

CHAPTER 2

Molecules, Ions, and Chemical Formulas

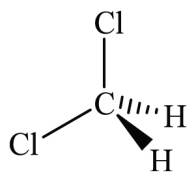
1. CHEMICAL COMPOUNDS

Conceptual Problems

1. The designation of ionic and covalent bonding is an arbitrary one. Bonding is most accurately described as a spectrum, ranging from purely covalent compounds such as H_2 to “ionic” compounds such as CsF . For now the following definition suffices: Covalent bonding involves the *sharing* of electron *pairs*, ignoring for now odd-electron molecules such as NO , whereas complete *transfer* of *one or more electrons* is assumed in ionic bonding. Thus, covalent bonding arises from the positive nuclei sharing electrons, and the electrons are not considered to “belong” to any one nucleus. Ionic bonding, however, involves an attraction between actual ions, where the transferred electron(s) is(are) “assigned” to one nucleus in particular. In the above examples, the two H nuclei share two electrons, whereas in CsF the bonding is attributed to the attraction between a Cs^+ and F^- ion pair.
2. Cations are positively charged ions, and anions are negatively charged ions. Cations become positively charged due to the loss of electrons. Therefore, a cation contains fewer electrons than the neutral atom. (An anion contains more electrons than the neutral atom.)
3. “Organic” compounds contain predominantly carbon and hydrogen and are held together by covalent bonds. Generally, “inorganic” compounds, the bonding for which can be either covalent or ionic, are those compounds that contain elements other than carbon and hydrogen. Note, however, that compounds such as HCl are classified as inorganic compounds, even though they contain hydrogen. Quite often, if you see carbon, the compound is organic, and if you see a metal, the compound is inorganic.
4. Condensed structural formulas can save time and space. For example, note how easy it is to write the condensed structural formula for decane, as opposed to its unwieldy structural formula. Additionally, condensed structural formulas typically focus on the atoms that are the key players in chemical reactions. An example of this can be seen among the family of carboxylic acids. The reactive part of carboxylic acids is the $-\text{COOH}$ group, and the hydrocarbon chain is often “just along for the ride.” Using condensed structural formulas omits the “clutter” of certain compounds, allowing for a more clear illustration of their reactions. The “R” notation, where the entire irrelevant portion of the molecule is denoted as an R as in $\text{R}-\text{COOH}$, is the ultimate in condensing.
5. The halogens exist as diatomic molecules. At room temperature and pressure, F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid.
6.
 - a. Covalent bonding arises from the positive nuclei *sharing* electrons, and the electrons are not considered to “belong” to any one nucleus. Ionic bonding involves attraction between actual ions, where the *transferred* electron(s) is(are) “assigned” to one nucleus in particular.
 - b. Covalent compounds generally have low melting points. Ionic compounds generally have high melting points. For example, Al_2O_3 , CaO , and NaCl have melting points of 2053, 2898, and 800.7 °C, respectively. The low melting points for covalent compounds are due to weak intermolecular forces between the *individual molecules* that comprise the solid. It is *not* correct simply to say that ionic solids have high boiling points due to strong intermolecular forces, however. Such reasoning is false because *no individual molecular units exist in ionic solids*. Ionic solids typically consist of a crystal lattice of ions. The strong ionic interactions that repeat throughout the crystal account for the high melting points.
 - c. The discussion of part (b) is relevant here. Thermal energy, energy that is available at room temperature to create random motion, is sufficient to overcome the weak intermolecular forces, or “bonds”, between the molecules of covalent compounds, whereas it is not able to destroy the crystal lattice of ionic solids. Therefore, many covalent compounds are gases or liquids at room temperature, and ionic solids remain so at room temperature and normal pressures. (The most important covalent compound that is not a gas at room temperature is H_2O .)
7. Covalent compounds generally melt at lower temperatures than ionic compounds because the intermolecular attractions that hold *individual 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 common name for CH_2Cl_2 is methylene chloride. The structure is provided.



2. The charge on ions is due to the loss or gain of *electrons* by the neutral atom. Remember that in any *chemical* reaction the number of protons does *not* change. In each case, then, first determine the number of electrons in the neutral atom. Next, add electrons to form anions or subtract electrons to form cations. For example, a neutral F atom has 9 electrons. To form an F^- ion, one must add one electron to F. Therefore, F^- has 10 electrons. Likewise for the rest: Rb^+ , 36 electrons; Ce^{3+} , 55 electrons; Zr^{4+} , 36 electrons; Zn^{2+} , 28 electrons; Kr^{2+} , 34 electrons; B^{3+} , 2 electrons.
3. The charge on ions is due to the loss or gain of *electrons* by the neutral atom. Remember that in any *chemical* reaction the number of protons does *not* change. In each case, then, first determine the number of electrons in the neutral atom. Next, add electrons to form anions or subtract electrons to form cations. For example, a neutral Ca atom has 20 electrons. To form a Ca^{2+} ion, one must remove two electrons from Ca. Therefore, Ca^{2+} has 18 electrons. Likewise for the rest: Se^{2-} , 36 electrons; In^{3+} , 46 electrons; Sr^{2+} , 36 electrons; As^{3+} , 30 electrons; N^{3-} , 10 electrons; Tl^+ , 80 electrons.
4. a. Because the atomic number of oxygen is 8, a neutral oxygen atom contains 8 electrons. Adding 2 more to form an ion with a -2 charge gives a total of 10 electrons.
- b. Because the atomic number of beryllium is 4, a neutral beryllium atom contains 4 electrons. Removing 2 electrons to form an ion with a $+2$ charge gives a total of 2 electrons.
- c. Because the atomic number of silver is 47, a neutral silver atom contains 47 electrons. Removing 1 electron to form an ion with a $+1$ charge gives a total of 46 electrons.
- d. Because the atomic number of selenium is 34, a neutral silver atom contains 34 electrons. Removing 4 electrons to form an ion with a $+4$ charge gives a total of 30 electrons.
- e. Because the atomic number of iron is 26, a neutral iron atom contains 26 electrons. Removing 2 electrons to form an ion with a $+2$ charge gives a total of 24 electrons.
- f. Because the atomic number of chlorine is 17, a neutral chlorine atom contains 17 electrons. Adding 1 more to form an ion with a -1 charge gives a total of 18 electrons.
5. a. Because the atomic number of copper is 29, a neutral copper atom contains 29 electrons. Removing 2 electrons to form an ion with a $+2$ charge gives a total of 27 electrons.
- b. Because the atomic number of molybdenum is 42, a neutral molybdenum atom contains 42 electrons. Removing 4 electrons to form an ion with a $+4$ charge gives a total of 38 electrons.
- c. Because the atomic number of iodine is 53, a neutral iodine atom contains 53 electrons. Adding 1 more to form an ion with a -1 charge gives a total of 54 electrons.
- d. Because the atomic number of gallium is 31, a neutral gallium atom contains 31 electrons. Removing 3 electrons to form an ion with a $+3$ charge gives a total of 28 electrons.
- e. Because the atomic number of ytterbium is 70, a neutral ytterbium atom contains 70 electrons. Removing 3 electrons to form an ion with a $+3$ charge gives a total of 67 electrons.
- f. Because the atomic number of scandium is 21, a neutral scandium atom contains 21 electrons. Removing 3 electrons to form an ion with a $+3$ charge gives a total of 18 electrons.
6. Nearly all of the main group elements exist only in a single charge state as an ion. Groups 1, 2, 13, 15, 16, and 17 form ions with charges $1+$, $2+$, $3+$, $3-$, $2-$, and $1-$, respectively. This attempt by the main group elements to obtain the same number of electrons as the noble gas nearest to each element is often referred to as the “octet rule.” The reason will be made clear later. Note that group 14 was not mentioned because it rarely forms ions, although C^{4-} is seen. Sc is a transition metal, but it still obeys the octet rule and forms only one ion. Thus, we predict (a) Cl^- ; (b) P^{3-} ; (c) Sc^{3+} ; (d) Mg^{2+} ; (e) As^{3-} ; (f) O^{2-} .
7. Nearly all of the main group elements exist only in a single charge state as an ion. Groups 1, 2, 13, 15, 16, and 17 form ions with charges $1+$, $2+$, $3+$, $3-$, $2-$, and $1-$, respectively. This attempt by the main group elements to obtain the same number of electrons as the noble gas nearest to each element is often referred to as the “octet rule.” The reason will be made clear later. Note that group 14 was not mentioned because it rarely forms ions, although carbide C^{4-} is seen. Thus, we predict (a) Na^+ ; (b) Se^{2-} ; (c) Ba^{2+} ; (d) Rb^+ ; (e) N^{3-} ; (f) Al^{3+} .
8. a. The left subscript of “4” identifies the element as beryllium, and the left superscript of “9” indicates the specific isotope of beryllium, beryllium-9. The ion is therefore Be^{2+} . The periodic table shows that beryllium is in the second period and Group 2.
- b. The left subscript of “1” identifies the element as hydrogen, and the left superscript of “1” indicates the specific isotope of hydrogen, hydrogen-1. The ion is therefore H^+ . The periodic table shows that hydrogen is in the first period and Group 1.
- c. The left subscript of “8” identifies the element as oxygen, and the left superscript of “16” indicates the specific isotope of oxygen, oxygen-16. The ion is therefore O^{2-} . The periodic table shows that oxygen is in the second period and Group 16.

9. a. The left subscript of “3” identifies the element as lithium, and the left superscript of “7” indicates the specific isotope of lithium, lithium-7. The ion is therefore Li^+ . The periodic table shows that lithium is in the second period and Group 1.
- b. The left subscript of “9” identifies the element as fluorine, and the left superscript of “19” indicates the specific isotope of fluorine, fluorine-19. The ion is therefore F^- . The periodic table shows that fluorine is in the second period and Group 17.
- c. The left subscript of “13” identifies the element as aluminum, and the left superscript of “27” indicates the specific isotope of aluminum, aluminum-27. The ion is therefore Al^{3+} . The periodic table shows that aluminum is in the third period and Group 13.

2. CHEMICAL FORMULAS

Conceptual Problems

1. One difference is that a polyatomic ions possess a charge, such as $-$, $2+$, or $3-$, whereas a molecule is not charged. A second difference is that polyatomic ions can form crystal lattice solids, whereas molecules form solids with weak intermolecular forces. Two similarities are that both contain two or more atoms and that both are held together by covalent bonds.
2. Those compounds that contain metals and nonmetals are ionic, and those compounds that contain only nonmetals are covalent. Thus, (a), (c), and (e) are ionic, and (b), (d), and (f) are covalent.
3. Those compounds that contain metals and nonmetals are ionic, and those compounds that contain only nonmetals are covalent. Thus, (b), (c), and (e) are ionic, and (a) and (d) are covalent. The criteria for organic and inorganic compounds were detailed in Section 2.1 in the text. Based on these criteria, (a) is organic, and the others are inorganic. (e) is tricky because it contains carbon, but since an ionic bond exists between the Na^+ and CH_3COO^- ions, we classify it as inorganic.
4. The molecular or molar mass is needed in addition to the empirical formula. Since the empirical formula is the smallest whole number ratio of elements in a compound, several compounds can share the same empirical formula. For example, benzene (C_6H_6) and acetylene (C_2H_2) both have the empirical formula CH. Each formula, however, is an integral multiple of the empirical formula. Thus, we can write *molecular formula* = (*empirical formula*)_{*n*}, where *n* is an integer. For benzene *n* = 6. For acetylene *n* = 2. An alternative formula is
$$\text{molecular formula mass} = n(\text{empirical formula mass})$$
where *n* is the same *n* as above.
5. A structural formula gives us both the number of atoms in the molecule and their connectivity, neither of which can be obtained from an empirical formula.
6. As we shall see, the presence of an $-\text{OH}$ group dominates the properties and chemistry of most alcohols. Writing the formula of an alcohol as ROH emphasizes the similarities in chemistry of different alcohols.
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 $(\text{CH}_3)_2\text{S}$.
8. Magnesium hydroxide is an ionic compound. 1 formula unit contains one Mg^{2+} ion and 2 OH^- ions. To obtain an electrically neutral compound, two hydroxide ions are needed to balance the one magnesium ion. Therefore, the correct formula for magnesium hydroxide is $\text{Mg}(\text{OH})_2$, where the parentheses, absent in MgOH_2 , indicate the presence of the hydroxide ions.
9. Magnesium cyanide is an ionic compound. 1 formula unit contains one Mg^{2+} ion and two CN^- ions. To obtain an electrically neutral compound, two cyanide ions are needed to balance the one magnesium ion. Therefore, the correct formula for magnesium cyanide is $\text{Mg}(\text{CN})_2$, where the parentheses, absent in MgCN_2 , indicate the presence of the cyanide ions.
10. Yes. For example, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ will always have the same ratio of water molecules to one formula unit of CuSO_4 . The CaSO_4 example discussed at the end of this section, however, shows that a particular compound can form several hydrated compounds with a different numbers of waters.

Numerical Problems

1. Compounds (a), (d), and (e) are ionic compounds, whose formulas are written as cation followed by anion. Compounds (b) and (c) are organic compounds, whose formulas are written as carbon followed by hydrogen and then other elements in alphabetical order. Thus:
 - a. MgSO_4
 - b. $\text{C}_2\text{H}_6\text{O}_2$
 - c. $\text{C}_2\text{H}_4\text{O}_2$
 - d. KClO_3
 - e. $\text{NaOCl} \cdot 5\text{H}_2\text{O}$
2. Compounds (a), (d), and (e) are ionic compounds, whose formulas are written as cation followed by anion. Compounds (b) and (c) are organic compounds, whose formulas are written as carbon followed by hydrogen and then other elements in alphabetical order. Thus:
 - a. $\text{Cd}(\text{C}_2\text{H}_3\text{O}_2)_2$
 - b. $\text{Ba}(\text{CN})_2$
 - c. $\text{FePO}_4 \cdot 2\text{H}_2\text{O}$
 - d. $\text{Mn}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$
 - e. Ag_3PO_4

3. In each case the total charges of all ions in the compound must add up to zero. Further, be sure to include parentheses only when more than one polyatomic ion is in the formula. The table is

Ion	K ⁺	Fe ³⁺	NH ₄ ⁺	Ba ²⁺
Cl ⁻	KCl	FeCl ₃	NH ₄ Cl	BaCl ₂
SO ₄ ²⁻	K ₂ SO ₄	Fe ₂ (SO ₄) ₃	(NH ₄) ₂ SO ₄	BaSO ₄
PO ₄ ³⁻	K ₃ PO ₄	FePO ₄	(NH ₄) ₃ PO ₄	Ba ₃ (PO ₄) ₂
NO ₃ ⁻	KNO ₃	Fe(NO ₃) ₃	NH ₄ NO ₃	Ba(NO ₃) ₂
OH ⁻	KOH	Fe(OH) ₃	NH ₄ OH	Ba(OH) ₂

4. The “most common” ion is the only ion for all of the elements in this problem. For (d) we estimate that Si forms Si⁴⁺ ions. Si is a semimetal, however, and to say it forms an ionic compound with O²⁻ ions is an oversimplification. Thus,
- ZnS
 - BaI₂
 - MgCl₂
 - SiO₂
 - Na₂S
5. The “most common” ion is the only ion for all of the elements in this problem. For (c) we estimate that Ge forms Ge⁴⁺ ions. Ge is a semimetal, however, and to say it forms an ionic compound with O²⁻ ions is an oversimplification. We should, likewise, be careful with As, another semimetal. Thus,
- Li₃N
 - CsCl
 - GeO₂
 - Rb₂S
 - Na₃As
6. An empirical formula is the smallest whole number ratio of elements in a compound. Thus, if there is a common denominator (other than 1, of course) in the subscripts of a compound, the empirical formula has not been found. For (a) there is a common denominator (cd) of 2. Dividing each subscript by 2 gives NaSO₂ as the empirical formula. Thus,
- cd = 2 NaSO₂
 - cd = 2 BH₃
 - cd = 6 CH₂O
 - cd = 2 P₂O₅
 - cd = 1 KMnO₄
7. An empirical formula is the smallest whole number ratio of elements in a compound. Thus, if there is a common denominator (other than 1, of course) in the subscripts of a compound, the empirical formula has not been found. For (a) there is a common denominator (cd) of 2. Dividing each subscript by 2 gives AlCl₃ as the empirical formula. Thus,
- cd = 2 AlCl₃
 - cd = 1 K₂Cr₂O₇
 - cd = 2 CH₂
 - cd = 1 CH₅N₃, following the usual order of C, H, others in alphabetical order
 - cd = 2 CH₂O

3. NAMING IONIC COMPOUNDS

Conceptual Problems

- Metals that form only one kind of cation simply retain the name of the neutral atom. Rule 2(a). Thus, adding “ion” to the neutral names, we find: (a) potassium ion; (b) aluminum ion; (c) ammonium ion; (d) magnesium ion; (e) lithium ion.
- Monatomic anions are named by adding the suffix *-ide* to the root of the name of the parent element. Rule 3(a). Thus, the names are (a) bromide; (b) carbonate; (c) sulfide; (d) nitrate; (e) formate; (f) fluoride; (g) hypochlorite; (h) oxalate.
- Monatomic anions are named by adding the suffix *-ide* to the root of the name of the parent element. Rule 3(a). Thus, the names are (a) phosphate; (b) chloride; (c) sulfite; (d) acetate; (e) hydrogen sulfate, or bisulfate; (f) perchlorate; (g) nitrite; (h) oxide.
- Monatomic anions are named by adding the suffix *-ide* to the root of the name of the parent element. Rule 3(a). Thus, the names are (a) sulfate; (b) cyanide; (c) dichromate; (d) nitride; (e) hydroxide; (f) iodide; (g) peroxide.
- All of the metal ions here form only one type of ion. The anions are a mixture of monatomic and polyatomic ions; N_3^- is a polyatomic ion, the azide ion. Lastly, all of these ions are singly charged, $(-)$ or $(+)$. Thus,
 - magnesium bromide
 - ammonium cyanide
 - calcium oxide
 - potassium chlorate
 - potassium phosphate
 - ammonium nitrite
 - sodium azide
- All of the metal ions here form only one type of ion, except copper. For copper, we indicate the charge on the copper ion using Roman numerals in parentheses. The anions are a mixture of monatomic and polyatomic ions. Lastly, some of the ions here are singly charged, $(-)$ or $(+)$, while others are multiply charged, $(n-)$ or $(n+)$. Be certain to identify the polyatomic ions correctly. Thus,
 - sodium nitrate
 - copper(II) phosphate
 - sodium hydroxide
 - lithium carbide
 - calcium fluoride
 - ammonium bromide
 - magnesium carbonate
- All of the metal ions here form only one type of ion, except manganese. For manganese, we indicate the charge on the manganese ion using Roman numerals in parentheses. The anions are a mixture of monatomic and polyatomic ions. Lastly, some of the ions here are singly charged, $(-)$ or $(+)$, while others are multiply charged, $(n-)$ or $(n+)$. Be certain to identify the polyatomic ions correctly. Note that the iodate ion is IO_3^- . Thus,
 - rubidium bromide
 - manganese(III) sulfate
 - sodium hypochlorite
 - ammonium sulfate
 - sodium bromide
 - potassium iodate
 - sodium chromate
- All of the metal ions here form only one type of ion, except tin and iron. For tin and iron, we indicate the charge on the metal ions using Roman numerals in parentheses. The anions are a mixture of monatomic and polyatomic ions. Lastly, some of the ions here are singly charged, $(-)$ or $(+)$, while others are multiply charged, $(n-)$ or $(n+)$. Be certain to identify the polyatomic ions correctly. Thus,
 - ammonium perchlorate
 - tin(IV) chloride
 - iron(II) hydroxide
 - sodium oxide
 - magnesium chloride
 - potassium sulfate
 - radium chloride
- All of the metal ions here form only one type of ion, except for nickel. Nickel is present as Ni^{2+} , its most common ion. The anions are all polyatomic ions, except for Cl^- . Lastly, some of the ions here are singly charged, $(-)$ or $(+)$, while others are multiply charged, $(n-)$ or $(n+)$. Be certain to identify the polyatomic ions correctly. Thus,
 - potassium cyanide
 - lithium hydroxide
 - calcium chloride
 - nickel(II) sulfate
 - ammonium chlorite
 - lithium perchlorate
 - lanthanum cyanide

Numerical Problems

- When naming the ions in these compounds, remember that the cations are always written first and the anions second in the formula. Further, always be mindful of polyatomic ions. For example, every compound in this problem contains only *two* ions even if a compound has three elements present.
 - beryllium ion, Be^{2+} oxide ion, O^{2-}
 - lead(II) ion, Pb^{2+} hydroxide ion, OH^-
 - barium ion, Ba^{2+} sulfide, S^{2-}
 - sodium ion, Na^+ dichromate, $\text{C}_2\text{O}_7^{2-}$
 - zinc ion, Zn^{2+} sulfate ion, SO_4^{2-}
 - potassium ion, K^+ hypochlorite ion, ClO^-
 - sodium ion, Na^+ dihydrogen phosphate, H_2PO_4^-
- When naming the ions in these compounds, remember that the cations are always written first and the anions second in the formula. Further, always be mindful of polyatomic ions. For example, every compound in this problem contains only *two* ions even if a compound has three elements present.
 - zinc ion, Zn^{2+} nitrate ion, NO_3^-
 - cobalt(II) ion, Co^{2+} sulfide ion, S^{2-}
 - beryllium ion, Be^{2+} carbonate ion, CO_3^{2-}
 - sodium ion, Na^+ sulfate ion, SO_4^{2-}
 - potassium ion, K^+ oxalate ion, $\text{C}_2\text{O}_4^{2-}$
 - sodium ion, Na^+ cyanide ion, CN^-
 - iron(II) ion, Fe^{2+} chloride ion, Cl^-
- For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “lead(IV)” and “titanium(IV)” immediately tells you the charge of each ion. Assume that Si^{4+} ions are present. The formulas are
 - MgCO_3
 - $\text{Al}_2(\text{SO}_4)_3$
 - K_3PO_4
 - PbO_2
 - Si_3N_4
 - NaClO
 - TiCl_4
 - $\text{Na}_2\text{NH}_4\text{PO}_4$
- For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “lead(II)” immediately tells you the charge of the lead ion. The formulas are
 - $\text{Pb}(\text{NO}_3)_2$
 - $(\text{NH}_4)_3\text{PO}_4$
 - Ag_2S
 - BaSO_4
 - CsI
 - NaHCO_3
 - $\text{K}_2\text{Cr}_2\text{O}_7$
 - NaClO
- For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “lead(II)” and “copper(II)” immediately tells you the charge of each ion. The formulas are
 - $\text{Zn}(\text{CN})_2$
 - Ag_2CrO_4
 - PbI_2
 - C_6H_6
 - $\text{Cu}(\text{ClO}_4)_2$
- For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “iron(III)” and “copper(II)” immediately tells you the charge of each ion. The formulas are
 - CaF_2
 - NaNO_3
 - Fe_2O_3
 - $\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$
 - NaNO_2

7. For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “nickel(II)” and “titanium(IV)” immediately tells you the charge of each ion. (Ni(II) is nickel’s most common oxidation state.). The formulas are
- NaOH
 - Ca(CN)₂
 - Mg₃(PO₄)₂
 - Na₂SO₄
 - NiBr₂
 - Ca(ClO₂)₂
 - TiBr₄
8. For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “tin(II)” and “iron(II)” immediately tells you the charge of each ion. The azide ion is N₃⁻. The formulas are
- NaClO₂
 - KNO₂
 - NaN₃
 - Ca₃P₂
 - SnCl₂
 - CaHPO₄
 - FeCl₂•2H₂O
9. For any neutral formula the sum of the charges must equal zero. Additionally, include parentheses only when more than one polyatomic ion is in the formula. Chemists label metal/nonmetal compounds as ionic because the metals donate electrons, forming cations, to the nonmetals in these compounds. For example, groups 1, 2, and 13 donate electrons, and the nonmetals from groups 15, 16, and 17 accept electrons, forming ions with charges 1+, 2+, 3+, 3-, 2-, and 1-, respectively. Transition metals form cations, too, but with much more widely varying charge states. The notation “chromium(III)”, “cobalt(II)”, and “nickel(II)” immediately tells you the charge of each ion. The formulas are
- K₂CO₃
 - Cr₂(SO₃)₃
 - Co₃(PO₄)₂
 - Mg(ClO)₂
 - Ni(NO₃)₂•6H₂O

4. NAMING COVALENT COMPOUNDS

Conceptual Problems

- While the mass of a molecule or ion can be a useful predictor of some properties of a substance, the nature and strength of the intermolecular or interionic interactions is much more important. Benzene is an organic compound that contains individual C_6H_6 molecules in the solid, liquid, and vapor phases. The individual molecules are held together by only weak intermolecular interactions. In fact, benzene is a liquid at room temperature. In contrast, KCl is an ionic compound, and strong interionic interactions cause it to form a crystal lattice that consists of a regular array of alternating positive and negative ions. We can therefore predict that:
 - The melting point of KCl will be much higher than that of benzene. (The actual values are 771°C and 5.5°C , respectively.)
 - Benzene is a covalent organic compound; KCl is an ionic inorganic compound.
 - Due to the presence of weak intermolecular forces, at room temperature benzene molecules can escape from a sample of liquid benzene much more readily than K^+ and Cl^- ions can escape from a KCl crystal. Thus, the rate of evaporation of benzene will be many orders of magnitude greater than that of KCl.
 - Benzene: weak intermolecular interactions cause it to be a volatile liquid at room temperature.
KCl: strong interionic interactions result in formation of an ordered crystal lattice.
- Be careful not to confuse “organic” compounds with “hydrocarbons”. Hydrocarbons contain *only* C and H. Since inorganic compounds must contain something *other than* C or H, it is not possible to classify an inorganic compound as a hydrocarbon.
- Hydrocarbons contain *only* C and H. Since this compound contains Na and O, it is not a hydrocarbon.
- Compounds (a) and (b) are inorganic compounds that are best viewed as ionic, and they are named using the rules for ionic compounds. Note that Ni and Ti can form more than one kind of positive ion, so their charge must be indicated. Compounds (c) through (g) are covalent compounds, and are named using the rules for covalent compounds. Thus:
 - nickel(II) oxide
 - titanium(IV) oxide
 - dinitrogen monoxide
 - carbon disulfide
 - sulfur trioxide
 - nitrogen trifluoride
 - sulfur hexafluoride
- Compounds (a) and (e) are inorganic compounds that are best viewed as ionic, and they are named using the rules for ionic compounds. Note that Hg can form more than one kind of positive ion, so its charge must be indicated. Compounds (b), (c), (d), and (f) are covalent compounds, and are named using the rules for covalent compounds. Thus:
 - mercury(II) chloride
 - iodine pentafluoride
 - dinitrogen pentoxide
 - dichlorine monoxide
 - mercury(II) sulfide
 - phosphorus pentachloride
- To identify the atoms in the three-dimensional models, you will have to refer to the color code used for different elements in this and the first section of this chapter.
 - Light purple indicates silicon (Si), and dark green is chlorine (Cl). The molecule is therefore silicon tetrachloride, SiCl_4 .
 - Oxygen is red and sulfur is yellow, so this is sulfur trioxide, SO_3 .
 - Oxygen is red and nitrogen is blue, so this is dinitrogen tetroxide, N_2O_4 .
 - This molecule contains a linear chain of five carbon atoms, with a double bond between the first and second (going counterclockwise from the top left). It also contains 10 hydrogen atoms. This is 1-pentene, C_5H_{10} .
 - This molecule contains a linear chain of four carbon atoms with an $-\text{OH}$ attached to one of the terminal carbon atoms. It also contains 9 hydrogen atoms bonded to carbon. This is *n*-butanol, $n\text{-C}_4\text{H}_9\text{OH}$.
- To identify the atoms in the three-dimensional models, you will have to refer to the color code used for different elements in this and the first section of this chapter.
 - Dark green represents chlorine and light green is fluorine. The molecule is therefore chlorine trifluoride, ClF_3 .
 - Dark green is chlorine and yellow is sulfur. The molecule is therefore disulfur dichloride, S_2Cl_2 .
 - Nitrogen is blue and oxygen is red. The molecule is therefore dinitrogen monoxide, N_2O (also known as nitrous oxide).
 - This molecule contains a linear chain of three carbon atoms with a triple bond between two of them; it also contains four hydrogen atoms. This is propyne, C_3H_4 .
- So far, we have said that metals join with nonmetals to form ionic compounds and that nonmetals bond with other nonmetals to form covalent compounds. Since P and Cl are both nonmetals, PCl_3 is predicted to be a covalent compound.
- Aliphatic hydrocarbons are any hydrocarbons that are not aromatic. For our purposes, aromatic compounds are those that contain carbon rings with alternating single and double bonds, similar to those in benzene.

10. C_nH_{2n+2} , C_nH_{2n} , and C_nH_{2n-2} are the general formulas for alkanes, alkenes, and alkynes, respectively. You should confirm these formulas for yourself using the various compounds mentioned in this section.

Numerical Problems

- The formulas of these compounds can be deduced from the prefixes indicating numbers of atoms and the rules for naming covalent compounds.
 - N_2O
 - SiF_4
 - BCl_3
 - NF_3
 - PBr_3
- The formulas of these compounds can be deduced from the prefixes indicating numbers of atoms and the rules for naming covalent compounds.
 - N_2O_3
 - IF_5
 - BBr_3
 - OF_2
 - $AsCl_3$
- The formulas of these compounds can be deduced from the prefixes indicating numbers of atoms and the Roman numerals indicating the charge on the metal ions.
 - Tl_2Se
 - NpO_2
 - FeS
 - $CuCN$
 - NCl_3
- All of the metals in these compounds can form more than one kind of positive ion, so their charge must be indicated by Roman numerals in parentheses.
 - ruthenium(VIII) oxide
 - lead(IV) oxide
 - molybdenum(VI) fluoride
 - mercury(I) nitrate dihydrate
 - tungsten(IV) chloride
- All of the metals in these compounds can form more than one kind of positive ion, so their charge must be indicated by Roman numerals in parentheses. P_4S_{10} is a covalent compound and is named accordingly.
 - niobium(IV) oxide
 - molybdenum(IV) sulfide
 - tetraphosphorus decasulfide
 - copper(I) oxide
 - rhenium(V) fluoride

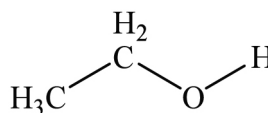
- ROH (where R is an alkyl group)
 - ROH (where R is an aryl group)

6. Remember that each carbon atom will make four bonds, each oxygen atom will make two bonds, and each hydrogen atom will make one bond. Often the hydrogen bonds to carbon are not shown with a line. Finally, since we are given the number of carbon atoms, our job is much easier. The structures are

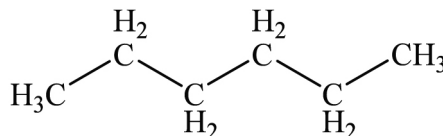
- CH_3CCH



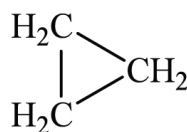
- CH_3CH_2OH



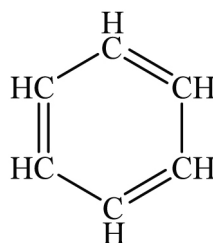
- $CH_3(CH_2)_4CH_3$



- C_3H_6

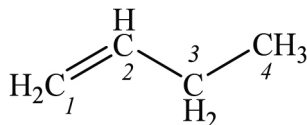


- C_6H_6

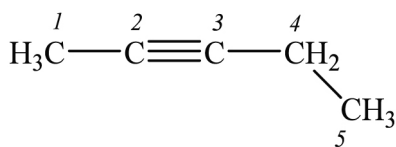


7. Remember that each carbon atom will make four bonds, each oxygen atom will make two bonds, and each hydrogen atom will make one bond. Often the hydrogen bonds to carbon are not shown with a line. Since we are given the number of carbon atoms, our job is much easier.

- a. This 4-carbon *alkene* has a double bond that appears after the first carbon, as indicated by the “1-” in front of the name. The double bond is placed so as to make the number in front of the name as small as possible. The carbon atoms are numbered in italics. Note that the erroneous “4-butene” is just 1-butene that has been numbered incorrectly, but that 2-butene is a different compound.

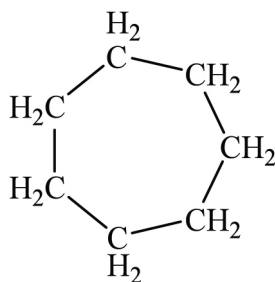


- b. This 5-carbon *alkyne* has a triple bond that appears after the second carbon, as indicated by the “2-” in front of the name. The triple bond is placed so as to make the number in front of the name as small as possible. The carbon atoms are numbered in italics. Note that the erroneous “4-pentyne” is just 2-pentyne that has been numbered incorrectly, but that 1-pentyne is a different compound.

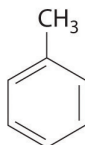


Note that alkynes are linear at the triple bond.

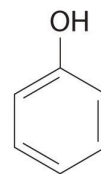
- c. This is a simple cyclic alkane. No numbering or other dilemmas.



- d. Toluene is simply benzene in which one hydrogen atom is replaced by a methyl group: $C_6H_5CH_3$. The structure is:



- e. Phenol is simply benzene in which one hydrogen atom is replaced by an $-OH$ group: C_6H_5OH . The structure is:

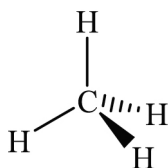


5. ACIDS AND BASES

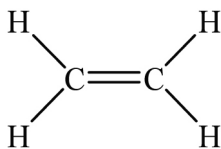
Conceptual Problems

- The chief distinction to consider when naming acids is whether the H atom is attached to an oxygen atom or an atom other than O. In oxoacids the H atom is attached to an O atom.
 - hydrochloric acid
 - bromic acid
 - nitric acid
 - sulfuric acid
 - iodic acid
- The chief distinction to consider when naming acids is whether the H atom is attached to an oxygen atom or an atom other than O. In oxoacids the H atom is attached to an O atom.
 - hydrobromic acid
 - sulfurous acid
 - chloric acid
 - hydrocyanic acid
 - phosphoric acid
- When hydrogen bromide gas dissolves in water, hydrobromic acid forms. Likewise for (b) and (c). Thus,
 - hydrobromic acid
 - hydrocyanic acid
 - hydroiodic acid
- This molecule contains a linear chain of 5 carbon atoms, with a carboxylic acid group on one end. The prefix for “5” is “penta-”, so this is a pentanoic acid. To indicate that the carbon chain is not branched, we can use the *n* for *normal* designation, giving *n*-pentanoic acid.
 - The pale orange color of the central atom indicates that it is phosphorus, with three –OH groups attached to phosphorus. The compound is H_3PO_3 , phosphorous acid.
- This molecule contains a 2-carbon chain attached to a nitrogen atom (shown in blue), which makes it an amine. Seven additional hydrogen atoms are attached to carbon and nitrogen. It is $\text{CH}_3\text{CH}_2\text{NH}_2$ (often written as $\text{C}_2\text{H}_5\text{NH}_2$), ethylamine.
 - The purple color of the central atom indicates that it is iodine, and there are two oxygen atoms and one –OH group attached to iodine. It is therefore HIO_3 , iodic acid.
- Compounds (b) and (d) are ionic compounds containing the OH^- ion, so they are strong bases. Compound (f) is sulfuric acid, a strong acid. Compounds (c) and (e) are both carboxylic acids, which dissociate in water to produce H^+ ions. Compound (a) is an alcohol, ethanol – alcohols are covalent compounds that do not dissociate in water to produce either H^+ or OH^- ions. We therefore predict that the solutions will be:
 - neutral
 - basic
 - acidic
 - basic
 - acidic
 - acidic

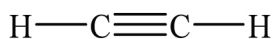
7. a. CH_4 , methane



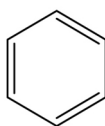
- b. C_2H_4 , ethylene



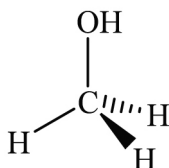
- c. C_2H_2 , acetylene



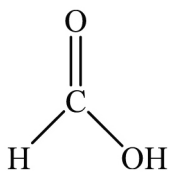
- d. C_6H_6 , benzene



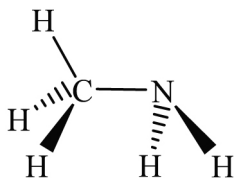
- e. CH_3OH , methanol



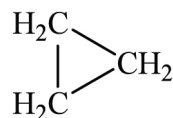
- f. HCOOH , formic acid



- g. methyl amine, CH_3NH_2



- h. C_3H_6 , cyclopropane



8. a. cycloalkane
b. alcohol
c. alkyne
d. cycloalkene
e. amine
f. alkene
g. alcohol
h. carboxylic acid
9. a. amine
b. alcohol
c. cycloalkane
d. cycloalkene
e. carboxylic acid
f. alkyne
g. amine
h. carboxylic acid

Numerical Problems

1.
 - a. HClO
 - b. HBrO_4
 - c. HBr
 - d. H_2SO_3
 - e. NaBrO_4
2.
 - a. HI
 - b. H_2S
 - c. H_3PO_3
 - d. HClO_4
 - e. $\text{Ca}(\text{BrO})_2$
3.
 - a. hydrobromic acid
 - b. sulfurous acid
 - c. hydrocyanic acid
 - d. perchloric acid
 - e. sodium hydrogen sulfate
4.
 - a. sulfuric acid
 - b. nitrous acid
 - c. potassium hydrogen phosphate
 - d. phosphorous acid
 - e. calcium dihydrogen phosphate monohydrate

6. INDUSTRIALLY IMPORTANT CHEMICALS

Conceptual Problems

1. Crude oil is converted into transportation fuels in three steps: distillation, cracking, reforming. Crude oil contains many organic compounds, many of which can be used for fuels. Since these compounds have different boiling points, crude oil is heated, separating the various components in the crude oil sample into *fractions*. This separation process is not ideal, however, and the less volatile, less valuable fractions (such as kerosene) are often converted into more volatile, more valuable hydrocarbons (such as gasoline) by “cracking” the larger, less volatile fractions into smaller, more volatile fractions. “Reforming” also enhances the value of the fractions that are obtained from distillation by converting straight chain alkanes into more valuable, branched-chain alkanes or mixtures of aromatic hydrocarbons. The “more valuable” products are those that can be sold as widely-used fuels, such as gasoline, *the* lifeblood of the world’s economy.
2. Gasoline with an octane rating that is too low for the engine can ignite prematurely, resulting in uneven burning that causes knocking and pinging.
3. Lead is released into the environment by the burning of tetraethyllead. This, of course, greatly increases the chances of ingestion of lead, a highly toxic metal. Further, lead destroys catalytic converters, which are designed to prevent air pollution. This is costly financially and environmentally.
4. The Frasch process is the method used today to extract elemental sulfur that is found in porous limestone rocks. In this process water at 160 °C and high pressure is pumped underground, melting the sulfur. Compressed air is then used to bring the sulfur-water mixture to the surface. This process procures 99% pure sulfur from the underground deposit.
5. Phosphate salts contain the highly-charged PO_4^{3-} ion, salts of which are often insoluble. Protonation of the PO_4^{3-} ion by strong acids such as H_2SO_4 leads to the formation of the HPO_4^{2-} and H_2PO_4^- ions. Because of their decreased negative charge, salts containing these anions are usually much more soluble, allowing the anions to be readily taken up by plants when they are applied as fertilizer.

Numerical Problem

1. As in the worked example, there are only two components, allowing us to set up the following equation, where x is the fraction of MTBE in the solution:

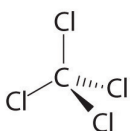
$$\begin{aligned}\text{final octane rating of mixture} &\times 100\% \\ &= 52(100) \\ &= (116)x + (0)(100 - x) \\ &= 116x\end{aligned}$$

Solving the equation $5200 = 116x$ gives
 $x = 5200 / 116 = 45\%$ MTBE. The percentage of *n*-heptane is, therefore, $100\% - x = 55\%$.

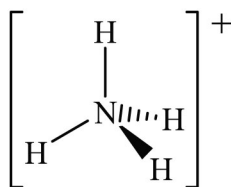
7. END-OF-CHAPTER MATERIAL

Application Problems

1. CCl_4 is just CH_4 with the H's replaced by Cl's. The structure is provided.



2. The ammonium ion is charged, and the structure reflects this with the (+) sign outside of a set of brackets. Each H makes its usual one bond, but N will make four bonds here instead of its usual three. The structure is geometrically the same as CH_4 , CH_2Cl_2 , CCl_4 , and so on, and it is provided:



3. There are only two reactants here, Mg and O_2 . (Remember that oxygen gas is a diatomic molecule.) Based on this one might (correctly) guess 'magnesium oxide' without much thought. The material presented so far, however, indicates that we can explain why this is so; this is a reaction between a *metal* and a *nonmetal*. Thus, ion formation is expected, and it should be clear by now that the ions are Mg^{2+} and O^{2-} , giving MgO as the formula of the reaction product.
4. Sodium hydrogen sulfite is an ionic compound, and sulfur dioxide is a covalent compound. Thus, the formulas are NaHSO_3 and SO_2 . (NaHSO_3 contains Na^+ and HSO_3^- ions.)
5. Carbonic acid is H_2CO_3 ; lithium carbonate is Li_2CO_3 .
6. If acetic acid is an organic acid, we expect to find the carboxylic acid group, $-\text{CO}_2\text{H}$, in the molecule. The other group present is a methyl group, $-\text{CH}_3$, which is attached to the carboxylic acid group to give $\text{CH}_3-\text{CO}_2\text{H}$.
7. a. Sodamide is NaNH_2 , and sodium cyanide is NaCN .
b. Sodium hydroxide (NaOH) and ammonia (NH_3).

8. If *n*-pentane comprises 30% of the solution, isooctane and *n*-heptane must comprise the other 70%. Therefore, we will have factors of x and $(70 - x)$. The following equation can be set up with x being the percentage of isooctane in the solution

$$\begin{aligned}\text{final octane rating of mixture} \times 100\% \\ &= 87(100) \\ &= 62(30) + (100)x + (0)(70 - x) \\ &= 1860 + 100x\end{aligned}$$

Dividing both sides of $8700 = 1860 + 100x$ by 100 gives $87 = 18.6 + x$, which is easily solved to give % isooctane = 68.4%. Thus,
% *n*-heptane = $70 - 68.4 = 1.6\%$.

9. a. We are given the % by mass of *n*-pentane and methanol, 60% and 25%, respectively. Therefore, the % by mass of *n*-hexane is $100 - 60 - 25 = 15\%$. With this information and the tabulated octane ratings, we can set up an equation to calculate the final octane rating of the mixture:

$$\begin{aligned}\text{final octane rating of mixture} \times 100\% \\ &= x(100) \\ &= 62(60) + 107(25) + 25(15) \\ &= 3720 + 2675 + 375 \\ &= 6770\end{aligned}$$

Thus, $x = 6770 / 100 \approx 68$.

- b. The mixture of part (a) has an octane rating of 68. We can now treat the solution of *n*-pentane, methanol, and *n*-hexane as one component of the two component mixture of part (b) for which an octane rating of 93 is desired. Thus, we can write, with x being the percentage of MTBE in the solution,

$$\begin{aligned}\text{final octane rating of mixture} \times 100\% \\ &= 93(100) \\ &= 68(100 - x) + 116(x) \\ &= 6800 - 68x + 116x \\ &= 6800 + 48x\end{aligned}$$

Solving this for x gives $x = (9300 - 6800) / 48 = 52\%$ MTBE *by mass*. This means that for *any* sample size 52% of the mass must be MTBE and 48% of the mass must be the three component solution of part (a). As an example, adding 52 g of MTBE to 48 g of the solution of part (a) would yield a solution with an octane rating of 93.

10. Ethanol has an octane rating of 108. The other component of the premium gasoline constitutes 90% of the solution, and, if we allow x to represent the octane rating of the “other component”, we can write,

$$\begin{aligned}\text{final octane rating of mixture} &\times 100\% \\ &= 93(100) \\ &= 108(10) + x(90) \\ &= 1080 + 90x\end{aligned}$$

Solving the final equation for x gives

$$x = (9300 - 1080) / 90 \approx 91.$$

