A general chemistry Libretexts Textmap organized around the textbook

**Chemistry: The Central Science**
by Brown, LeMay, Busten, Murphy, and Woodward

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These are homework exercises to accompany the Textmap created for "Chemistry: The Central Science" by Brown et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.
2.1: The Atomic Theory of Matter

Conceptual Problems

1. Which of the following elements exist as diatomic molecules?
   a. helium
   b. hydrogen
   c. iodine
   d. gold

2. Which of the following elements exist as diatomic molecules?
   a. chlorine
   b. potassium
   c. silver
   d. oxygen

3. Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H?

4. Why is it proper to represent the elemental form of chlorine as Cl₂ but improper to represent the elemental form of calcium as Ca₂?

Conceptual Solutions

1. a. no
   b. yes
   c. yes
   d. no

2. Hydrogen exists as a diatomic molecule in its elemental form; helium does not exist as a diatomic molecule.

Exercises

1. Which of the following elements exist as diatomic molecules?
   a. helium
   b. hydrogen
   c. iodine
   d. gold

2. Which of the following elements exist as diatomic molecules?
3. Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H? 

4. Why is it proper to represent the elemental form of chlorine as Cl₂ but improper to represent the elemental form of calcium as Ca₂?

Answers
1. a. no
   b. yes
   c. yes
   d. no

2.2: The Discovery of Atomic Structure

Exercises
1. What is the modern atomic theory?
2. What are atoms?

Conceptual Answers
1. The modern atomic theory states that all matter is composed of atoms.
2. Atoms are the smallest parts of an element that maintain the identity of that element.

Numerical Problems
1. (Basic concept check) When 32.0 grams (g) of methane are burned in 128.0 g of oxygen, 88.0 g of carbon dioxide and 72.0 g of water are produced. Which law is this an example of? (a) Law of definite proportions (b) Law of conservation of mass or (c) Law of multiple proportions.

2. (Law of Conservation of Mass) 8.00 grams (g) of methane are burned in 32.00 g of oxygen. The reaction produces 22.00 g of carbon dioxide and an unmeasured mass of water. What mass of water is produced?

3. (Law of Definite Proportions) Two experiments using sodium and chlorine are performed. In the first experiment, 4.36 grams (g) sodium are reacted with 32.24 g of chlorine, using up all the sodium. 11.08 g of sodium chloride was produced in the first experiment. In the second experiment, 4.20 g of chlorine reacted with 20.00 g of sodium, using up all the chlorine. 6.92 g of sodium chloride was produced in the second experiment. Show that these
results are consistent with the law of constant composition.

4. (Law of Conservation of Mass): 36.0 grams (g) of wood are burned in oxygen. The products of this reaction weigh 74.4 g. (a) What mass of oxygen is needed in this reaction? (b) What mass of oxygen is needed to burn 8.00 lb of wood? 1 lb = 453.59237 g.

5. (Law of Definite Proportions): A sample of methane contains only carbon and hydrogen, with 3.00 grams (g) of carbon for every 1.00 g of hydrogen. How much hydrogen should be present in a different, 50.0 g same of methane?

Numerical Solutions

1. The answer is (b) Law of conservation of mass. The number of grams of reactants (32.0 g of methane and 128.0 g of oxygen = 160.0 g total) is equal to the number of grams of product (88.0 g of carbon dioxide and 72.0 g of water = 160.0 g total).

2. The answer is 18.00 g of water. Because the only products are water and carbon dioxide, their total mass must equal the total mass of the reactants, methane and oxygen. 8.00 g of methane + 32.00 g of oxygen = 40.00 total g of reactants. Because the total mass of the reactants equals the total mass of the products, the total mass of the products is also 40.00 g. Thus, 40.00 total g of products = 22.00 g carbon dioxide + unknown mass water. 40.00 total g of products - 22.00 g carbon dioxide = 18.00 g water.

3. To solve, determine the percent of sodium in each sample of sodium chloride. There is 4.36 g sodium for every 11.08 g of sodium chloride in the first experiment. The amount of sodium in the sodium chloride for the second experiment must be found. This is found by subtracted the known amount of reacted chlorine (4.20 g) from the amount of sodium chloride (6.92 g). 6.92 g sodium chloride - 4.20 g chlorine = 2.72 g sodium.

Thus, the percent of sodium in each sample is represented below:

% Na = (4.36 g Na)/(11.08 g NaCl) x 100% = 39.4% Na % Na = (2.72 g Na)/(6.92 g NaCl) x 100% = 39.3%

The slight difference in compositions is due to significant figures: each percent has an uncertainty of .01% in either direction. The two samples of sodium chloride have the same composition.

4. a. The answer is 38.4 g of oxygen. The total mass of the products is 74.4 g. Thus, the total mass of the reactants must equal 74.4 g as well. Thus, 74.4 g products - 36.0 g wood reactant = 38.4 g oxygen reactant.

b. The answer is 8.53 lb of oxygen. From, (a) that it takes 38.4 g of oxygen to burn 18.0 g of wood. First, convert both of these values to pounds (alternatively, the 8.00 lb can be converted to grams).

36.0 g wood x (1 lb)/(453.59237 g) = 0.0793664144 lb wood
38.4 g oxygen x (1 lb)/(453.59237 g) = .0846575087 lb oxygen

Now two ratios equal to each other can be set up to determine the unknown mass of oxygen.

(0.0793664144 lb wood)/(.0846575087 lb oxygen) = (8.00 lb wood)/(unknown mass oxygen)

Solving reveals that it requires 8.53 lb of oxygen to burn 8.00 lb of wood.

5. The answer is 12.5 g of hydrogen. If there are 3.00 g of carbon present for every 1.00 g of hydrogen, we can assume the smallest whole number combination of these elements in that ratio to be 4.00 g of methane: 50.0 g methane x (1.00 g hydrogen)/(4.00 g methane) = 12.5 g of hydrogen.
2.3: The Modern View of Atomic Structure

Conceptual Problems

1. Describe the experiment that provided evidence that the proton is positively charged.
2. What observation led Rutherford to propose the existence of the neutron?
3. What is the difference between Rutherford’s model of the atom and the model chemists use today?
4. If cathode rays are not deflected when they pass through a region of space, what does this imply about the presence or absence of a magnetic field perpendicular to the path of the rays in that region?
5. Describe the outcome that would be expected from Rutherford’s experiment if the charge on α particles had remained the same but the nucleus were negatively charged. If the nucleus were neutral, what would have been the outcome?
6. Describe the differences between an α particle, a β particle, and a γ ray. Which has the greatest ability to penetrate matter?

Numerical Problems

Please be sure you are familiar with the topics discussed in Section 1.6 before proceeding to the Numerical Problems.

1. Using the data in Table 1.3 and the periodic table, calculate the percentage of the mass of a silicon atom that is due to
   a. electrons.
   b. protons.
2. Using the data in Table 1.3 and the periodic table, calculate the percentage of the mass of a helium atom that is due to
   a. electrons.
   b. protons.
3. The radius of an atom is approximately 10^4 times larger than the radius of its nucleus. If the radius of the nucleus were 1.0 cm, what would be the radius of the atom in centimeters? in miles?
4. The total charge on an oil drop was found to be 3.84 \times 10^{-18} coulombs. What is the total number of electrons contained in the drop?

2.4: Atomic Mass

Conceptual Problems

1. Complete the following table for the missing elements, symbols, and numbers of electrons.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Number of Electrons</th>
</tr>
</thead>
</table>
2. Complete the following table for the missing elements, symbols, and numbers of electrons.

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>molybdenum</td>
<td></td>
<td>19</td>
</tr>
<tr>
<td>titanium</td>
<td>B</td>
<td>53</td>
</tr>
<tr>
<td>Sm</td>
<td></td>
<td>89</td>
</tr>
<tr>
<td>helium</td>
<td></td>
<td>14</td>
</tr>
</tbody>
</table>

3. Is the mass of an ion the same as the mass of its parent atom? Explain your answer.

4. What isotopic standard is used for determining the mass of an atom?

5. Give the symbol \(^{\text{A}}_{\text{Z}} \text{X}\) for these elements, all of which exist as a single isotope.
   a. beryllium
   b. ruthenium
   c. phosphorus
   d. aluminum
   e. cesium
6. Give the symbol \(_Z^AX\) for these elements, all of which exist as a single isotope.

a. fluorine
b. helium
c. terbium
d. iodine
e. gold
f. scandium
g. sodium
h. niobium
i. manganese

7. Identify each element, represented by X, that have the given symbols.

a. \(_{26}^{55}X\)
b. \(_{33}^{74}X\)
c. \(_{12}^{24}X\)
d. \(_{53}^{127}X\)
e. \(_{18}^{40}X\)
f. \(_{63}^{152}X\)

Numerical Problems

Please be sure you are familiar with the topics discussed in Section 1.6 before proceeding to the Numerical Problems.

1. The isotopes \(^{131}\text{I}\) and \(^{60}\text{Co}\) are commonly used in medicine. Determine the number of neutrons, protons, and electrons in a neutral atom of each.

2. Determine the number of protons, neutrons, and electrons in a neutral atom of each isotope:

a. \(^{97}\text{Tc}\)
b. \(^{113}\text{In}\)
c. \(^{63}\text{Ni}\)
d. \(^{55}\text{Fe}\)

3. Both technetium-97 and americium-240 are produced in nuclear reactors. Determine the number of protons, neutrons, and electrons in the neutral atoms of each.
4. The following isotopes are important in archaeological research. How many protons, neutrons, and electrons does a neutral atom of each contain?
   a. \(^{207}\text{Pb}\)
   b. \(^{16}\text{O}\)
   c. \(^{40}\text{K}\)
   d. \(^{137}\text{Cs}\)
   e. \(^{40}\text{Ar}\)

5. Copper, an excellent conductor of heat, has two isotopes: \(^{63}\text{Cu}\) and \(^{65}\text{Cu}\). Use the following information to calculate the average atomic mass of copper:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance (%)</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{63}\text{Cu})</td>
<td>69.09</td>
<td>62.9298</td>
</tr>
<tr>
<td>(^{65}\text{Cu})</td>
<td>30.92</td>
<td>64.9278</td>
</tr>
</tbody>
</table>

6. Silicon consists of three isotopes with the following percent abundances:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance (%)</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{28}\text{Si})</td>
<td>92.18</td>
<td>27.976926</td>
</tr>
<tr>
<td>(^{29}\text{Si})</td>
<td>4.71</td>
<td>28.976495</td>
</tr>
<tr>
<td>(^{30}\text{Si})</td>
<td>3.12</td>
<td>29.973770</td>
</tr>
</tbody>
</table>

Calculate the average atomic mass of silicon.

7. Complete the following table for neon. The average atomic mass of neon is 20.1797 amu.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Abundance (%)</th>
<th>Atomic Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{20}\text{Ne})</td>
<td>90.92</td>
<td>19.99244</td>
</tr>
<tr>
<td>(^{21}\text{Ne})</td>
<td>0.257</td>
<td>20.99395</td>
</tr>
<tr>
<td>(^{22}\text{Ne})</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

8. Are \(^{28}\text{X}\) and \(^{29}\text{X}\) isotopes of the same element? Explain your answer.

9. Complete the following table:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
10. Complete the following table:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{238}\text{X}$</td>
<td></td>
<td></td>
<td>95</td>
</tr>
<tr>
<td>$^{238}\text{U}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>75</td>
<td>112</td>
<td></td>
</tr>
</tbody>
</table>

11. Using a mass spectrometer, a scientist determined the percent abundances of the isotopes of sulfur to be 95.27% for $^{32}\text{S}$, 0.51% for $^{33}\text{S}$, and 4.22% for $^{34}\text{S}$. Use the atomic mass of sulfur from the periodic table (see Chapter 32 "Appendix H: Periodic Table of Elements") and the following atomic masses to determine whether these data are accurate, assuming that these are the only isotopes of sulfur: 31.972071 amu for $^{32}\text{S}$, 32.971459 amu for $^{33}\text{S}$, and 33.967867 amu for $^{34}\text{S}$.

12. The percent abundances of two of the three isotopes of oxygen are 99.76% for $^{16}\text{O}$, and 0.204% for $^{18}\text{O}$. Use the atomic mass of oxygen given in the periodic table and the following data to determine the mass of $^{17}\text{O}$: 15.994915 amu for $^{16}\text{O}$ and 17.999160 amu for $^{18}\text{O}$.

13. Which element has the higher proportion by mass in NaI?

14. Which element has the higher proportion by mass in KBr?

**2.5: The Periodic Table**

**Conceptual Problems**

1. Classify each element in Conceptual Problem 1 as a metal, a nonmetal, or a semimetal. If a metal, state whether it is an alkali metal, an alkaline earth metal, or a transition metal.

2. Classify each element in Conceptual Problem 2) as a metal, a nonmetal, or a semimetal. If a metal, state whether it is an alkali metal, an alkaline earth metal, or a transition metal.

3. Classify each element as a metal, a semimetal, or a nonmetal. If a metal, state whether it is an alkali metal, an alkaline earth metal, or a transition metal.
a. iron
b. tantalum
c. sulfur
d. silicon
e. chlorine
f. nickel
g. potassium
h. radon
i. zirconium

4. Which of these sets of elements are all in the same period?
   a. potassium, vanadium, and ruthenium
   b. lithium, carbon, and chlorine
   c. sodium, magnesium, and sulfur
   d. chromium, nickel, and krypton

5. Which of these sets of elements are all in the same period?
   a. barium, tungsten, and argon
   b. yttrium, zirconium, and selenium
   c. potassium, calcium, and zinc
   d. scandium, bromine, and manganese

6. Which of these sets of elements are all in the same group?
   a. sodium, rubidium, and barium
   b. nitrogen, phosphorus, and bismuth
   c. copper, silver, and gold
   d. magnesium, strontium, and samarium

7. Which of these sets of elements are all in the same group?
   a. iron, ruthenium, and osmium
   b. nickel, palladium, and lead
   c. iodine, fluorine, and oxygen
   d. boron, aluminum, and gallium

8. Indicate whether each element is a transition metal, a halogen, or a noble gas.
   a. manganese
   b. iridium
   c. fluorine
d. xenon  
e. lithium  
f. carbon  
g. zinc  
h. sodium  
i. tantalum  
j. hafnium  
k. antimony  
l. cadmium

9. Which of the elements indicated in color in the periodic table shown below is most likely to exist as a monoatomic gas? As a diatomic gas? Which is most likely to be a semimetal? A reactive metal?

10. Based on their locations in the periodic table, would you expect these elements to be malleable? Why or why not?

   a. phosphorus  
   b. chromium  
   c. rubidium  
   d. copper  
   e. aluminum  
   f. bismuth  
   g. neodymium

11. Based on their locations in the periodic table, would you expect these elements to be lustrous? Why or why not?

   a. sulfur  
   b. vanadium  
   c. nickel  
   d. arsenic  
   e. strontium  
   f. cerium  
   g. sodium
### Conceptual Solution

3.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe</td>
<td>metal: transition metal</td>
</tr>
<tr>
<td>Ta</td>
<td>metal: transition metal</td>
</tr>
<tr>
<td>S</td>
<td>nonmetal</td>
</tr>
<tr>
<td>Si</td>
<td>semimetal</td>
</tr>
<tr>
<td>Cl</td>
<td>nonmetal (halogen)</td>
</tr>
<tr>
<td>Ni</td>
<td>metal: transition metal</td>
</tr>
<tr>
<td>K</td>
<td>metal: alkali metal</td>
</tr>
<tr>
<td>Rn</td>
<td>nonmetal (noble gas)</td>
</tr>
<tr>
<td>Zr</td>
<td>metal: transition meta</td>
</tr>
</tbody>
</table>

### 2.6: Molecules and Molecular Compounds

#### Conceptual Problems

1. Ionic and covalent compounds are held together by electrostatic attractions between oppositely charged particles. Describe the differences in the nature of the attractions in ionic and covalent compounds. Which class of compounds contains pairs of electrons shared between bonded atoms?

2. Which contains fewer electrons than the neutral atom—the corresponding cation or the anion?

3. What is the difference between an organic compound and an inorganic compound?

4. What is the advantage of writing a structural formula as a condensed formula?

5. The majority of elements that exist as diatomic molecules are found in one group of the periodic table. Identify the group.

6. Discuss the differences between covalent and ionic compounds with regard to
   a. the forces that hold the atoms together.
   b. melting points.
   c. physical states at room temperature and pressure.

7. Why do covalent compounds generally tend to have lower melting points than ionic compounds?

#### Conceptual Answer

7. Covalent compounds generally melt at lower temperatures than ionic compounds because the intermolecular
interactions that hold the molecules together in a molecular solid are weaker than the electrostatic attractions that hold oppositely charged ions together in an ionic solid.

**Numerical Problems**

1. The structural formula for chloroform (CHCl₃) was shown in Example 2.6.2. Based on this information, draw the structural formula of dichloromethane (CH₂Cl₂).

2. What is the total number of electrons present in each ion?
   a. F⁻
   b. Rb⁺
   c. Ce³⁺
   d. Zr⁴⁺
   e. Zn²⁺
   f. Kr²⁺
   g. B³⁺

3. What is the total number of electrons present in each ion?
   a. Ca²⁺
   b. Se²⁻
   c. In³⁺
   d. Sr²⁺
   e. As³⁺
   f. N³⁻
   g. Ti⁺

4. Predict how many electrons are in each ion.
   a. an oxygen ion with a −2 charge
   b. a beryllium ion with a +2 charge
   c. a silver ion with a +1 charge
   d. a selenium ion with a +4 charge
   e. an iron ion with a +2 charge
   f. a chlorine ion with a −1 charge

5. Predict how many electrons are in each ion.
   a. copper ion with a +2 charge
   b. a molybdenum ion with a +4 charge
c. an iodine ion with a −1 charge
d. a gallium ion with a +3 charge
e. an ytterbium ion with a +3 charge
f. a scandium ion with a +3 charge

6. Predict the charge on the most common monatomic ion formed by each element.
   a. chlorine
   b. phosphorus
c. scandium
d. magnesium
e. arsenic
f. oxygen

7. Predict the charge on the most common monatomic ion formed by each element.
   a. sodium
   b. selenium
c. barium
d. rubidium
e. nitrogen
f. aluminum

8. For each representation of a monatomic ion, identify the parent atom, write the formula of the ion using an appropriate superscript, and indicate the period and group of the periodic table in which the element is found.
   a. \(_{4}^{9}X^{2+} \)
b. \(_{1}^{1}X^- \)
c. \(_{8}^{16}X^{2-} \)

9. For each representation of a monatomic ion, identify the parent atom, write the formula of the ion using an appropriate superscript, and indicate the period and group of the periodic table in which the element is found.
   a. \(_{3}^{7}X^+ \)
b. \(_{9}^{19}X^- \)
c. \(_{13}^{27}X^{3+} \)

**Numerical Answers**

5.
   a. 27
   b. 38
   c. 54
9.

a. Li, Li\(^+\), 2nd period, group 1
b. F, F\(^-\), 2nd period, group 17
c. Al, Al\(^3+\), 3rd period, group 13

2.7: Ions and Ionic Compounds

2.8: Naming Inorganic Compounds

Conceptual Problems

1. What are the differences and similarities between a polyatomic ion and a molecule?

2. Classify each compound as ionic or covalent.
   a. Zn\(_3\)(PO\(_4\))\(_2\)
   b. C\(_6\)H\(_5\)CO\(_2\)H
   c. K\(_2\)Cr\(_2\)O\(_7\)
   d. CH\(_3\)CH\(_2\)SH
   e. NH\(_4\)Br
   f. CCl\(_2\)F\(_2\)

3. Classify each compound as ionic or covalent. Which are organic compounds and which are inorganic compounds?
   a. CH\(_3\)CH\(_2\)CO\(_2\)H
   b. CaCl\(_2\)
   c. Y(NO\(_3\))\(_3\)
   d. H\(_2\)S
   e. NaC\(_2\)H\(_3\)O\(_2\)

4. Generally, one cannot determine the molecular formula directly from an empirical formula. What other information is needed?

5. Give two pieces of information that we obtain from a structural formula that we cannot obtain from an empirical formula.
6. The formulas of alcohols are often written as ROH rather than as empirical formulas. For example, methanol is generally written as CH$_3$OH rather than CH$_4$O. Explain why the ROH notation is preferred.

7. The compound dimethyl sulfide has the empirical formula C$_2$H$_6$S and the structural formula CH$_3$SCH$_3$. What information do we obtain from the structural formula that we do not get from the empirical formula? Write the condensed structural formula for the compound.

8. What is the correct formula for magnesium hydroxide—MgOH$_2$ or Mg(OH)$_2$? Why?

9. Magnesium cyanide is written as Mg(CN)$_2$, not MgCN$_2$. Why?

10. Does a given hydrate always contain the same number of waters of hydration?

**Conceptual Solutions**

7. The structural formula gives us the connectivity of the atoms in the molecule or ion, as well as a schematic representation of their arrangement in space. Empirical formulas tell us only the ratios of the atoms present. The condensed structural formula of dimethylsulfide is (CH$_3$)$_2$S.

**Numerical Problems**

1. Write the formula for each compound.
   a. magnesium sulfate, which has 1 magnesium atom, 4 oxygen atoms, and 1 sulfur atom
   b. ethylene glycol (antifreeze), which has 6 hydrogen atoms, 2 carbon atoms, and 2 oxygen atoms
   c. acetic acid, which has 2 oxygen atoms, 2 carbon atoms, and 4 hydrogen atoms
   d. potassium chlorate, which has 1 chlorine atom, 1 potassium atom, and 3 oxygen atoms
   e. sodium hypochlorite pentahydrate, which has 1 chlorine atom, 1 sodium atom, 6 oxygen atoms, and 10 hydrogen atoms

2. Write the formula for each compound.
   a. cadmium acetate, which has 1 cadmium atom, 4 oxygen atoms, 4 carbon atoms, and 6 hydrogen atoms
   b. barium cyanide, which has 1 barium atom, 2 carbon atoms, and 2 nitrogen atoms
   c. iron(III) phosphate dihydrate, which has 1 iron atom, 1 phosphorus atom, 6 oxygen atoms, and 4 hydrogen atoms
   d. manganese(II) nitrate hexahydrate, which has 1 manganese atom, 12 hydrogen atoms, 12 oxygen atoms, and 2 nitrogen atoms
   e. silver phosphate, which has 1 phosphorus atom, 3 silver atoms, and 4 oxygen atoms

3. Complete the following table by filling in the formula for the ionic compound formed by each cation-anion pair.

<table>
<thead>
<tr>
<th>Ion</th>
<th>K$^+$</th>
<th>Fe$^{3+}$</th>
<th>NH$_4^+$</th>
<th>Ba$^{2+}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl$^-$</td>
<td>KCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>SO$_4^{2-}$</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
4. Write the empirical formula for the binary compound formed by the most common monatomic ions formed by each pair of elements.

a. zinc and sulfur  
b. barium and iodine  
c. magnesium and chlorine  
d. silicon and oxygen  
e. sodium and sulfur

5. Write the empirical formula for the binary compound formed by the most common monatomic ions formed by each pair of elements.

a. lithium and nitrogen  
b. cesium and chlorine  
c. germanium and oxygen  
d. rubidium and sulfur  
e. arsenic and sodium

6. Write the empirical formula for each compound.

a. Na₂S₂O₄  
b. B₂H₆  
c. C₆H₁₂O₆  
d. P₄O₁₀  
e. KMnO₄

7. Write the empirical formula for each compound.

a. Al₂Cl₆  
b. K₂Cr₂O₇  
c. C₂H₄  
d. (NH₂)₂CNH  
e. CH₃COOH
Numerical Answers

1.
   a. MgSO$_4$
   b. C$_2$H$_6$O$_2$
   c. C$_2$H$_4$O$_2$
   d. KClO$_3$
   e. NaOCl·5H$_2$O

3.

<table>
<thead>
<tr>
<th>Ion</th>
<th>K$^+$</th>
<th>Fe$^{3+}$</th>
<th>NH$_4^+$</th>
<th>Ba$^{2+}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl$^-$</td>
<td>KCl</td>
<td>FeCl$_3$</td>
<td>NH$_4$Cl</td>
<td>BaCl$_2$</td>
</tr>
<tr>
<td>SO$_4^{2-}$</td>
<td>K$_2$SO$_4$</td>
<td>Fe$_2$(SO$_4$)$_3$</td>
<td>(NH$_4$)$_2$SO$_4$</td>
<td>BaSO$_4$</td>
</tr>
<tr>
<td>PO$_4^{3-}$</td>
<td>K$_3$PO$_4$</td>
<td>FePO$_4$</td>
<td>(NH$_4$)$_3$PO$_4$</td>
<td>Ba$_3$(PO$_4$)$_2$</td>
</tr>
<tr>
<td>NO$_3^-$</td>
<td>KNO$_3$</td>
<td>Fe(NO$_3$)$_3$</td>
<td>NH$_4$NO$_3$</td>
<td>Ba(NO$_3$)$_2$</td>
</tr>
<tr>
<td>OH$^-$</td>
<td>KOH</td>
<td>Fe(OH)$_3$</td>
<td>NH$_4$OH</td>
<td>Ba(OH)$_2$</td>
</tr>
</tbody>
</table>

5.
   a. Li$_3$N
   b. CsCl
   c. GeO$_2$
   d. Rb$_2$S
   e. Na$_3$As

7.
   a. AlCl$_3$
   b. K$_2$Cr$_2$O$_7$
   c. CH$_2$
   d. CH$_3$N$_3$
   e. CH$_2$O

2.9: Some Simple Organic Compounds
Conceptual Problems

1. Benzene (C₆H₆) is an organic compound, and KCl is an ionic compound. The sum of the masses of the atoms in each empirical formula is approximately the same. How would you expect the two to compare with regard to each of the following? What species are present in benzene vapor?

   a. melting point
   b. type of bonding
   c. rate of evaporation
   d. structure

2. Can an inorganic compound be classified as a hydrocarbon? Why or why not?

3. Is the compound NaHCO₃ a hydrocarbon? Why or why not?

4. Name each compound.
   a. NiO
   b. TiO₂
   c. N₂O
   d. CS₂
   e. SO₃
   f. NF₃
   g. SF₆

5. Name each compound.
   a. HgCl₂
   b. IF₅
   c. N₂O₅
   d. Cl₂O
   e. HgS
   f. PCl₅

6. For each structural formula, write the condensed formula and the name of the compound.

   a. 
   ![Image]

   b. 
   ![Image]
7. For each structural formula, write the condensed formula and the name of the compound.

a.

b.

c.
8. Would you expect PCl₃ to be an ionic compound or a covalent compound? Explain your reasoning.

9. What distinguishes an aromatic hydrocarbon from an aliphatic hydrocarbon?

10. The following general formulas represent specific classes of hydrocarbons. Refer to Table 2.7 "The First 10 Straight-Chain Alkanes" and Table 2.8 "Some Common Acids That Do Not Contain Oxygen" and Figure 2.16 and identify the classes.
   
   a. \( C_nH_{2n} + 2 \)
   b. \( C_nH_{2n} \)
   c. \( C_nH_{2n} - 2 \)

11. Using R to represent an alkyl or aryl group, show the general structure of an
   
   a. alcohol.
   b. phenol.

**Conceptual Answer**

11.

   a. ROH (where R is an alkyl group)
   b. ROH (where R is an aryl group)

**Numerical Problems**

1. Write the formula for each compound.
   
   a. dinitrogen monoxide
   b. silicon tetrafluoride
   c. boron trichloride
   d. nitrogen trifluoride
   e. phosphorus tribromide
2. Write the formula for each compound.
   a. dinitrogen trioxide
   b. iodine pentafluoride
   c. boron tribromide
   d. oxygen difluoride
   e. arsenic trichloride

3. Write the formula for each compound.
   a. thallium(I) selenide
   b. neptunium(IV) oxide
   c. iron(II) sulfide
   d. copper(I) cyanide
   e. nitrogen trichloride

4. Name each compound.
   a. RuO$_4$
   b. PbO$_2$
   c. MoF$_6$
   d. Hg$_2$(NO$_3$)$_2$$\cdot$2H$_2$O
   e. WCl$_4$

5. Name each compound.
   a. NbO$_2$
   b. MoS$_2$
   c. P$_4$S$_{10}$
   d. Cu$_2$O
   e. ReF$_5$

6. Draw the structure of each compound.
   a. propyne
   b. ethanol
   c. n-hexane
   d. cyclopropane
   e. benzene

7. Draw the structure of each compound.
   a. 1-butene
   b. 2-pentyne
c. cycloheptane
d. toluene
e. phenol

Numerical Answers

1.
   a. N$_2$O
   b. SiF$_4$
   c. BCl$_3$
   d. NF$_3$
   e. PBr$_3$

2.
   a. Tl$_2$Se
   b. NpO$_2$
   c. FeS
   d. CuCN
   e. NCl$_3$

5.
   a. niobium (IV) oxide
   b. molybdenum (IV) sulfide
   c. tetraphosphorus decasulfide
   d. copper(I) oxide
   e. rhenium(V) fluoride

7.
   a.

   b.

   c.

   d.
e.

\[ \text{CH}_3 \]

\[ \text{OH} \]