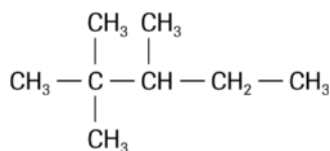
(a) methylpentane



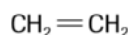
(b) octane



(c) 2,3,4-trimethylpentane



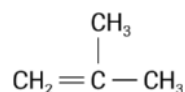
(d) ethene



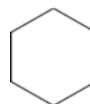
(e) propyne



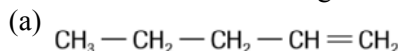
(f) methylpropene



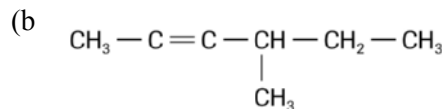
(g) cyclohexane



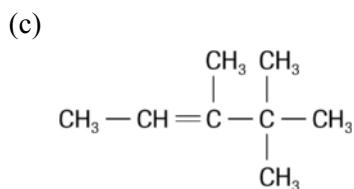
2. For each of the following structural diagrams, write the IUPAC name.



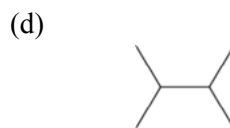
pent-1-ene



4-methylhex-2-yne



3,4,5-trimethylpent-2-ene



2,3-dimethylbutane

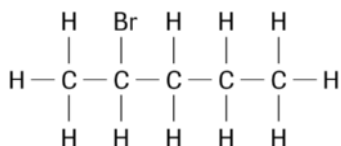
Chapter 9 Review, Part 1

1. C
2. D
3. A
4. 1, 4, 5, 8
5. 3, 6, 9
6. 1, 4, 5, 7
7. B
8. D
9. A
10. C
11. A
12. B
13. C
14. C
15. A & C

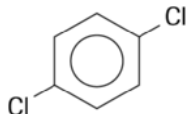
Solutions for Identification and Naming of Organic Halides, Extra Exercises

1. In the following questions, the IUPAC names of a variety of organic compounds are provided. Draw a structural diagram for each compound.

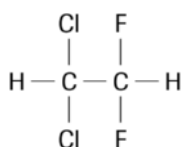
(a) 2-bromopentane



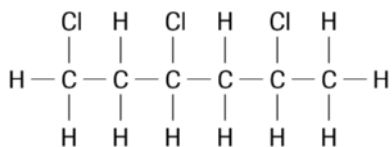
(b) 1,4-dichlorobenzene



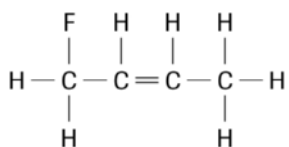
(c) 1,1-dichloro-2,2-difluoroethane



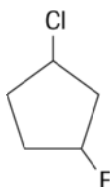
(d) 1,3,5-trichlorohexane



(e) 1-fluorobut-2-ene

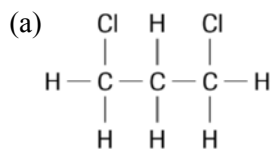


(f) 1-chloro-3-fluorocyclopentane

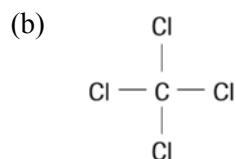


LSM 10.2B (cont'd)

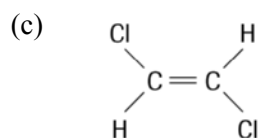
2. Give the IUPAC name for the following organic compound represented by the given structural formula.



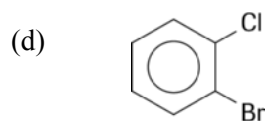
1, 3-dichloropropane



carbon tetrachloride or 1,2,3,4-tetrachloromethane



1,2-dichloroethene or *trans*-1,2-dichloroethene



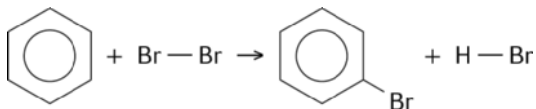
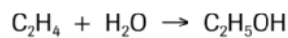
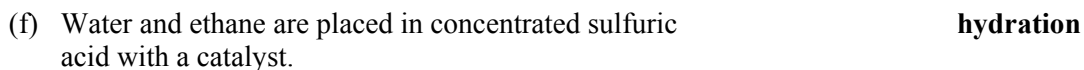
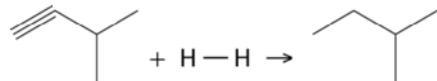
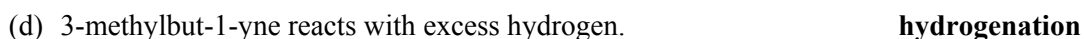
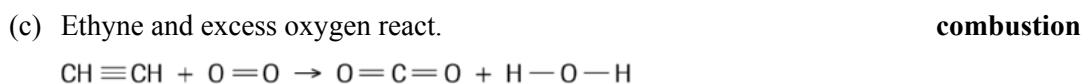
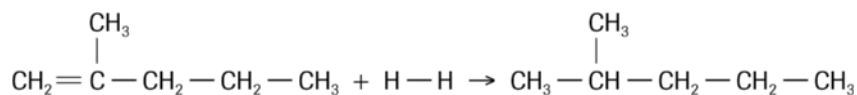
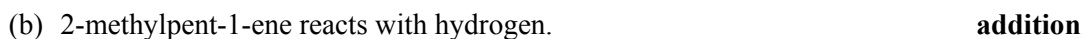
1-bromo-2-chlorobenzene

Student Worksheet Solutions LSM 10.2D

Solutions for Reactions of Hydrocarbons, Extra Exercises

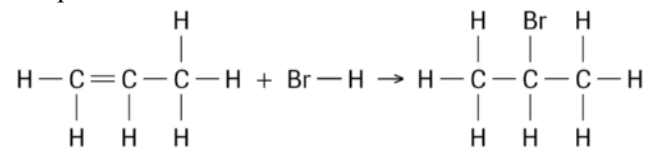
1. Write an equation for each of the following reactions. Use molecular or structural formulas and classify the reaction as combustion, addition, substitution, hydrogenation, halogenation, hydration, or hydrohalogenation.

[Note: Molecular, complete structural, condensed structural, or line structural formulas should all be considered acceptable.]



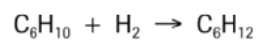
LSM 10.2D (cont'd)

(h) Propene and HBr combine.



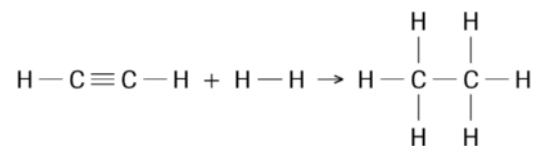
hydrohalogenation

(i) Cyclohexene reacts with hydrogen gas.



addition/hydrogenation

(j) Acetylene reacts with excess hydrogen.



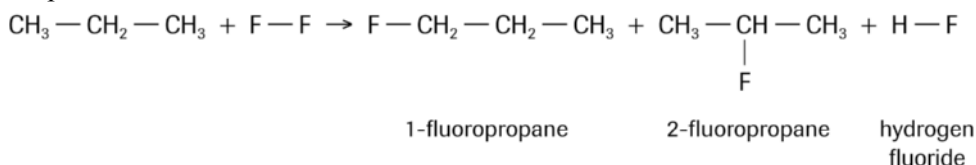
addition

Student Worksheet Solutions LSM 10.2G

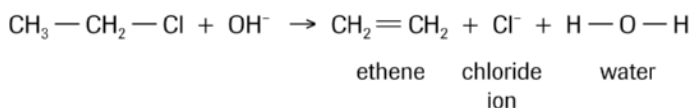
Solutions for Organic Reactions, Extra Exercises

For each of the following questions, state the organic reaction type, draw condensed structural diagrams for all reactants and products, and name all products. (You do not need to balance the equations.)

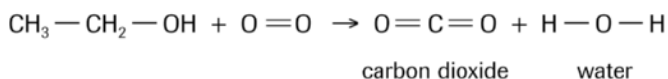
1. Propane reacts with fluorine **Substitution**



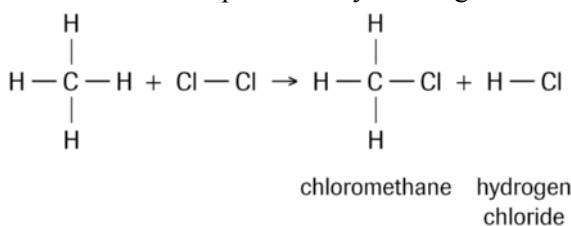
2. Chloroethane reacts with hydroxide ions **Elimination**



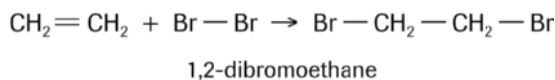
3. Ethanol, $\text{CH}_3\text{CH}_2\text{OH}$, burns in an automobile engine. **Combustion**



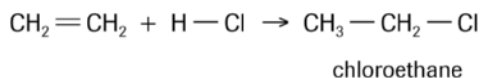
4. Chloromethane is produced by reacting methane and chlorine. **Substitution**



5. Bromine and ethene react to form an organic halide. **Addition**



6. Hydrogen chloride and ethene react to produce an organic halide **Addition**



Solutions for Applications of Organic Reactions, Extra Exercises

1. Research examples of the use of organic chemistry to address health, safety, or environmental problems, and write a report or present one such case study. Examples of topics include leaded and unleaded gasoline, solvents in dry cleaning, aerosol propellants, and pesticides and fertilizers.

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Methanol Fuel Cell

Los Alamos National Laboratory and Motorola are developing a fuel cell that produces electricity using methanol as a fuel. This fuel cell is expected to last the lifetime of a cellular phone (or other small electronic devices). Methanol is a clean and potentially renewable fuel. The methanol fuel cell would eliminate the production and disposal of rechargeable batteries that consume non-renewable resources and often create disposal problems because of toxic contents.

Production of Maleic Anhydride

Maleic anhydride is not a familiar organic substance but it is widely used in the manufacture of important products such as polyester resins and paints, polyurethane, and various solvents. The traditional method to produce maleic anhydride used benzene as a starting material but this method is only 67% efficient (based on carbon) and also produces carbon dioxide, a greenhouse gas. The same product can be produced from either butane or butene with 100% efficiencies (based on carbon) because it produces no carbon dioxide product. Overall, the butane or butene method is not only better for the environment but also less expensive. Research and development continues. Some chemical manufacturers are now producing maleic anhydride as a by-product of the oxidation of naphthalene in another industrial process. This means that a useful product is now being obtained from what would otherwise be waste.

Green Solvents

There are many harmful solvents that have traditionally been used in industry and research. Solvents such as benzene, carbon tetrachloride and chloroform ($\text{CHCl}_3(\text{l})$) have been widely used and are harmful to workers and also the environment when released. The replacement of harmful solvents in various processes is a good example of the use of chemistry to address health, safety and environmental problems. Most of this research is done by industrial process chemist and engineers, resulting in the development of several potential solutions. These include the use of liquid carbon dioxide, ionic liquids, and high temperature water.

LSM 10.2I (cont'd)

2. Why was mustard gas such an effective weapon, both during the First World War and more recently? Research its properties and effects, and defences developed against it.

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Mustard gas, $\text{ClCH}_2\text{CH}_2\text{SCH}_2\text{CH}_2\text{Cl}(\text{g})$, when inhaled, reacts with water in the lungs to produce hydrochloric acid and other intermediate products, many of which damage lung and skin tissue. Mustard gas causes mustard-coloured blisters, burns to mucous membranes, burning and swelling of the eyes, bronchitis, choking, and high levels can lead to death. During World War I, mustard gas could penetrate the materials used for clothing and masks. In more recent years, urethane has been discovered and used to resist exposure to mustard gas at a wide range of temperatures.

3. Long-term replacements for CFCs and HCFCs include hydrofluorocarbons (HFCs), which have no chlorine atoms.

HFCs are now being used in refrigerators, automobile air conditioners, aerosol cans, and as foaming agents.

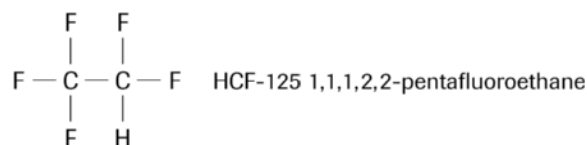
- (a) Research the latest information on HFCs and on alternative chemicals or technologies that are being used in refrigerators and air conditioners.

Present your findings in a brief report.

HCFCs and HFCs are alternatives to CFCs that are still in use. They appear to be lessdamaging to the ozone layer than CFCs (although the “hole” in the ozone layer is still growing). However, HCFCs and HFCs may act as greenhouse gases, possibly leading to global warming.

- (b) Name, and draw structural formulas for, two chlorine-free organic halides that could potentially be used in place of CFCs and HCFCs.

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Student Worksheet Solutions LSM 10.3D

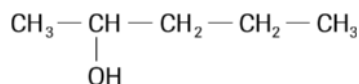
Solutions for Alcohols and Elimination Reactions, Extra Exercises

1. Draw a structural diagram for each compound named.

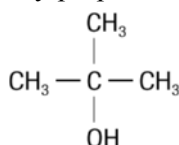
(a) butan-1-ol



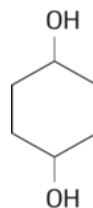
(d) pentan-2-ol



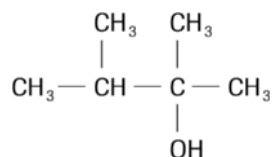
(b) 2-methylpropan-2-ol



(e) cyclohexane-1,4-diol

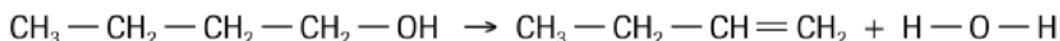


(c) 1,2-dimethylbutan-2-ol



2. For each of the following questions, draw condensed structural diagrams for all reactants and products, and name all products. (You do not need to balance the equations.)

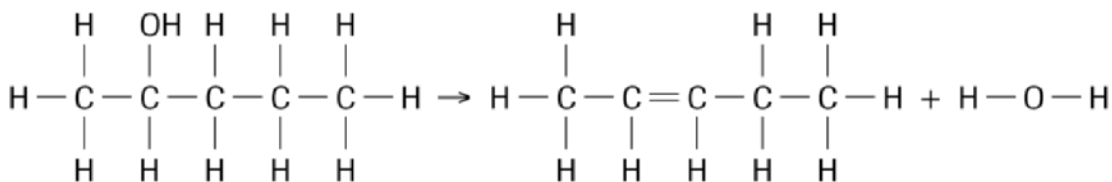
(a) Butan-1-ol reacts in the presence of concentrated sulfuric acid.



1-butene

water

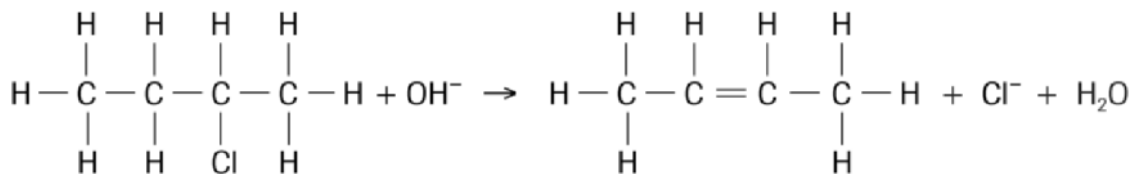
(b) A 5-carbon alkene is formed when a secondary alcohol undergoes an elimination reaction.



pent-2-ene

water

(c) In the presence of a strong base, 2-chlorobutane forms releases water and another small entity as it is transformed to an unsaturated compound



but-2-ene

chloride ion

water

LSM 10.3D (cont'd)

3. Considering the technological difficulties and the economic cost of producing ethene, which of the following solutions would you choose? Justify your selection:
- burn and/or export the natural gas without removing the ethane
 - remove the ethane and export it separately in a special pipeline
 - crack the ethane to ethene and e
 - use the ethene in Alberta to produce tertiary products, such as plastics

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Alberta should produce more tertiary products such as plastics. The reason tertiary products are desirable is that their production would create a large number of jobs and bring many economic advantages to Alberta. Using Figure 3, page 7, as a reference, if ethylene production requires 11 jobs, while the styrene monomer production requires 100 jobs, then the tertiary polystyrene plastic production would provide 322 jobs. This multiplier effect would create an economic windfall for Alberta once the industry is established.

Student Worksheet Solutions LSM 10.CR

Chapter 10 Review, Part 1

1. B
2. A
3. B
4. 1, 2, 6, 9
5. 4, 5, 7, 8
6. D
7. C
8. D
9. 4, 2, 5, 7
10. 7, 5, 2, 4
11. 9, 1, 8, 3
12. 4, 2 (*or 2, 4*)
13. D
14. C

Student Worksheet Solutions LSM 10.UR
Unit 5 Review, Part 1

1. C
2. D
3. 3, 1, 4, 2
4. B
5. A
6. D
7. C
8. B
9. A
10. D
11. B
12. 4, 9
13. 1, 8
14. 5, 7
15. 2, 6
16. 2, 6
17. 2
18. 5, 7
19. 4

Student Worksheet Solutions LSM 11.0B

Solutions for Thermal Energy Calculations, Extra Exercises

1. Calculate the quantity of heat required to warm 1.25 L of water from 22.0 °C to 98.0 °C in an electric kettle.

$$\begin{aligned}Q &= mc\Delta T \\ &= 1250 \text{ g} \times 4.19 \text{ J}/(\text{g}\cdot^\circ\text{C}) \times (98.0 - 22.0) \text{ }^\circ\text{C} \\ &= 398 \text{ kJ}\end{aligned}$$

2. What mass of aluminum in a car engine will absorb 1.00 MJ of heat when the temperature rises from 22 °C to 102 °C after the car is started?

$$\begin{aligned}Q &= mc\Delta t \\ 1.00 \text{ MJ} &= m \times 0.900 \text{ J}/(\text{g}\cdot^\circ\text{C}) \times (102 - 22) \text{ }^\circ\text{C} \\ m &= 14 \text{ kg}\end{aligned}$$

3. In a laboratory experiment, 2.00 kJ of heat flowed to a 100 g sample of a liquid solvent, causing a temperature increase from 15.40 °C to 21.37 °C. Calculate the specific heat capacity of the liquid.

$$\begin{aligned}Q &= mc\Delta t \\ 2.00 \text{ MJ} &= 100 \text{ g} \times c \times (21.37 - 15.40) \text{ }^\circ\text{C} \\ c &= 3.35 \text{ J}/(\text{g}\cdot^\circ\text{C})\end{aligned}$$

4. A human body loses about 360 kJ of heat every hour. Assuming that an average human body is equivalent to about 60 kg of water, what temperature decrease would this heat transfer cause? (Of course, this heat is replaced by body metabolism.)

$$\begin{aligned}Q &= mc\Delta t \\ 360 \text{ kJ} &= 60 \text{ kg} \times 4.19 \text{ J}/(\text{g}\cdot^\circ\text{C}) \times \Delta t \\ \Delta t &= 1.4 \text{ }^\circ\text{C}\end{aligned}$$

Solutions for Bomb Calorimetry,
Extra Exercise (Extension)

1. An oxygen bomb calorimeter has a heat capacity of 6.49 kJ/°C. The complete combustion of 1.12 g of ethyne, C₂H₂(g), produces a temperature change from 18.60 °C to 27.15 °C. Calculate the molar enthalpy of combustion, Δ_cH_m, for ethyne.

$$C = 6.49 \text{ kJ/}^\circ\text{C}$$

$$\Delta t = 27.15 - 18.60 \text{ }^\circ\text{C} = 8.55 \text{ }^\circ\text{C}$$

$$n_{\text{C}_2\text{H}_2} = 1.12 \text{ g} \times \frac{1 \text{ mol}}{26.04 \text{ g}} = 0.0430 \text{ mol}$$

$$Q = C\Delta t$$

$$= 6.49 \text{ kJ/}^\circ\text{C} \times 8.55 \text{ }^\circ\text{C}$$

$$= 55.4895 \text{ kJ}$$

$$\Delta_c H = Q = 55.4895 \text{ kJ}$$

$$\Delta_c H_{\text{mol}} = \frac{55.4895 \text{ kJ}}{0.0430 \text{ mol}} = 1290 \text{ kJ/mol or } 1.29 \text{ MJ/mol}$$

The reaction is exothermic, so the molar enthalpy of combustion for ethyne is **-1.29 MJ/mol**.

2. Canadian inventors have developed zeolite, a natural aluminium silicate mineral, as a storage medium for solar heat. Zeolite releases heat when hydrated with water. In a test, zeolite is used to heat water in a tank that has a heat capacity of 157 kJ/°C. What is the enthalpy change of hydration (ΔH_h) for zeolite if the temperature of the water increases from 27 °C to 73 °C?

$$C = 157 \text{ kJ/}^\circ\text{C}$$

$$\Delta t = 73 - 27 \text{ }^\circ\text{C} = 46 \text{ }^\circ\text{C}$$

$$Q = C\Delta t$$

$$= 157 \text{ kJ/}^\circ\text{C} \times 46 \text{ }^\circ\text{C}$$

$$= 7200 \text{ kJ}$$

$$\Delta_h H = Q = 7200 \text{ kJ or } 7.2 \text{ MJ}$$

The reaction is exothermic, so the enthalpy change of hydration for zeolite is **-7.2 MJ**.

3. Besides the molar enthalpy of combustion as determined in a bomb calorimeter, what other properties or factors are involved in evaluating automobile fuels such as propane, ethanol, and hydrogen?

Properties or factors to consider include:

- availability
- effectiveness as a fuel
- cost of production
- ease of distribution and storage
- acceptance by consumers
- impact on the environment

Solutions for Enthalpy Changes and Calorimetry, Extra Exercises

- An experiment produces evidence that the evaporation of 4.00 g of liquid butane, $C_4H_{10}(l)$, requires a gain in enthalpy of 1.67 kJ. Find the molar enthalpy of vaporization for butane from this evidence.

$$\Delta H = n\Delta_v H$$

$$1.67 \text{ kJ} = 4.00 \text{ g} \times (1 \text{ mol}/58.14 \text{ g}) \times (6.03 \text{ kJ/mol}) \times \Delta_{\text{vap}} H$$

$$\Delta_{\text{vap}} H = 24.3 \text{ kJ/mol}$$
- A 40.0 kg industrial calorimeter has an overall specific heat capacity of $1.00 \text{ J}/(\text{g}\cdot^\circ\text{C})$. Complete combustion of 1.00 g of hydrogen in this calorimeter causes a temperature increase of $3.54 \text{ }^\circ\text{C}$. Calculate the molar enthalpy of combustion for hydrogen from this evidence.

$$\Delta H = Q$$

H_2 (calorimeter)

$$n\Delta_c H_m = mc\Delta t$$

$$1.00 \text{ g} \times (1 \text{ mol}/2.02 \text{ g}) \times \Delta_c H_m = 40.0 \times 10^3 \text{ g} \times 1.00 \text{ J}/(\text{g}\cdot^\circ\text{C}) \times 3.54 \text{ }^\circ\text{C}$$

$$\Delta_c H_m = 286 \text{ kJ/mol}$$

Because this is an exothermic reaction, the molar enthalpy of combustion for hydrogen, according to this evidence, is -286 kJ/mol .
- Combustion of 3.50 g of ethanol, $C_2H_5OH(l)$, in a calorimeter with an overall specific heat capacity of $1.52 \text{ J}/(\text{g}\cdot^\circ\text{C})$ causes a temperature increase from $19.88 \text{ }^\circ\text{C}$ to $26.18 \text{ }^\circ\text{C}$. The mass of the calorimeter, including water, is 10.0 kg. Find the molar enthalpy of combustion for ethanol from this evidence.

$$\Delta H = Q$$

C_2H_5OH (calorimeter)

$$n\Delta_c H_m = mc\Delta t$$

$$3.50 \text{ g} \times (1 \text{ mol}/46.08 \text{ g}) \times \Delta_c H_m = 10.0 \times 10^3 \text{ g} \times 1.52 \text{ J}/(\text{g}\cdot^\circ\text{C}) \times (26.18 - 19.88) \text{ }^\circ\text{C}$$

$$\Delta_c H_m = 1.26 \text{ MJ/mol}$$

Because this is an exothermic reaction, the molar enthalpy of combustion for ethanol, according to this evidence, is -1.26 MJ/mol .

Solutions for Representing Enthalpy Changes, Extra Exercises

- Iron(II) sulfide ore is roasted according to the following chemical equation.

$$4 \text{FeS(s)} + 7 \text{O}_2\text{(g)} \rightarrow 2 \text{Fe}_2\text{O}_3\text{(s)} + 4 \text{SO}_2\text{(g)} \quad \Delta_c H = -2456 \text{ kJ}$$
 - Rewrite this chemical equation including the energy as a term in the balanced equation.

$$4 \text{FeS(s)} + 7 \text{O}_2\text{(g)} \rightarrow 2 \text{Fe}_2\text{O}_3\text{(s)} + 4 \text{SO}_2\text{(g)} + 2456 \text{ kJ}$$
 - What is the molar enthalpy for iron(II) sulfide in this reaction?

$$\Delta_c H = -2456 \text{ kJ}/4 \text{ mol} = -614 \text{ kJ/mol}$$

FeS
 - What is the molar enthalpy for iron(III) oxide in this reaction?

$$\Delta_c H = -2456 \text{ kJ}/2 \text{ mol} = -1228 \text{ kJ/mol}$$

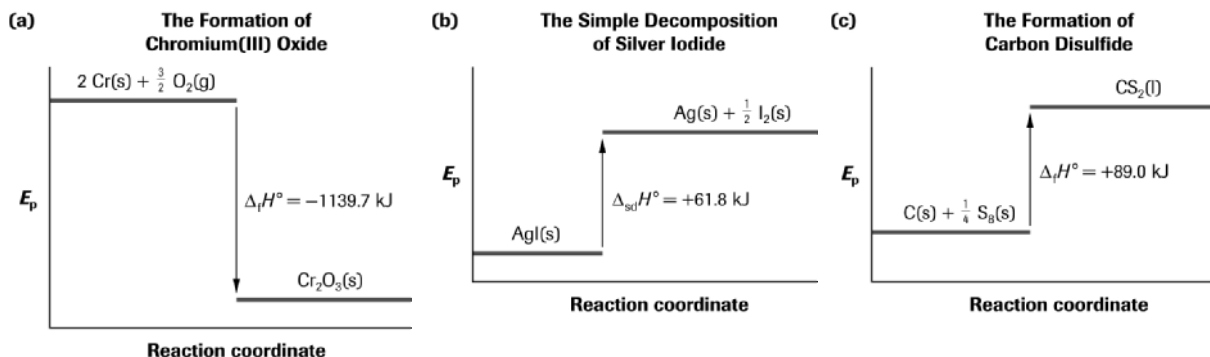
Fe₂O₃
- Boron reacts with hydrogen to form diboron hexahydride (diborane) gas. The molar enthalpy of reaction for boron is +15.7 kJ/mol. Write the balanced chemical equation using whole number coefficients and
 - including the energy change as a $\Delta_r H$.

$$2 \text{B(s)} + 3 \text{H}_2\text{(g)} \rightarrow \text{B}_2\text{H}_6\text{(g)} \quad \Delta_c H = 2 \text{ mol} \times 15.7 \text{ kJ/mol} = +31.4 \text{ kJ}$$
 - including the energy change as a term in the balanced equation.

$$2 \text{B(s)} + 3 \text{H}_2\text{(g)} + 31.4 \text{ kJ} \rightarrow \text{B}_2\text{H}_6\text{(g)}$$
- The molar enthalpy of combustion for octane, C₈H₁₈(l), is reported to be -1.3 MJ/mol. Write the balanced chemical equation using whole number coefficients and
 - including the energy change as a $\Delta_r H$.

$$2 \text{C}_8\text{H}_{18}\text{(l)} + 25 \text{O}_2\text{(g)} \rightarrow 16 \text{CO}_2\text{(g)} + 18 \text{H}_2\text{O(g)} \quad \Delta_c H = 2 \text{ mol} \times -1.3 \text{ MJ/mol} = -2.6 \text{ MJ}$$
 - including the energy change as a term in the balanced equation.

$$2 \text{C}_8\text{H}_{18}\text{(l)} + 25 \text{O}_2\text{(g)} \rightarrow 16 \text{CO}_2\text{(g)} + 18 \text{H}_2\text{O(g)} + 2.6 \text{ MJ}$$
- Draw potential energy diagrams to communicate the following chemical reactions. Assume SATP conditions.

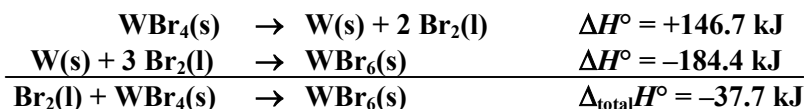
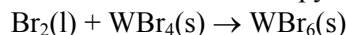


Solutions for Predicting ΔH Using Hess' Law, Extra Exercises

1. The enthalpy changes for the formation of two tungsten bromides are shown below.

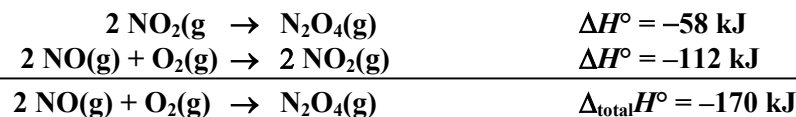
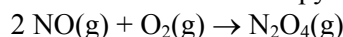


Calculate the standard enthalpy change for the following reaction.

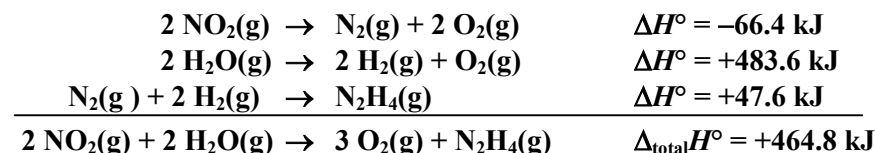
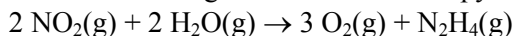


2. Given: (1) $\text{N}_2\text{O}_4(\text{g}) \rightarrow 2 \text{NO}_2(\text{g}) \qquad \Delta_1 H^\circ = +58 \text{ kJ}$
 (2) $\text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \qquad \Delta_2 H^\circ = -56 \text{ kJ}$

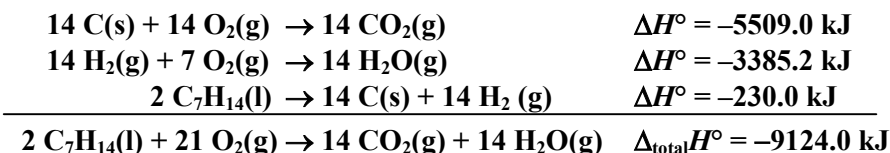
Calculate the standard enthalpy change for the following reaction.



3. Use the following reactions and enthalpy changes to predict the standard enthalpy change for:



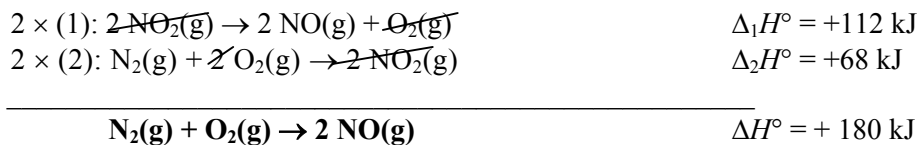
4. Use the following formation reaction evidence to calculate the standard enthalpy change for the complete combustion of cycloheptane.



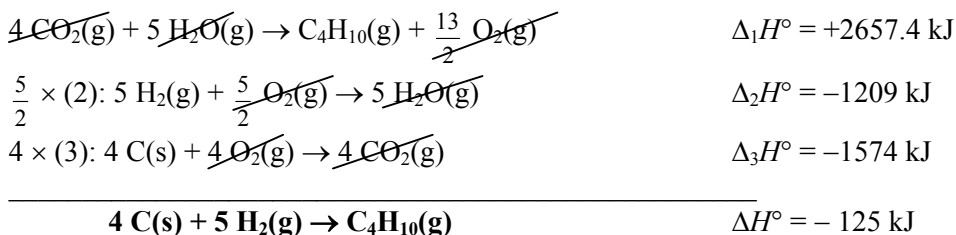
Student Worksheet Solutions LSM 11.4H

Solutions for Hess' Law, Extra Exercises

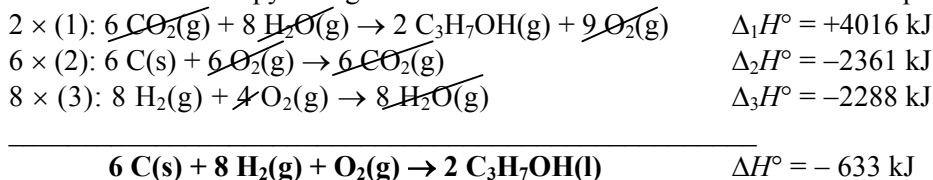
- (a) What is the enthalpy change for the formation of two moles of nitrogen monoxide from its elements?



- (b) What is the enthalpy change for the formation of one mole of butane, C₄H₁₀(g), from its elements?

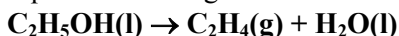


- (c) Determine the enthalpy change involved in the formation of two moles of liquid propanol.



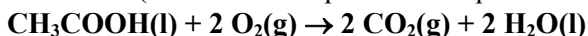
Solutions for Predicting ΔH Using Standard Enthalpies of Formation, Extra Exercises

1. Laboratory quantities of ethylene can be prepared by an elimination reaction of ethanol using an acid catalyst. Calculate the enthalpy change for the conversion of ethanol into ethylene and liquid water using standard enthalpies of formation.



$$\begin{aligned} \Delta_{\text{r}}H^{\circ} &= \sum n\Delta_{\text{fP}}H^{\circ} - \sum n\Delta_{\text{fR}}H^{\circ} \\ &= \left(1 \text{ mol} \times \frac{+52.4 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-277.6 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -233.4 \text{ kJ} - (-277.6 \text{ kJ}) \\ &= +44.2 \text{ kJ} \end{aligned}$$

2. Calculate the molar enthalpy of combustion for acetic acid using standard enthalpies of formation. (Assume that liquid water is produced.)



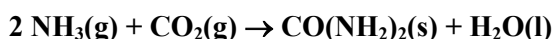
$$\begin{aligned} \Delta_{\text{r}}H^{\circ} &= \sum n\Delta_{\text{fP}}H^{\circ} - \sum n\Delta_{\text{fR}}H^{\circ} \\ &= \left(2 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-277.6 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -1358.6 \text{ kJ} - (-277.6 \text{ kJ}) \\ &= -1081.0 \text{ kJ} \end{aligned}$$

3. An initial step in the production of iron in a blast furnace involves the conversion of iron(III) oxide and carbon monoxide into iron(II, III) oxide, $\text{Fe}_3\text{O}_4(\text{s})$, and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.



$$\begin{aligned} \Delta_{\text{r}}H^{\circ} &= \sum n\Delta_{\text{fP}}H^{\circ} - \sum n\Delta_{\text{fR}}H^{\circ} \\ &= \left(2 \text{ mol} \times \frac{-1118.4 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} \right) - \left(3 \text{ mol} \times \frac{-824.2 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-110.5 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -2630.3 \text{ kJ} - (-277.6 \text{ kJ}) \\ &= -47.2 \text{ kJ} \end{aligned}$$

4. The fertilizer urea is produced along with liquid water by the reaction of ammonia and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.



$$\begin{aligned} \Delta_{\text{r}}H^{\circ} &= \sum n\Delta_{\text{fP}}H^{\circ} - \sum n\Delta_{\text{fR}}H^{\circ} \\ &= \left(1 \text{ mol} \times \frac{-333.5 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(2 \text{ mol} \times \frac{-45.9 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -619.3 \text{ kJ} - (-485.3 \text{ kJ}) \\ &= -134.0 \text{ kJ} \end{aligned}$$

Solutions for Multi-Step Energy Calculations, Extra Exercises

1. For the following combustion reaction, what mass of carbon dioxide is produced when 1500 kJ of energy is released?



$$\Delta_{\text{c}}H_{\text{mol}} = \frac{\Delta_{\text{c}}H}{n} = \frac{-2502 \text{ kJ}}{4 \text{ mol}} = -626 \text{ kJ/mol CO}_2$$

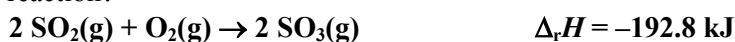
$$\Delta_{\text{c}}H = n\Delta_{\text{c}}H_{\text{mol}}$$

$$1500 \text{ kJ} = m \times \frac{1 \text{ mol}}{44.01 \text{ g}} \times 626 \text{ kJ/mol}$$

$$m = 105 \text{ g}$$

According to the evidence given, 106 g carbon dioxide will be produced for every 1500 kJ of energy released.

2. How much energy is released when 1.00 t of sulfur trioxide is produced by the following reaction?



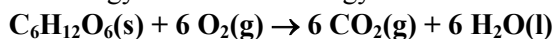
$$\Delta_{\text{r}}H_{\text{mol}} = \frac{\Delta_{\text{r}}H}{n} = \frac{-192.8 \text{ kJ}}{2 \text{ mol}} = -96.4 \text{ kJ/mol SO}_3$$

$$\Delta_{\text{r}}H = n\Delta_{\text{r}}H_{\text{mol}}$$

$$= 1.00 \text{ Mg} \times \frac{1 \text{ mol}}{80.06 \text{ g}} \times \frac{96.4 \text{ kJ}}{1 \text{ mol}} = 1.20 \text{ GJ}$$

According to the evidence given, 1.20 GJ of energy is released when 1.00 t of sulfur trioxide is produced.

3. In respiration, glucose is oxidized by oxygen gas to produce carbon dioxide gas, liquid water, and energy. What is the energy released when 18.0 g of glucose is consumed?



$$\Delta_{\text{r}}H^{\circ} = \sum n\Delta_{\text{fP}}H^{\circ} - \sum n\Delta_{\text{fR}}H^{\circ}$$

$$= \left(6 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-1273.3 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{-0 \text{ kJ}}{1 \text{ mol}} \right)$$

$$= -4075.8 \text{ kJ} - (-1273.3 \text{ kJ})$$

$$= -2802.5 \text{ kJ}$$

$$\Delta_{\text{c}}H_{\text{mol}} = \frac{-2802.5 \text{ kJ}}{1 \text{ mol}} = 2802.5 \text{ kJ/mol C}_6\text{H}_{12}\text{O}_6$$

$$\Delta_{\text{c}}H = n\Delta_{\text{c}}H_{\text{mol}}$$

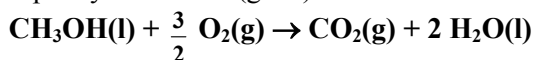
$$= 18.0 \text{ g} \times \frac{1 \text{ mol}}{180.18 \text{ g}} \times \frac{-2802.5 \text{ kJ}}{1 \text{ mol}}$$

$$= -280 \text{ kJ}$$

According to referenced information, 280 kJ of energy is released when 18.0 g of glucose is consumed.

LSM 11.5G (cont'd)

4. Methanol is burned in a bomb calorimeter. Liquid water is formed as a product. If 3.40 g of methanol reacts, what is the expected temperature change in a 6.00 kg calorimeter with a heat capacity of 1.125 J/(g·°C)?



$$\begin{aligned} \Delta_r H^\circ &= \sum n \Delta_{\text{fP}} H^\circ - \sum n \Delta_{\text{fR}} H^\circ \\ &= \left(1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-277.6 \text{ kJ}}{1 \text{ mol}} + \frac{3}{2} \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -965.1 \text{ kJ} - (-277.6 \text{ kJ}) \\ &= -687.5 \text{ kJ} \end{aligned}$$

$$\Delta_c H_{\text{mol}} = \frac{-687.5 \text{ kJ}}{1 \text{ mol}} = -687.5 \text{ kJ/mol CH}_3\text{OH}$$

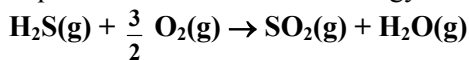
$$\begin{array}{l} \Delta_c H = mc\Delta t \\ \text{(methanol)} \quad \text{(calorimeter)} \end{array}$$

$$3.40 \text{ g} \times \frac{1 \text{ mol}}{32.05 \text{ g}} \times \frac{-687.5 \text{ kJ}}{1 \text{ mol}} = 1.00 \text{ kg} \times 1.125 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times \Delta t$$

$$\Delta t = 10.8 ^\circ\text{C}$$

According to referenced information and the evidence provided, the temperature of the calorimeter will change by 10.8 °C is released when 3.40 g of methanol is burned.

5. A waste heat exchanger is used to absorb the energy from the complete combustion of hydrogen sulfide gas. What volume of water undergoing a temperature change of 64 °C is required to absorb all of the energy from the burning of 15 kg of hydrogen sulfide?



$$\begin{aligned} \Delta_r H^\circ &= \sum n \Delta_{\text{fP}} H^\circ - \sum n \Delta_{\text{fR}} H^\circ \\ &= \left(1 \text{ mol} \times \frac{-296.8 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-241.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-20.6 \text{ kJ}}{1 \text{ mol}} + \frac{3}{2} \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -538.6 \text{ kJ} - (-20.6 \text{ kJ}) \\ &= -508.0 \text{ kJ} \end{aligned}$$

$$\Delta_c H_{\text{mol}} = \frac{-508.0 \text{ kJ}}{1 \text{ mol}} = -508.0 \text{ kJ/mol H}_2\text{S}$$

$$\begin{array}{l} \Delta_c H = Q \\ \text{(H}_2\text{S)} \quad \text{(water)} \end{array}$$

$$n \Delta_c H_{\text{mol}} = mc\Delta t$$

$$15 \text{ kg} \times \frac{1 \text{ mol}}{34.09 \text{ g}} \times \frac{-508.0 \text{ kJ}}{1 \text{ mol}} = m \times \frac{4.19 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times 64 ^\circ\text{C}$$

$$m = 834 \text{ kg}$$

$$V = 8.3 \times 10^2 \text{ L}$$

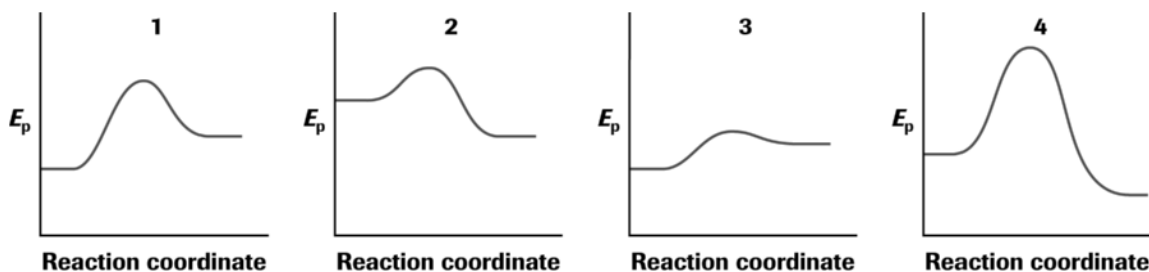
According to referenced information and the evidence provided, 8.3 × 10² L of water will warm by 64 °C when 15 kg of hydrogen sulfide is burned.

Student Worksheet Solutions LSM 11.CR

Chapter 11 Review, Part 1

1. B
2. 1, 4, 3, 2
3. C
4. 585
5. D
6. 98.9
7. C
8. D
9. 1, 3, 6
10. 2, 4, 5
11. 3, 4, 2, 1
12. 3, 2, 1, 4
13. 1, 2, 4, 3
14. 136

Solutions for Energy Changes During a Chemical Reaction, Extra Exercises



For each of the diagrams above, explain the *reasoning* in complete sentence statements, in terms of the changes in potential energy and kinetic energy over the duration of the reaction.

- In reaction 1, the activation energy is high, so the kinetic energy of the reactant molecules will likely not be sufficient to begin the reaction. Once underway, the increase in potential energy of this reaction results in kinetic (heat) energy being absorbed from the surroundings (in an endothermic reaction). Considerable energy would be required to keep this reaction going.**
- In reaction 2, the activation energy is low, so the kinetic energy of the reactant molecules will likely be sufficient to begin the reaction. Once underway, the decrease in potential energy of this reaction results in kinetic (heat) energy being released to the surroundings (in an exothermic reaction) to help maintain or even increase the rate of the reaction.**
- In reaction 3, the activation energy is low, so the kinetic energy of the reactant molecules will likely be sufficient to begin the reaction. Once underway, the increase in potential energy of this reaction results in kinetic (heat) energy being absorbed from the surroundings (in an endothermic reaction). An external energy source might be needed to keep the initially spontaneous reaction going.**
- In reaction 4, the activation energy is high, so the kinetic energy of the reactant molecules will likely not be sufficient to begin the reaction. Once underway, the decrease in potential energy of this reaction results in kinetic (heat) energy being released to the surroundings (in an exothermic reaction) to maybe make the initially nonspontaneous reaction self sustaining.**

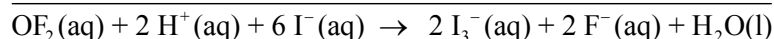
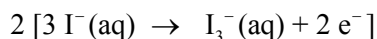
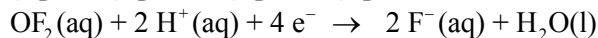
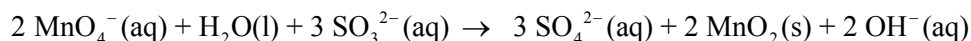
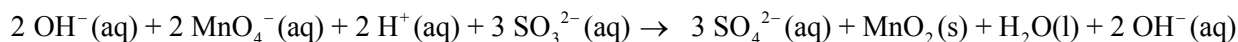
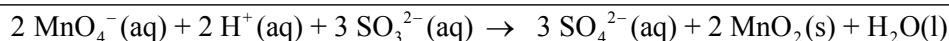
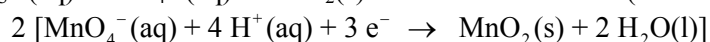
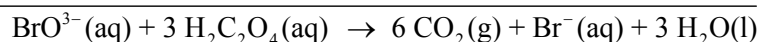
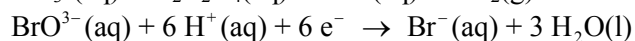
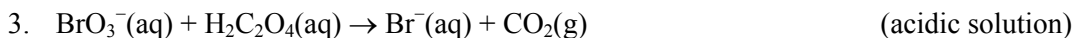
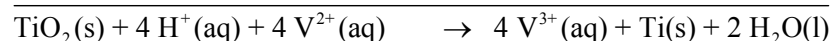
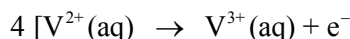
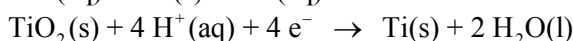
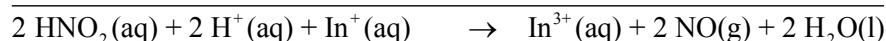
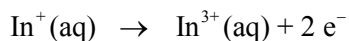
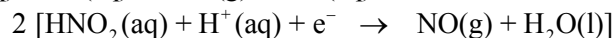
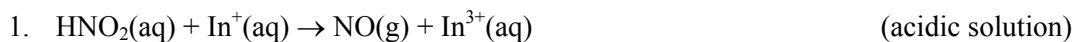
Student Worksheet Solutions LSM 12.CR
Chapter 12 Review, Part 1

1. C
2. A
3. D
4. 2, 6, 7, 8
5. A
6. D
7. B
8. D
9. B
10. B

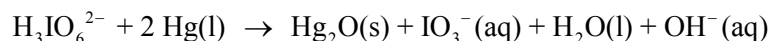
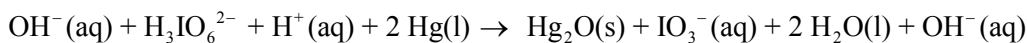
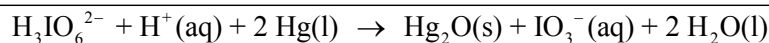
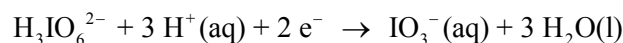
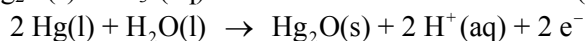
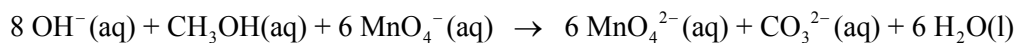
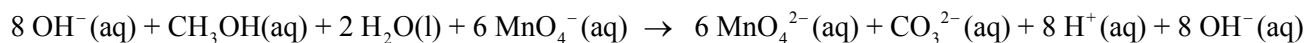
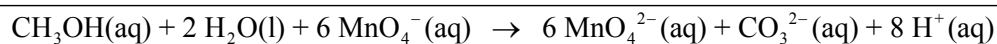
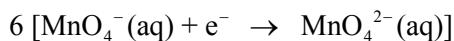
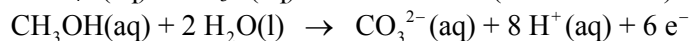
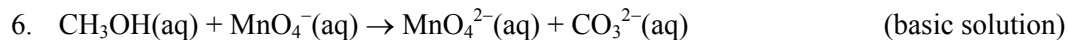
Student Worksheet Solutions LSM 12.UR
Unit 6 Review, Part 1

1. A
2. 572
3. 7.82
4. 1, 2, 3, 5
5. 1, 2, 4, 6
6. A
7. B
8. D
9. 1442
10. 749
11. A
12. D
13. 3, 3, 1, 2
14. 8, 9, 7, 6
15. D
16. 3, 2, 4, 1

Solutions for Balancing Redox Equations with Half-Reaction Equations, Extra Exercises

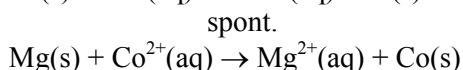
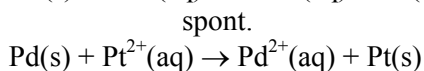
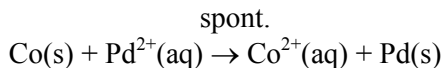


LSM13.1E (cont'd)

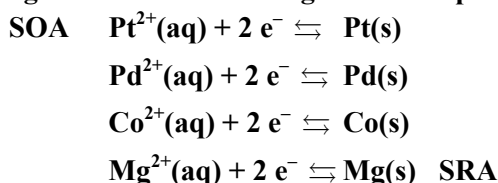


Solutions for Creating a Redox Table, Extra Exercises

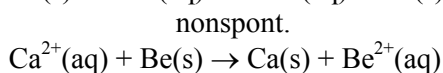
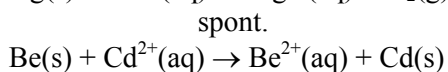
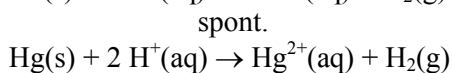
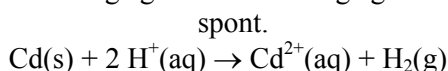
1. The following equations represent spontaneous reactions. From this evidence, set up a table of relative strengths of oxidizing and reducing agents. Write half-reaction equations and label the strongest oxidizing agent and reducing agent.



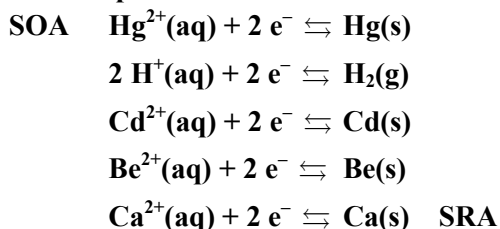
According to the evidence and the spontaneity rule, a table of oxidizing and reducing agents in order of strength is developed as follows:



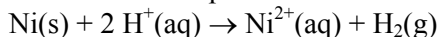
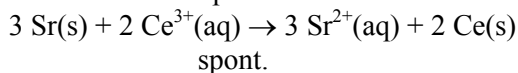
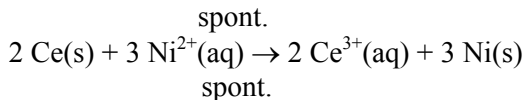
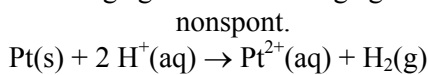
2. The following equations are interpretations of the evidence from the reactions of four metals with various cation solutions. Make a table of half-reactions and arrange the four metallic ions and the hydrogen ion in order of their decreasing tendency to react. Label the strongest oxidizing agent and reducing agent.



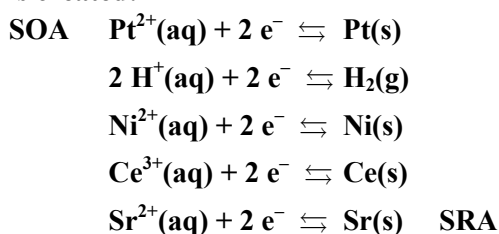
According to the evidence and the spontaneity rule, the following table of half-reactions is developed:



3. Four metals were placed into test tubes containing various ion solutions. Their resulting behaviour is communicated by the equations below. Create a half-reaction table and order the metallic ions and the hydrogen ion according to their tendency to react. Label the strongest oxidizing agent and reducing agent.



According to the evidence and the spontaneity rule, the following table of half-reactions is created:

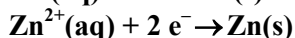
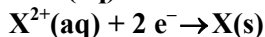
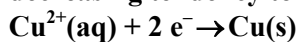


4. A student is required to store an aqueous solution of iron(III) nitrate. She has a choice of a copper, tin, iron, or silver container. Use the table of relative strengths of oxidizing and reducing agents in Appendix I and an appropriate generalization to predict which container would be most suitable for storing the solution.

Based upon the spontaneity rule, she should choose the silver container for storing the iron(III) nitrate solution. (This is, of course, from a scientific perspective. From an economic perspective, one would likely be concerned about the cost of using a silver container to store an inexpensive solution which could also be stored in glass.)

5. An analytical chemist reacts an unknown metal X with a copper(II) sulfate solution, plating out copper metal. Metal X does not react with aqueous zinc nitrate. What is the order for these metal ions in decreasing tendency to react? What groups of metals are eliminated as a possible identity of the unknown metal? What other solutions might next be chosen to help identify the unknown metal?

According to the evidence and the spontaneity rule, the metal ions in order of decreasing tendency to react are listed below:



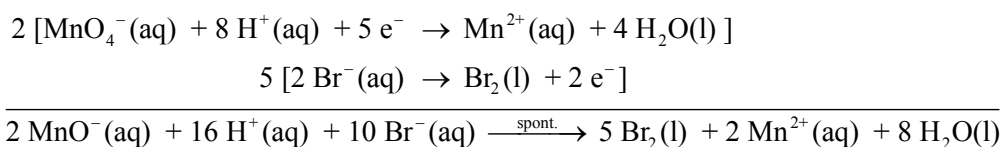
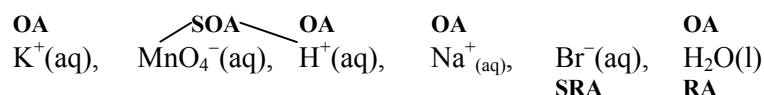
Metal X must be between copper and zinc in a table of half-reactions. This eliminates any Group 1, 2, and 3 metals as a possible identity for Metal X.

Groups 1, 2, and 3 metals are stronger reducing agents than zinc. The aqueous solutions of metal ions between copper and zinc ions in the table of relative strengths of oxidizing and reducing agents in Appendix G may be used to help narrow down the possibility of what Metal X may be. Some possible solutions that can be used are solutions of lead(II), nickel(II), cobalt(II), cadmium(II), chromium(III), and iron(III) aqueous ions.

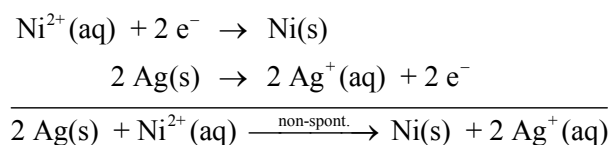
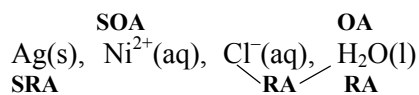
Solutions for Predicting Redox Reactions, Extra Exercises

For each of the following questions, use the five-step method to predict and communicate the most likely redox reaction, and use the spontaneity rule to predict whether the reaction will occur spontaneously.

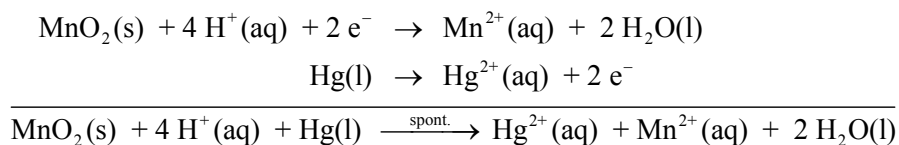
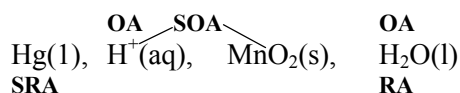
1. An aqueous solution of potassium permanganate was reacted with an acidic solution of sodium bromide and an orange-brown substance was formed.



2. A strip of silver metal is placed in a solution of aqueous nickel(II) chloride.

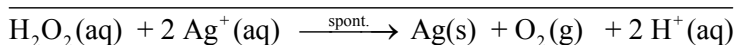
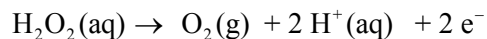
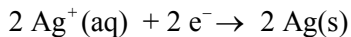
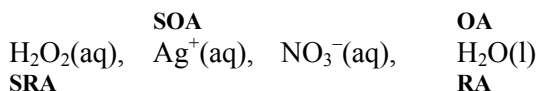


3. Liquid mercury is mixed with a paste of acidic manganese(IV) oxide.

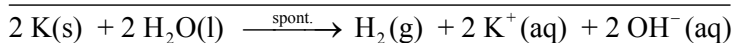
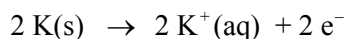
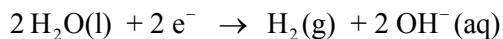
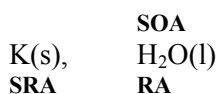


LSM13.2G (cont'd)

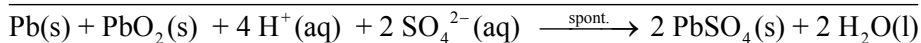
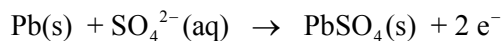
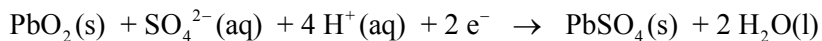
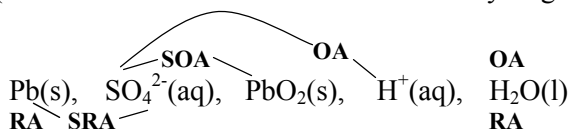
4. Hydrogen peroxide and silver nitrate are mixed.



5. Potassium metal is placed into some water.

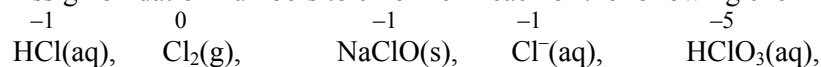


6. In a car battery, lead and lead(IV) oxide electrodes are exposed to a sulfuric acid electrolyte. (Assume that the sulfuric acid ionizes to hydrogen and sulfate ions.)

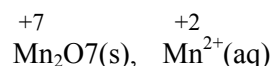
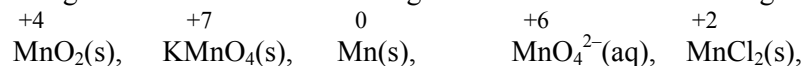


Solutions for Oxidation Numbers, Extra Exercises

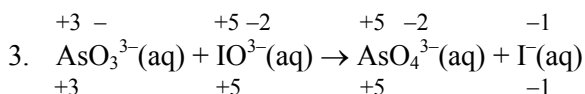
1. Assign oxidation numbers to chlorine in each of the following chemicals.



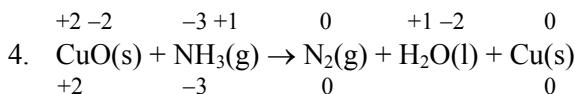
2. Assign oxidation numbers to manganese in each of the following chemicals.



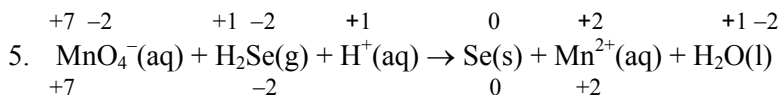
For the following reaction equations, use oxidation numbers to identify the oxidation (the atom oxidized) and the reduction (the atom reduced).



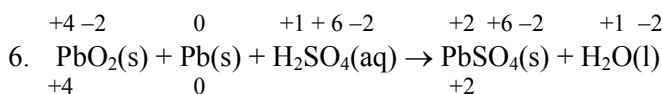
As is oxidized; I is reduced



N is oxidized; Cu is reduced

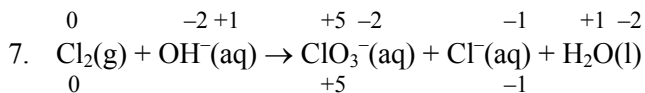


Se is oxidized; Mn is reduced



Pb is oxidized; Pb is reduced

(Note that the lead(IV) ions in PbO_2 and the lead atoms in the metal element are both converted into lead(II) ions in the course of this reaction, which represents a car battery discharging.)

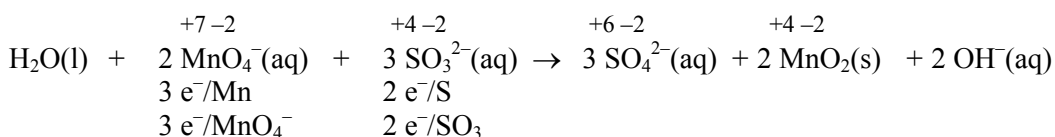
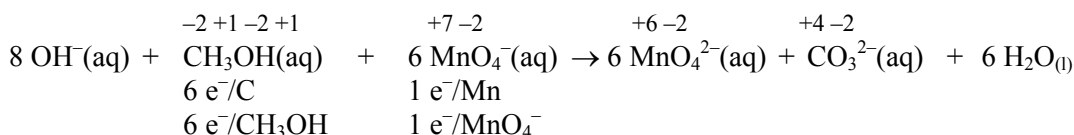
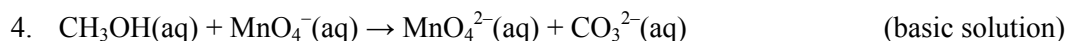
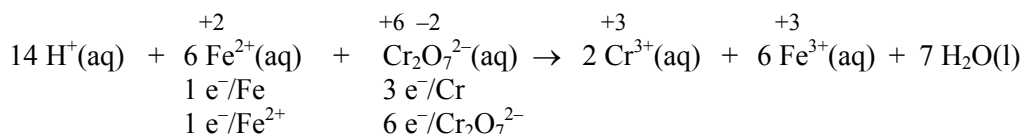
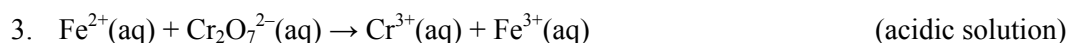
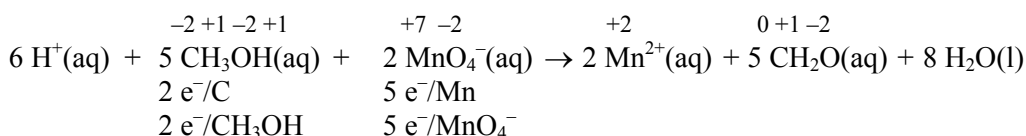
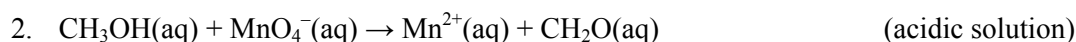
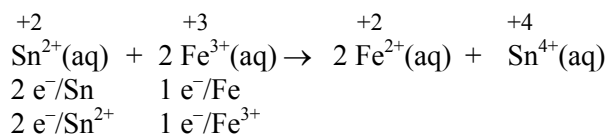
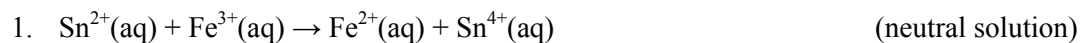


Cl is oxidized; Cl is reduced

(Note that some of the chlorine gas atoms are oxidized while others are reduced.)

Solutions for Balancing Redox Equations Using Oxidation Numbers, Extra Exercises

Balance the following reaction equations.



Student Worksheet Solutions LSM 13.CR

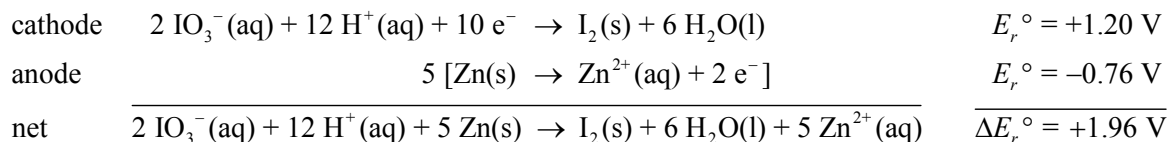
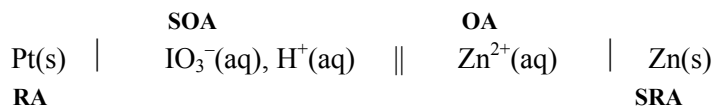
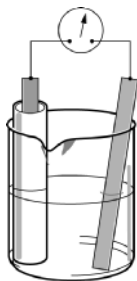
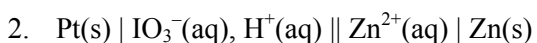
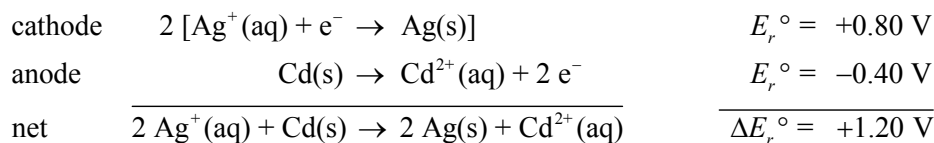
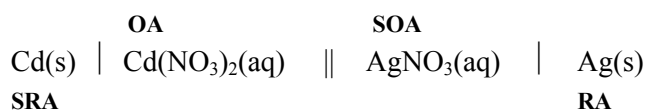
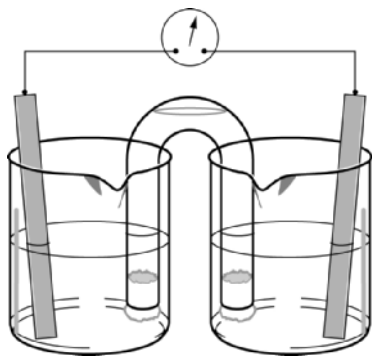
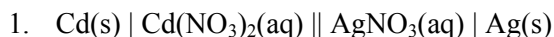
Chapter 13 Review, Part 1

1. B
2. A
3. D
4. 1, 3, 5, 6
5. B
6. D
7. C
8. C
9. A
10. 3, 1, 1, 3
11. B
12. B
13. A
14. D
15. 65.6

Student Worksheet Solutions LSM 14.2H

Solutions for Voltaic Cells, Extra Exercises

For each of the following cells, use the given cell notation to identify the strongest oxidizing and reducing agents. Write chemical equations to represent the cathode, anode, and net cell reactions. Label electrodes, electrolytes, electron flow, and ion movement. Predict the cell potential.

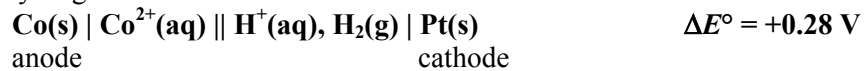


Solutions for Cell Potentials, Extra Exercises

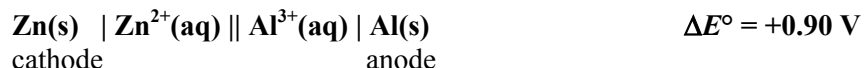
1. Assume that the reference half-cell is changed to a standard mercury–mercury(II) half-cell.
- (a) What would be the reduction potential of a standard chlorine half-cell?
+0.51 V
- (b) What would be the reduction potential of a standard nickel half-cell?
–1.11 V
- (c) What would be the net cell potential of a standard chlorine–nickel cell?
+1.62 V
- (d) Why is the answer to part c the same as the answer obtained using the standard hydrogen half-cell as the reference?
A cell potential is the difference in reduction potentials. This difference (how far apart the half-cells are in the table) does not depend on the individual reduction potentials.

2. For each of the following standard cells, write the cell notation, label electrodes, and determine the cell potential.

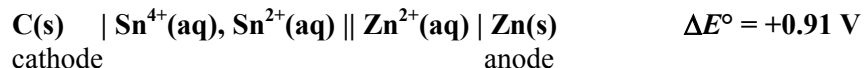
- (a) cobalt–hydrogen standard cell



- (b) zinc–aluminium standard cell

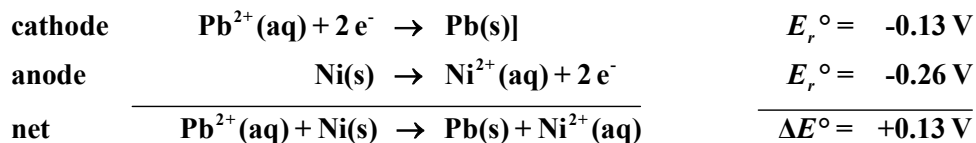


- (c) tin(IV)–zinc standard cell

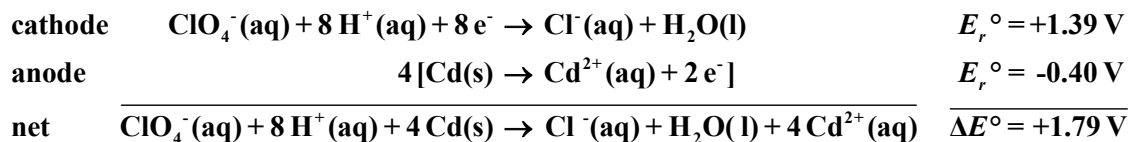
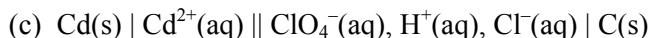
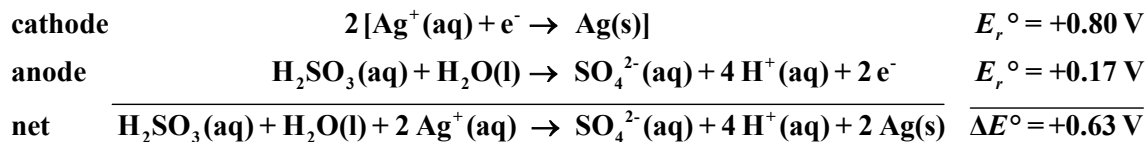
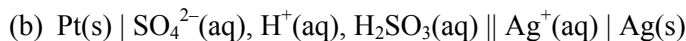


3. For each of the following cell notations, write the cathode, anode, and net cell reaction equations and calculate the cell potential.

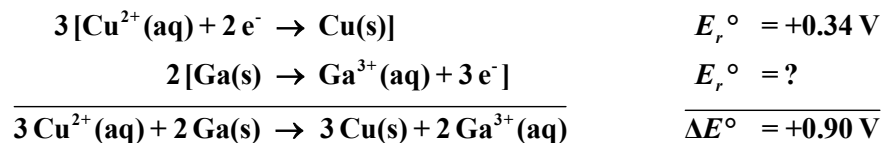
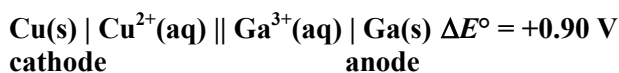
- (a) $\text{Pb(s)} \mid \text{Pb}^{2+}(\text{aq}) \parallel \text{Ni}^{2+}(\text{aq}) \mid \text{Ni(s)}$



LSM 14.2J (cont'd)



4. Use the standard cell described below to determine the standard reduction potential of the gallium half-cell.



$$\Delta E^\circ = E_r^\circ - E_r^\circ$$

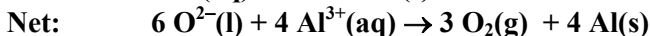
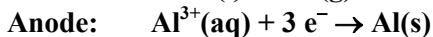
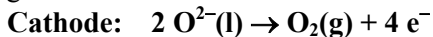
cathode anode

$$+0.90 \text{ V} = +0.34 \text{ V} - E_r^\circ$$

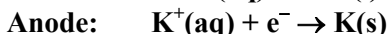
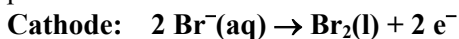
$$E_r^\circ = -0.56 \text{ V}$$

Solutions for Electrolytic Cells, Extra Exercises

1. Write the equations for the anode, cathode, and net reactions for the electrolysis of molten aluminium oxide. One half-reaction will involve conversion of liquid oxide ions into oxygen gas.



2. Write equations for the anode, cathode, and net reactions for the electrolysis of molten potassium bromide.



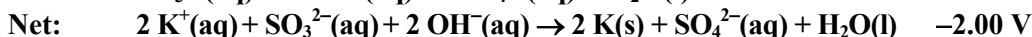
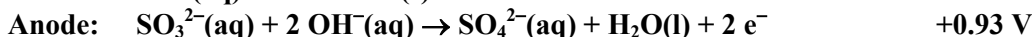
3. For the following electrolytic cells, write equations for half-reactions and the net cell reaction, and calculate the minimum electrical potential difference that would be required to cause the electrolysis to occur.

- (a) an aqueous solution of nickel (II) chloride is electrolyzed



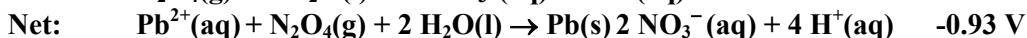
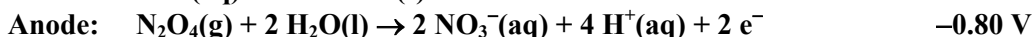
Electrical potential > 1.10 V required

- (b) an aqueous solution of potassium sulfate is electrolyzed



Electrical potential > 2.00 V required

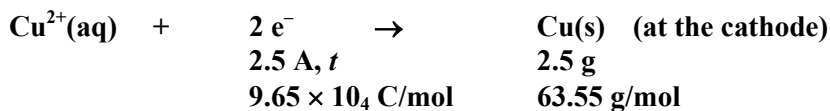
- (c) an aqueous solution of lead (II) nitrate is electrolyzed



Electrical potential > 0.93 V required

Solutions for Stoichiometry of Cell Reactions, Extra Exercises

1. A student wishes to set up an electrolytic cell to plate copper onto a belt buckle. Predict the length of time it will take to plate out 2.5 g of copper from a copper(II) nitrate solution using a 2.5 A current. At which electrode should the buckle be attached?



$$n_{\text{Cu}} = 2.5 \text{ g} \times \frac{1 \text{ mol}}{63.55 \text{ g}}$$

$$n_{\text{Cu}} = 0.039 \text{ mol}$$

$$n_{\text{e}^-} = 0.039 \text{ mol Cu} \times \frac{2 \text{ mol e}^-}{1 \text{ mol Cu}}$$

$$n_{\text{e}^-} = 0.079 \text{ mol}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

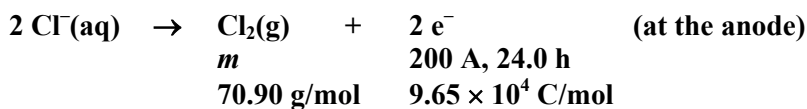
$$t = \frac{nF}{I}$$

$$= \frac{0.079 \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{2.5 \frac{\text{C}}{\text{s}}}$$

$$t = 3.0 \times 10^3 \text{ s} \quad (\text{or } 51 \text{ min})$$

The time required is 3.0 ks (or 51 mins). The buckle should be attached to the cathode.

2. Determine the mass of chlorine produced when a 200 A current flows for 24.0 h through a cell containing molten sodium chloride (a Downs cell). At which electrode is the chlorine produced?



$$n_{\text{e}^-} = \frac{It}{F}$$

$$= \frac{200 \frac{\text{C}}{\text{s}} \times 24.0 \text{ h} \times 3600 \frac{\text{s}}{\text{h}}}{9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}$$

$$n_{\text{e}^-} = 179 \text{ mol}$$

$$n_{\text{Cl}_2} = 179 \text{ mol e}^- \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol e}^-}$$

$$n_{\text{Cl}_2} = 89.5 \text{ mol}$$

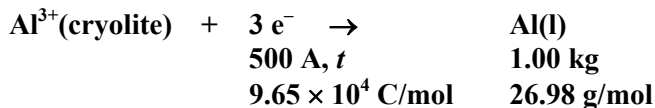
$$m_{\text{Cl}_2} = 89.5 \text{ mol} \times \frac{70.90 \text{ g}}{1 \text{ mol}}$$

LSM 14.4D (cont'd)

$$m_{\text{Cl}_2} = 6.35 \text{ kg}$$

The mass of chlorine produced is 6.35 kg at the anode of the cell.

3. How long would it take a 500 A current to produce 1.00 kg of aluminium from aluminium oxide dissolved in molten cryolite (Hall-Héroult cell)?



$$n_{\text{Al}} = 1.00 \text{ kg} \times \frac{1 \text{ mol}}{26.98 \text{ g}}$$

$$n_{\text{Al}} = 0.0371 \text{ kmol}$$

$$n_{\text{e}^-} = 0.03371 \text{ kmol Al} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Al}}$$

$$n_{\text{e}^-} = 0.111 \text{ kmol}$$

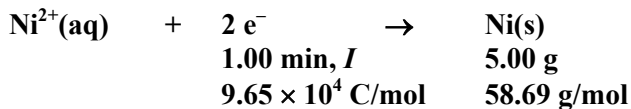
$$n_{\text{e}^-} = \frac{It}{F}$$

$$t = \frac{nF}{I}$$

$$= \frac{0.111 \text{ kmol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{500 \frac{\text{C}}{\text{s}}}$$

$$t = 21.5 \text{ ks} \quad (\text{or } 358 \text{ min})$$

4. A trophy company is setting up a nickel-plating cell using an electrolyte containing nickel(II) ions. Predict the current required to produce nickel metal at the rate of 5.00 g/min.



$$n_{\text{Ni}} = 5.00 \text{ g} \times \frac{1 \text{ mol}}{58.69 \text{ g}}$$

$$n_{\text{Ni}} = 0.0852 \text{ mol}$$

$$n_{\text{e}^-} = 0.0852 \text{ mol Ni} \times \frac{2 \text{ mol e}^-}{1 \text{ mol Ni}}$$

$$n_{\text{e}^-} = 0.170 \text{ mol}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

$$t = \frac{nF}{I}$$

$$= \frac{0.170 \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{1.00 \text{ min} \times 60 \frac{\text{s}}{\text{min}}}$$

$$I = 274 \text{ A}$$

The current required is 274 A.

Student Worksheet Solutions LSM 14.CR

Chapter 14 Review, Part 1

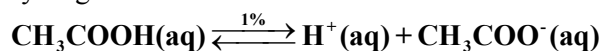
1. A
2. B
3. D
4. C
5. 1, 3, 4, 7
6. 2, 3, 5, 7
7. 1.49
8. B
9. D
10. A
11. A
12. C
13. 1.16
14. 1.51

Student Worksheet Solutions LSM 14.UR
Unit 7 Review, Part 1

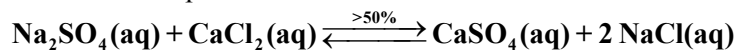
1. B
2. A
3. 3, 1, 4, 2
4. C
5. A
6. D
7. B
8. C
9. 2, 1, 3, 4
10. D
11. 0.85
12. 87.4
13. B
14. A
15. D
16. C
17. B
18. C
19. D
20. 0.65

Solutions for Chemical Equilibrium, Extra Exercises

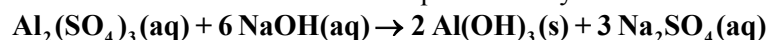
1. For each of the following, write the chemical reaction equation with appropriate equilibrium arrows.
- (a) pH measurements indicate that acetic acid in vinegar is approximately 1% ionized into hydrogen ions and acetate ions.



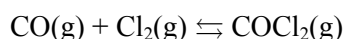
- (b) Quantitative analysis of the reaction of sodium sulfate and calcium chloride solutions shows that the products are favoured



- (c) Aluminium sulfate solution reacts quantitatively with a sodium hydroxide solution.



2. Chlorine and carbon monoxide gases are mixed in a 1.00 L container and the following equilibrium is established



Initially, 1.50 mol of chlorine was present with 2.55 mol of carbon monoxide. At equilibrium, 0.80 mol of $\text{COCl}_2(\text{g})$ was found.

- (a) Create an ICE table and calculate the percent reaction of chlorine.

Concentration	[CO(g)] (mol/L)	[Cl ₂ (g)] (mol/L)	[COCl ₂ (g)] (mol/L)
Initial	2.55	1.50	0.00
Change	0.80	0.80	0.80
Equilibrium	1.75	0.75	0.80

$$\% \text{ reaction} = \frac{0.80 \text{ mol}}{1.50 \text{ mol}} \times 100 = 53\%$$

- (b) Write the equilibrium law for this reaction.

$$K_c = \frac{[\text{COCl}_2(\text{g})]}{[\text{CO}(\text{g})][\text{Cl}_2(\text{g})]}$$

- (c) At equilibrium, 1.75 mol of carbon monoxide and 0.70 mol of chlorine were present. Calculate the equilibrium constant.

$$\begin{aligned} K_c &= \frac{0.80}{(1.75)(0.70)} \\ &= 0.65 \end{aligned}$$

3. Write the equilibrium law for each of the following chemical reaction equations.

- (a) $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$

$$K_c = \frac{[\text{SO}_3(\text{g})]^2}{[\text{SO}_2(\text{g})]^2 [\text{O}_2(\text{g})]}$$

- (b) $2 \text{NO}_2(\text{g}) \rightleftharpoons 2 \text{NO}(\text{g}) + \text{O}_2(\text{g})$

$$K_c = \frac{[\text{NO}(\text{g})]^2 [\text{O}_2(\text{g})]}{[\text{NO}_2(\text{g})]^2}$$

LSM 15.1E (cont'd)

- (c) At equilibrium, 1.75 mol of carbon monoxide and 0.70 mol of chlorine were present. Calculate the equilibrium constant.

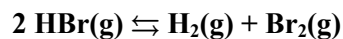
$$K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3}$$

4. In an experiment at a high temperature, 0.500 mol/L of hydrogen bromide gas is placed into a sealed container and decomposes into hydrogen and bromine gases.
- (a) Write the equilibrium equation and law for this reaction.



$$K = \frac{[\text{H}_2(\text{g})][\text{Br}_2(\text{g})]}{[\text{HBr}(\text{g})]^2}$$

- (b) The equilibrium concentration for $\text{Br}_2(\text{g})$ in this system is found from calorimetry to be 0.130 mol/L. Use an ICE table to help calculate the equilibrium constant.



Concentration	[HBr(g)] (mol/L)	[H ₂ (g)] (mol/L)	[Br ₂ (g)] (mol/L)
Initial	0.500	0.000	0.000
Change	0.260	0.130	0.130
Equilibrium	0.240	0.130	0.130

$$K_c = \frac{(0.130)^2}{(0.240)^2} = 0.293$$

Solutions for Le Châtelier's Principle, Extra Exercises

- Nitrogen monoxide, a major air pollutant, is formed in automobile engines from the endothermic reaction of nitrogen and oxygen gases.

 - Write the equilibrium reaction equation including the term "energy" in the equation.

$$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) + \text{energy} \rightleftharpoons 2 \text{NO}(\text{g})$$
 - Describe the direction of the equilibrium shift if the concentration of oxygen is increased.
The equilibrium shifts to the right.
 - Describe the direction of the equilibrium shift if the pressure is increased.
There will be no shift in the equilibrium because the number of gas molecules is the same on both sides of the equilibrium equation.
 - Gasoline burns better at higher temperatures. What are some disadvantages of the operation of automobile engines at higher temperatures?
At high temperatures, the equilibrium shifts to the right, producing more of the pollutant nitrogen monoxide. There may also be additional technical and safety problems as the temperature is increased.

- In a sealed container, nitrogen monoxide and oxygen gases are in equilibrium with nitrogen dioxide gas. The reaction of nitrogen monoxide and oxygen is exothermic.

$$2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g}) + \text{energy}$$

Predict the equilibrium shift when the following changes are made.

 - the temperature is decreased
The equilibrium shifts to the right.
 - the concentration of $\text{NO}_{(\text{g})}$ is decreased
The equilibrium shifts to the left.
 - the concentration of $\text{NO}_{2(\text{g})}$ is increased
The equilibrium shifts to the left.
 - the volume of the system is decreased
The equilibrium shifts to the right.

- The equilibrium of the iron(III)–thiocyanate system is convenient to study.

$$\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightleftharpoons \text{FeSCN}^{2+}(\text{aq})$$

almost colourless
colourless
red

Predict the colour change in the equilibrium mixture when each of the following changes is made.

 - A crystal of $\text{KSCN}(\text{s})$ is added to the system.
The red colour becomes more intense as the equilibrium shifts to the right.
 - A crystal of $\text{CeCl}_3(\text{s})$ is added to the system.
The red colour becomes more intense as the equilibrium shifts to the right.
 - A crystal of $\text{NaOH}(\text{s})$ is added to the system.
The red colour becomes less intense as the equilibrium shifts to the left. (Iron(III) ions precipitate as iron(III) hydroxide.)

Student Worksheet Solutions LSM 15.CR

Chapter 15 Review, Part 1

1. B
2. C
3. 70.3
4. D
5. A
6. B
7. C
8. B
9. A
10. 0.400

Student Worksheet Solutions LSM 16.1D

Solutions for K_w , pH, and pOH Calculations, Extra Exercises

1. Calculate the $[\text{OH}^-(\text{aq})]$ in limes which have a $[\text{H}_3\text{O}^+(\text{aq})]$ of $1.3 \times 10^{-2} \text{ mol/L}$.

$$[\text{OH}^-(\text{aq})] = \frac{K_w}{[\text{H}_3\text{O}^+(\text{aq})]} = \frac{1.0 \times 10^{-14} (\text{mol/L})^2}{1.3 \times 10^{-2} \text{ mol/L}} = 7.7 \times 10^{-13} \text{ mol/L}$$

2. Calculate the $[\text{H}_3\text{O}^+(\text{aq})]$ in lemons which have a $[\text{OH}^-(\text{aq})]$ of $2.0 \times 10^{-12} \text{ mol/L}$.

$$[\text{H}_3\text{O}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]} = \frac{1.0 \times 10^{-14} (\text{mol/L})^2}{2.0 \times 10^{-12} \text{ mol/L}} = 5.0 \times 10^{-3} \text{ mol/L}$$

3. A sodium hydroxide solution is prepared by dissolving 2.50 g to make 2.00 L of solution. Calculate the hydroxide and hydronium ion concentrations.

$$n_{\text{NaOH}} = 2.50 \text{ g} \times \frac{1 \text{ mol}}{40.00 \text{ g}} = 6.25 \times 10^{-2} \text{ mol}$$

$$[\text{NaOH}(\text{aq})] = \frac{6.25 \times 10^{-2} \text{ mol}}{2.00 \text{ L}} = 3.13 \times 10^{-2} \text{ mol/L}$$

$$[\text{OH}^-(\text{aq})] = [\text{NaOH}(\text{aq})] = 3.13 \times 10^{-2} \text{ mol/L}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]} = \frac{1.0 \times 10^{-14} (\text{mol/L})^2}{3.13 \times 10^{-2} \text{ mol/L}} = 3.2 \times 10^{-13} \text{ mol/L}$$

4. A 0.728 g sample of hydrogen chloride gas is dissolved in 200 mL of solution. Calculate the hydronium and hydroxide ion concentrations.

$$n_{\text{HCl}} = 0.728 \text{ g} \times \frac{1 \text{ mol}}{36.46 \text{ g}} = 2.00 \times 10^{-2} \text{ mol}$$

$$[\text{HCl}(\text{aq})] = \frac{2.00 \times 10^{-2} \text{ mol/L}}{0.200 \text{ L}} = 9.98 \times 10^{-2} \text{ mol/L}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = [\text{HCl}(\text{aq})] = 9.98 \times 10^{-2} \text{ mol/L}$$

$$[\text{OH}^-(\text{aq})] = \frac{K_w}{[\text{H}_3\text{O}^+(\text{aq})]} = \frac{1.0 \times 10^{-14} (\text{mol/L})^2}{9.98 \times 10^{-2} \text{ mol/L}} = 1.0 \times 10^{-13} \text{ mol/L}$$

5. Vinegar has a hydronium ion concentration of $1.5 \times 10^{-3} \text{ mol/L}$. Calculate the pH.

$$\text{pH} = -\log[\text{H}_3\text{O}^+(\text{aq})] = -\log(1.5 \times 10^{-3}) = 2.82$$

6. An ammonia solution has a pOH of 2.92. What is the concentration of hydroxide ions in the solution?

$$[\text{OH}^-(\text{aq})] = 10^{-\text{pOH}} = 10^{-2.92} \text{ mol/L} = 1.2 \times 10^{-3} \text{ mol/L}$$

LSM 16.1D (cont'd)

7. Calculate the pOH and pH of a solution made by dissolving 7.50 g of strontium hydroxide to make 500 mL of solution.

$$n_{\text{Sr(OH)}_2} = 7.50 \text{ g} \times \frac{1 \text{ mol}}{121.64 \text{ g}} = 0.0617 \text{ mol}$$

$$[\text{Sr(OH)}_2(\text{aq})] = \frac{0.0617 \text{ mol}}{0.500 \text{ L}} = 0.123 \text{ mol/L}$$

$$[\text{OH}^-(\text{aq})] = 2[\text{Sr(OH)}_2(\text{aq})] = 2 \times 0.123 \text{ mol/L} = 0.247 \text{ mol/L}$$

$$\text{pOH} = -\log[\text{OH}^-(\text{aq})] = -\log(0.247) = 0.608$$

$$\text{pH} = 14.000 - 0.608 = 13.392$$

8. Complete the following table.

Table 1 pH of Common Substances

Substance	$[\text{H}_3\text{O}^+(\text{aq})]$ (mol/L)	pH	$[\text{OH}^-(\text{aq})]$ (mol/L)	pOH	Acidic, basic, or neutral
milk	3.1×10^{-7}	6.51	3.2×10^{-4}	7.49	acidic
pure water	1×10^{-7}	7.0	1×10^{-7}	7.0	neutral
blood	4.0×10^{-8}	7.40	2.5×10^{-7}	6.60	basic
cleaner	1.6×10^{-11}	10.80	6.3×10^{-4}	3.20	basic

Solutions for Brønsted–Lowry Definitions, Extra Exercises

- According to the Brønsted–Lowry concept, define an acid and a base in terms of an acid–base reaction.
According to the Brønsted–Lowry concept, a proton is transferred from an acid to a base in an acid–base reaction.
- What is meant by an “acidic” solution?
An acidic solution is an empirical classification of a solution with properties such as turning blue litmus red and having a pH less than 7.
 - Does a Brønsted–Lowry acid have to form an acidic solution? Explain.
No. A Brønsted–Lowry acid is defined in terms of proton transfer in a reaction and not in terms of the empirical properties of a solution. For example, the hydrogen carbonate ion forms a basic solution, but sometimes acts as a Brønsted–Lowry acid.
- There are many species that are classified as amphiprotic.
 - What does amphiprotic mean?
An amphiprotic substance is one that appears to act as a Brønsted–Lowry acid in some reactions and as a Brønsted–Lowry base in other reactions.
 - What general type of species is amphiprotic?
Hydrogen-polyatomic species, such as $\text{HCO}_3^-(\text{aq})$, are generally amphiprotic.
- $$\text{HOCCOO}^-(\text{aq}) + \text{PO}_4^{3-}(\text{aq}) \rightleftharpoons \text{HPO}_4^{2-}(\text{aq}) + \text{OCCOO}^{2-}(\text{aq})$$

A
B
A
B

 - Label each species as A or B for both forward and reverse reactions.
 - Identify all conjugate acid–base pairs.
 **$\text{HOCCOO}^-(\text{aq}) / \text{OCCOO}^{2-}(\text{aq})$
 $\text{HPO}_4^{2-}(\text{aq}) / \text{PO}_4^{3-}(\text{aq})$**
 - According to the Brønsted–Lowry concept, what determines the position of this equilibrium?
According to the Brønsted–Lowry concept, the position of the equilibrium is determined by the competition for proton transfer between the reacting species in the forward reaction and the reacting species in the reverse reaction.
- What is the generalization about the strength of an acid relative to its conjugate base?
In general, the stronger an acid, the weaker its conjugate base; and conversely, the weaker the acid, the stronger its conjugate base.

Solutions for Strengths of Acids, Extra Exercises

1. List three empirical properties that may be measured to distinguish among acids of different strengths.

conductivity, rate of reaction, pH

2. Calculate the hydronium ion concentration and pH of a 0.10 mol/L solution of nitrous acid.

$$[\text{H}_3\text{O}^+(\text{aq})] = \frac{8.1}{100} \times 0.10 \text{ mol/L} = 8.1 \times 10^{-3} \text{ mol/L or}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+(\text{aq})] = -\log(8.1 \times 10^{-3}) = 2.09$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \sqrt{\frac{7.2 \times 10^{-4} \text{ mol}}{1 \text{ L}} \times \frac{0.10 \text{ mol}}{1 \text{ L}}}$$

$$= 8.5 \times 10^{-3} \text{ mol/L}$$

$$\text{pH} = -\log(8.5 \times 10^{-3}) = 2.07$$

3. Calculate the hydrogen ion concentration and pH of a solution prepared by dissolving 10.70 g of ammonium chloride to make 2.00 L of solution.

$$n_{\text{NH}_4\text{Cl}} = 10.70 \text{ g} \times \frac{1 \text{ mol}}{53.50 \text{ g}} = 0.2000 \text{ mol}$$

$$[\text{NH}_4\text{Cl}(\text{aq})] = \frac{0.2000 \text{ mol}}{2.00 \text{ L}} = 0.100 \text{ mol/L}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \frac{0.0076}{100} \times 0.100 \text{ mol/L} = 7.6 \times 10^{-6} \text{ mol/L or}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+(\text{aq})] = -\log(7.6 \times 10^{-6}) = 5.12$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \sqrt{\frac{5.8 \times 10^{-10} \text{ mol}}{1 \text{ L}} \times \frac{0.100 \text{ mol}}{1 \text{ L}}}$$

$$= 7.6 \times 10^{-6} \text{ mol/L}$$

$$\text{pH} = -\log(7.6 \times 10^{-6}) = 5.12$$

4. Use the percent reaction value to determine the mass of sodium hydrogen sulfate required to prepare 500 mL of solution with a pH of 1.57.

$$[\text{H}_3\text{O}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-1.57} \text{ mol/L} = 2.7 \times 10^{-2} \text{ mol/L or}$$

$$\frac{[\text{H}_3\text{O}^+(\text{aq})]}{[\text{HSO}_4^-(\text{aq})]} \times 100 = 27\%$$

$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})]^2}{[\text{HSO}_4^-(\text{aq})] - [\text{H}^+(\text{aq})]}$$

$$= 1.0 \times 10^{-2} \text{ mol/L}$$

$$[\text{HSO}_4^-(\text{aq})] = \frac{2.7 \times 10^{-2} \text{ mol/L}}{27} \times 100 = 0.10 \text{ mol/L}$$

$$[\text{NaHSO}_4(\text{aq})] = [\text{HSO}_4^-(\text{aq})] = 0.10 \text{ mol/L}$$

LSM 16.3E (cont'd)

$$[\text{HSO}_4^-(\text{aq})] = \frac{(2.7 \times 10^{-2} \text{ mol/L})^2}{1.0 \times 10^{-2} \text{ mol/L} + 2.7 \times 10^{-2} \text{ mol/L}}$$

$$= 9.9 \times 10^{-2} \text{ mol/L}$$

$$n_{\text{NaHSO}_4} = 500 \text{ mL} \times \frac{0.10 \text{ mol}}{1 \text{ L}} = 50 \text{ mmol}$$

$$m_{\text{NaHSO}_4} = 50 \text{ mmol} \times \frac{120.06 \text{ g}}{1 \text{ mol}} = 6.0 \text{ g}$$

$$[\text{NaHSO}_4(\text{aq})] = [\text{HSO}_4^-(\text{aq})] = 9.9 \times 10^{-2} \text{ mol/L}$$

$$n_{\text{NaHSO}_4} = 500 \text{ mL} \times \frac{9.9 \times 10^{-2} \text{ mol}}{1 \text{ L}} = 50 \text{ mmol}$$

$$m_{\text{NaHSO}_4} = 50 \text{ mmol} \times \frac{120.06 \text{ g}}{1 \text{ mol}} = 6.0 \text{ g}$$

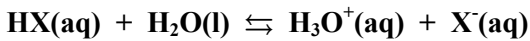
5. A 0.80 mol/L solution of an unknown acid, HX(aq), has a pH of 3.75.

(a) Calculate the percent reaction.

$$[\text{H}_3\text{O}^+(\text{aq})] = 10^{-\text{pH}} = 10^{-3.75} \text{ mol/L} = 1.8 \times 10^{-4} \text{ mol/L}$$

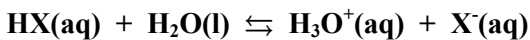
$$p = \frac{1.8 \times 10^{-4} \text{ mol/L}}{0.80 \text{ mol/L}} \times 100 = 2.2 \times 10^{-2}\%$$

(b) Calculate the acid ionization constant.



$$K = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{X}^-(\text{aq})]}{[\text{HX}(\text{aq})]} = \frac{(1.8 \times 10^{-4} \text{ mol/L})^2}{0.80 \text{ mol/L}} = 4.0 \times 10^{-8} \text{ mol/L}$$

6. Calculate the pH of a solution containing 0.25 mol/L of an acid with an acid ionization constant of 3.2×10^{-6} mol/L.



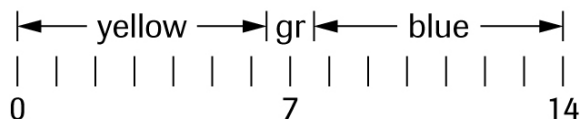
$$K = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{X}^-(\text{aq})]}{[\text{HX}(\text{aq})]}$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \sqrt{3.2 \times 10^{-6} \text{ mol/L} \times 0.25 \text{ mol/L}} = 8.9 \times 10^{-4} \text{ mol/L}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+(\text{aq})] = -\log(8.9 \times 10^{-4}) = 3.05$$

Solutions for Indicators, Extra Exercises

1. Use the following pH scale to label the colours for bromothymol blue over the 0–14 pH range. Identify the form of the indicator for each distinct colour using conventional symbols.



yellow: ≤ 6 ; HBb(aq)
green: in 6–7.6 range; HBb(aq) and $\text{Bb}^{\text{-}}(\text{aq})$
blue: ≥ 7.6 ; $\text{Bb}^{\text{-}}(\text{aq})$

2. **Problem**

What is the approximate pH of an unknown solution?

Evidence

Separate samples of the unknown solution turned blue litmus to red, congo red to blue, and orange IV to yellow.

Analysis

If litmus turns red, the pH is 6.0 or lower.

If congo red turns blue, the pH is 3.0 or lower.

If orange IV turns yellow, the pH is 2.8 or higher.

According to the evidence, the approximate pH of the unknown solution is between 2.8 and 3.0

3. Three unknown solutions in unlabelled beakers have pH values of 5.8, 7.8, and 9.8. Write two diagnostic tests using indicators to identify the pH of each solution.

If all three solutions are tested with red and blue litmus paper, the solution that turns the blue litmus to red has a pH of 5.8.

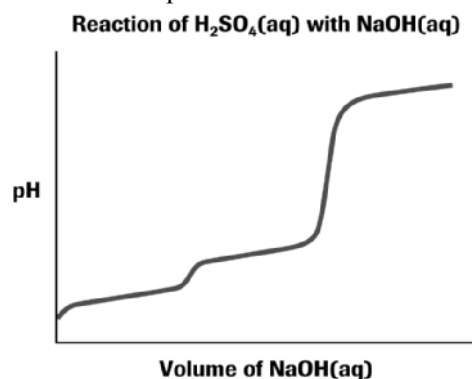
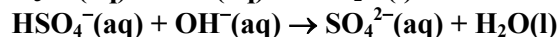
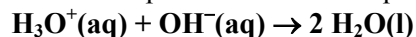
If the remaining two solutions are tested with thymol blue, the one that turns yellow has a pH of 7.8 and the one that turns blue has a pH of 9.8.

[Other acceptable diagnostic tests are possible.]

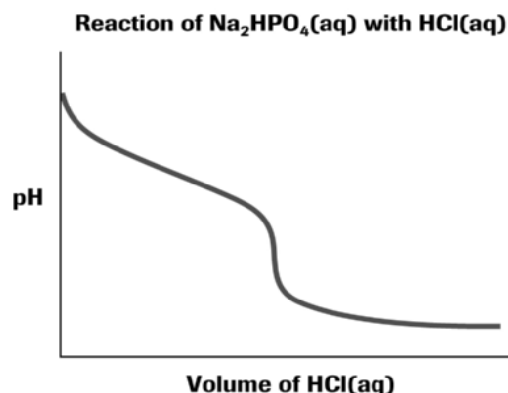
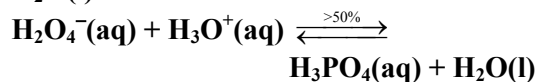
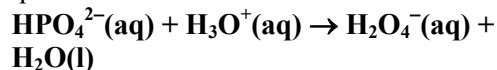
Solutions for pH Curves, Extra Exercises

- Does the burette contain the acid or the base?
The burette contains the base.
 - Is the sample reacted an acid or a base?
The sample is the acid.
 - How many equivalence points are represented? Estimate the pH of each equivalence point.
Two endpoints are apparent at about pH 4.0 and pH 8.5.
 - How many quantitative reactions have occurred?
Two quantitative reactions have occurred.
 - Choose the best indicator for each equivalence point?
Methyl orange or congo red are suitable for the first endpoint, and metacresol purple or thymol blue are suitable for the second endpoint.

- Sketch a pH curve for the reaction of sulfuric acid with sodium hydroxide solution. All reaction steps are known to be quantitative. Include reaction equations.



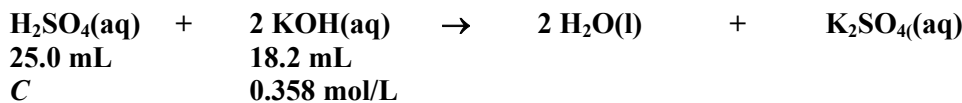
- A sodium hydrogen phosphate solution is to be titrated with hydrochloric acid. Only one quantitative reaction is observed. Sketch the pH curve and write equilibrium equations.



- What are two advantages of detecting an endpoint using an indicator as opposed to identifying it with a pH meter?
Two advantages are that the method of analysis does not require any expensive equipment, and the endpoint is more visible and can be seen during the titration process. (It does not require that you watch a meter while performing the titration.)

Solutions for Acid–Base Stoichiometry, Extra Exercises

1. In a chemical analysis, 25.0 mL of sulfuric acid solution was titrated to the second endpoint with 0.358 mol/L KOH(aq). In the titration, an average volume of 18.2 mL was required. Calculate the molar concentration of the sulfuric acid.



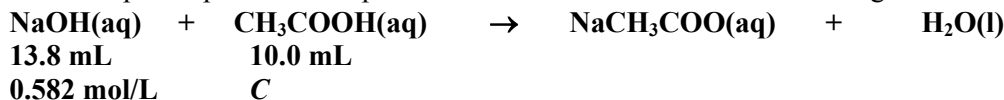
$$n_{\text{KOH}} = 18.2 \text{ mL} \times \frac{0.358 \text{ mol}}{1 \text{ L}} = 6.52 \text{ mmol}$$

$$n_{\text{H}_2\text{SO}_4} = 6.52 \text{ mmol} \times \frac{1}{2} = 3.26 \text{ mmol}$$

$$C_{\text{H}_2\text{SO}_4} = \frac{3.26 \text{ mmol}}{25.0 \text{ mL}} = 0.130 \text{ mol/L}$$

$$\text{or } C_{\text{H}_2\text{SO}_4} = 18.2 \text{ mL KOH} \times \frac{0.358 \text{ mol KOH}}{1 \text{ L KOH}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} \times \frac{1}{250 \text{ mL}} = 0.130 \text{ mol/L}$$

2. Several 10.0 mL vinegar samples were titrated with a standardized 0.582 mol/L solution of sodium hydroxide. An average volume of 13.8 mL of sodium hydroxide was required to reach the phenolphthalein endpoint. What is the concentration of the vinegar solution?



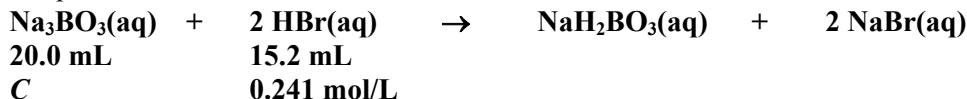
$$n_{\text{NaOH}} = 13.8 \text{ mL} \times \frac{0.582 \text{ mol}}{1 \text{ L}} = 8.03 \text{ mmol}$$

$$n_{\text{CH}_3\text{COOH}} = 8.03 \text{ mmol} \times \frac{1}{1} = 8.03 \text{ mmol}$$

$$C_{\text{CH}_3\text{COOH}} = \frac{8.03 \text{ mmol}}{10.0 \text{ mL}} = 0.803 \text{ mol/L}$$

$$\text{or } C_{\text{CH}_3\text{COOH}} = 13.8 \text{ mL NaOH} \times \frac{0.582 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol NaOH}} \times \frac{1}{10.0 \text{ mL}} = 0.803 \text{ mol/L}$$

3. A sodium borate solution was titrated to the second endpoint with 0.241 mol/L hydrobromic acid. An average volume of 15.2 mL of hydrobromic acid was required to react with 20.0 mL samples of sodium borate. Calculate the molar concentration of sodium borate.



$$n_{\text{HBr}} = 15.2 \text{ mL} \times \frac{0.241 \text{ mol}}{1 \text{ L}} = 3.66 \text{ mmol}$$

$$n_{\text{Na}_3\text{BO}_3} = 3.66 \text{ mmol} \times \frac{1}{2} = 1.83 \text{ mmol}$$

LSM 16.4I (cont'd)

$$C_{\text{Na}_3\text{BO}_3} = \frac{1.83 \text{ mmol}}{20.0 \text{ mL}} = 91.6 \text{ mmol/L}$$

$$\text{or } C_{\text{Na}_3\text{BO}_3} = 15.2 \text{ mL HBr} \times \frac{0.241 \text{ mol HBr}}{1 \text{ L HBr}} \times \frac{1 \text{ mol Na}_3\text{BO}_3}{2 \text{ mol HBr}} \times \frac{1}{20.0 \text{ mL}} = 91.6 \text{ mmol/L}$$

4. Problem

What is the molar concentration of a hydrochloric acid solution?

Design

100.0 mL of a standard solution of sodium oxalate was prepared using 1.85 g of the dry solid. Using the second endpoint, 10.0 mL samples were titrated with hydrochloric acid.

Evidence

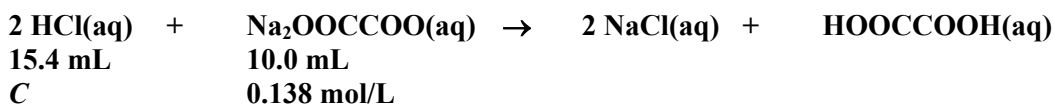
Table 1 Titration of 10.0 mL Samples of $\text{Na}_2\text{O}(\text{OCCOO})_2(\text{aq})$ with $\text{HCl}(\text{aq})$

Trial	1	2	3	4
Final burette reading (mL)	16.1	31.5	46.9	16.9
Initial burette reading (mL)	0.3	16.1	31.5	1.5
Comment on endpoint	poor	good	good	good

Analysis

$$n_{\text{Na}_2\text{O}(\text{OCCOO})_2} = 1.85 \text{ g} \times \frac{1 \text{ mol}}{134.00 \text{ g/mol}} = 0.0138 \text{ mol}$$

$$C_{\text{Na}_2\text{O}(\text{OCCOO})_2} = \frac{0.0138 \text{ mol}}{0.1000 \text{ L}} = 0.138 \text{ mol/L} \quad V_{\text{HCl}} = \frac{(15.4 + 15.4 + 15.4) \text{ mL}}{3} = 15.4 \text{ mL}$$



$$n_{\text{Na}_2\text{O}(\text{OCCOO})_2} = 10.0 \text{ mL} \times \frac{0.138 \text{ mol}}{1 \text{ L}} = 1.38 \text{ mmol}$$

$$n_{\text{HCl}} = 1.38 \text{ mmol} \times \frac{2}{1} = 2.76 \text{ mmol}$$

$$C_{\text{HCl}} = \frac{2.76 \text{ mmol}}{15.4 \text{ mL}} = 0.179 \text{ mol/L}$$

or

$$\begin{aligned} C_{\text{HCl}} &= 10.0 \text{ mL Na}_2\text{O}(\text{OCCOO})_2 \times \frac{0.138 \text{ mol Na}_2\text{O}(\text{OCCOO})_2}{1 \text{ L Na}_2\text{O}(\text{OCCOO})_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Na}_2\text{O}(\text{OCCOO})_2} \times \frac{1}{15.4 \text{ mL}} \\ &= 0.179 \text{ mol/L} \end{aligned}$$

Student Worksheet Solutions LSM 16.CR

Chapter 16 Review, Part 1

1. A
2. D
3. B
4. B
5. C
6. 5, 4, 1, 6
7. 2, 3, 6, 5
8. B
9. C
10. D
11. B
12. A

Student Worksheet Solutions LSM 16.UR
Unit 8 Review, Part 1

1. B
2. A
3. C
4. A
5. A
6. B
7. A
8. A
9. 3, 6, 1, 1
10. C
11. A
12. C
13. A
14. D