Chapter 8. Basic Concepts of Chemical Bonding

Media Resources

Figures and Tables in Transparency Pack:
- Figure 8.2 Reaction of Sodium Metal with Chlorine Gas to Form the Ionic Compound Sodium Chloride
- Table 8.2 Lattice Energies for Some Ionic Compounds
- Figure 8.5 Born-Haber Cycle for Formation of NaCl
- Figure 8.7 Electronegativity Values Based on Pauling’s Thermochemical Data
- Figure 8.8 Electron Density Distribution
- Figure 8.11 Oxidation Number, Formal Charge, and Electron Density Distribution for the HCl Molecule
- Table 8.4 Average Bond Enthalpies (kJ/mol)
- Figure 8.15 Using Bond Enthalpies to Calculate $\Delta H_{\text{rxn}}$
- Table 8.5 Average Bond Lengths for Some Single, Double, and Triple Bonds

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- Octet Rule
- Coulomb’s Law
- Ion Electron Configurations
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- Periodic Trends: Electronegativity
- Formal Charges

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- Formation of Sodium Chloride

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- 8.2 Ionic Bonding

3-D Models:
- Chlorine
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Section:
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- 8.2 Ionic Bonding
Methane
Carbon dioxide
Hydrogen Chloride
Ozone
Benzene
Phosphorus Pentachloride
Chloromethane

**Other Resources**

**Further Readings:**
The Chemical Bond as an Atomic Tug-of-War
Gilbert Newton Lewis and the Amazing Electron Dots
The Chemical Bond
Grade-12 Students’ Misconceptions of Covalent Bonding and Structure
The Role of Lewis Structures in Teaching Covalent Bonding
Abegg, Lewis, Langmuir, and the Octet Rule
G. N. Lewis and the Chemical Bond
The Use of Dots in Chemical Formulas
Electronegativity and Bond Type: Predicting Bond Type
Electronegativity from Avogadro to Pauling Part I: Origins of the Electronegativity Concept
Electronegativity from Avogadro to Pauling Part II: Late Nineteenth- and Early Twentieth-Century Developments
Electron Densities, Atomic Charges, and Ionic, Covalent, and Polar Bonds
Drawing Lewis Structures from Lewis Symbols: A Direct Electron Pairing Approach
Lewis Structures Are Models for Predicting Molecular Structure, Not Electronic Structure
Teaching a Model for Writing Lewis Structures
Drawing Lewis Structures without Anticipating Octets
The ‘6N + 2 Rule’ for Writing Lewis Octet Structures
Another Procedure for Writing Lewis Structures
Using Formal Charges in Teaching Descriptive Inorganic Chemistry

**Section:**
8.1 Lewis Symbols, and the Octet Rule
8.2 Ionic Bonding
8.3 Covalent Bonding
8.4 Bond Polarity and Electronegativity
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Lewis Structures, Formal Charge, and Oxidation Numbers: A More User-Friendly Approach

Valence, Oxidation Number, and Formal Charge: Three Related but Fundamentally Different Concepts

Lost in Lewis Structures: An Investigation of Student Difficulties in Developing Representational Competence

If It’s Resonance, What Is Resonating?

Aromatic Bagels: An Edible Resonance Analogy

The Concept of Resonance

Explaining Resonance—A Colorful Approach

A Visual Aid for Teaching the Resonance Concept

The Origin of the Circle Symbol for Aromaticity

Nitric Oxide—Some Old and New Perspectives

Biological Roles of Nitric Oxide

The Relative Explosive Power of Some Explosives

Exothermic Bond Breaking: A Persistent Misconception

8.5 Drawing Lewis Structures

8.6 Resonance Structures

8.7 Exceptions to the Octet Rule

8.8 Strengths of Covalent Bonds
Chapter 8. Basic Concepts of Chemical Bonding

Common Student Misconceptions

• Students often think that a triple bond is three times as strong as a single bond. The fact that the second and third bonds (π bonds) are weaker than the first (σ bond) needs to be emphasized.
• Students confuse formal charges with real charges on atoms.
• The only place the ↔ arrow is used is for resonance; students often want to use this to indicate equilibrium.
• Students do not appreciate that the exceptions to the octet rule are almost as common as the examples of substances that obey it.
• Students often confuse the octet rule with having any 8 valence electrons (e.g., 4s² 3d⁶ for iron).
• Students often think that all polar substances conduct electricity.

Teaching Tips

• Students need to be able to count the number of valence electrons in order to get the correct Lewis structure.
• Students need to be reminded that several correct Lewis structures can be often drawn for a molecule or an ion, but not all correct Lewis structures are equally good.

Lecture Outline

8.1 Lewis Symbols and the Octet Rule

• The properties of many materials can be understood in terms of their microscopic properties.
• Microscopic properties of molecules include:
  • the connectivity between atoms and
  • the 3-D shape of the molecule.
• When atoms or ions are strongly attracted to one another, we say that there is a chemical bond between them.
  • In chemical bonds, electrons are shared or transferred between atoms.
• Types of chemical bonds include:
  • ionic bonds (electrostatic forces that hold ions together, e.g., NaCl);
  • covalent bonds (result from sharing electrons between atoms, e.g., Cl₂);
  • metallic bonds (refers to metal nuclei floating in a sea of electrons, e.g., Na).
• The electrons involved in bonding are called valence electrons.
  • Valence electrons are found in the incomplete, outermost shell of an atom.
• As a pictorial understanding of where the electrons are in an atom, we represent the electrons as dots around the symbol for the element.
  • The number of valence electrons available for bonding are indicated by unpaired dots.
  • These symbols are called Lewis symbols or Lewis electron-dot symbols.
  • We generally place the electrons on four sides of a square around the element’s symbol.

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1 “The Chemical Bond as an Atomic Tug-of-War” from Further Readings
2 “Chlorine” 3-D Model from Instructor’s Resource CD/DVD
3 “Phosphorus” 3-D Model from Instructor’s Resource CD/DVD
The Octet Rule

- Atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons; this is known as the **octet rule**.
  - An octet consists of full s and p subshells.
  - We know that \( ns^2np^6 \) is a noble gas configuration.
  - We assume that an atom is stable when surrounded by eight electrons (four electron pairs).

**FORWARD REFERENCES**
- The octet rule will be brought up again in Chapters 22 and 23 for main group nonmetals and metals, respectively, and in Chapter 24 for carbon.

8.2 Ionic Bonding

- Consider the reaction between sodium and chlorine:
  \[ \text{Na}(s) + \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{NaCl}(s) \quad \Delta H_f = -410.9 \text{ kJ/mol} \]
  - The reaction is violently exothermic.
  - We infer that the NaCl is more stable than its constituent elements.
    - Sodium has lost an electron to become Na⁺ and chlorine has gained the electron to become Cl⁻.
    - Note that Na⁺ has an Ne electron configuration and Cl⁻ has an Ar configuration.
    - That is, both Na⁺ and Cl⁻ have an octet of electrons.
  - NaCl forms a very regular structure in which each Na⁺ ion is surrounded by six Cl⁻ ions.
    - Similarly, each Cl⁻ ion is surrounded by six Na⁺ ions.
    - There is a regular arrangement of Na⁺ and Cl⁻.
    - Note that the ions are packed as closely as possible.
  - Ionic substances are often crystalline, brittle compounds with high melting points.

Energetics of Ionic Bond Formation

- The heat of formation of NaCl(s) is exothermic:
  \[ \text{Na}(s) + \frac{1}{2} \text{Cl}_2(g) \rightarrow \text{NaCl}(s) \quad \Delta H_f = -410.9 \text{ kJ/mol} \]
  - Separation of the NaCl into sodium and chloride ions is endothermic:
    \[ \text{NaCl}(s) \rightarrow \text{Na}^+(g) + \text{Cl}^-(g) \quad \Delta H^\circ = +788 \text{ kJ/mol} \]
    - The energy required to separate one mole of a solid ionic compound into gaseous ions is called the **lattice energy**, \( \Delta H_{\text{lattice}} \).
    - Lattice energy depends on the charge on the ions and the size of the ions.
    - The stability of the ionic compound comes from the attraction between ions of unlike charge.
    - The specific relationship is given by Coulomb’s equation:
      \[ E = k \frac{Q_1 Q_2}{d} \]
      - where \( E \) is the potential energy of the two interacting charged particles, \( Q_1 \) and \( Q_2 \) are the charges on the particles, \( d \) is the distance between their centers, and \( k \) is a constant:
        \[ k = 8.99 \times 10^9 \text{ J-m/C}^2. \]
• As \( Q_1 \) and \( Q_2 \) increase, \( E \) increases, and as \( d \) increases, \( E \) decreases.

**Electron Configuration of Ions of the s-and p-Block Elements**

• These are derived from the electron configuration of elements with the required number of electrons added or removed from the most accessible orbital.

• Electron configuration of ions can predict stable ion formation:
  - Na: \([\text{Ne}]3s^1\)
  - Na\(^+\): \([\text{Ne}]\)
  - Cl: \([\text{Ne}]3s^23p^5\)
  - Cl\(^-\): \([\text{Ne}]3s^23p^6 = [\text{Ar}]\)

**Transition-Metal Ions**

• Lattice energies compensate for the loss of up to three electrons.
• We often encounter cations with charges of 1+, 2+, or 3+ in ionic compounds.
• However, transition metals can’t attain a noble gas conformation (>3 electrons beyond a noble gas core).
  - Transition metals tend to lose the valence shell electrons first and then as many \( d \) electrons as are required to reach the desired charge on the ion.
• Thus, electrons are removed from 4\( s \) before the 3\( d \), etc.

**FORWARD REFERENCES**

• The link between lattice energy and solubility of ionic compounds will be made in Chapter 17 (section 17.4).
• The formation reaction of NaCl(s) from elements will be brought up in Chapter 19 (section 19.5) as an example of highly exothermic spontaneous processes with decreasing entropy of the system.
• Electron configurations of the first transition-series elements and chemistry of select main group and transition metals will be further discussed in Chapter 23 (sections 23.1).

**8.3 Covalent Bonding**

• The majority of chemical substances do not have characteristics of ionic compounds.
• We need a different model for bonding between atoms.
• A chemical bond formed by sharing a pair of electrons is called a **covalent** bond.
• Both atoms acquire noble-gas electronic configurations.
• This is the “glue” to bind atoms together.

**Lewis Structures**

• Formation of covalent bonds can be represented using Lewis symbols.
  - The structures are called **Lewis structures**.
  - We usually show each electron pair shared between atoms as a line and show unshared electron pairs as dots.
  - Each pair of shared electrons constitutes one chemical bond.

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12 “Ion Electron Configurations” Activity from Instructor’s Resource CD/DVD
13 “The Chemical Bond” from Further Readings
14 “Grade-12 Students’ Misconceptions of Covalent Bonding and Structure” from Further Readings
15 “H\(_2\) Bond Formation” Animation from Instructor’s Resource CD/DVD
16 “The Role of Lewis Structures in Teaching Covalent Bonding” from Further Readings
18 “Abegg, Lewis, Langmuir, and the Octet Rule” from Further Readings
19 “G.N. Lewis and the Chemical Bond” from Further Readings
20 “The Use of Dots in Chemical Formulas” from Further Readings

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• Example: \( \bullet \text{H} + \bullet \text{H} \rightarrow \text{H}^+: \text{H} \) has electrons on a line connecting the two H nuclei (H–H).

**Multiple Bonds**

• It is possible for more than one pair of electrons to be shared between two atoms (e.g., multiple bonding):
  • One shared pair of electrons is a **single bond** (e.g., H\(_2\));
  • Two shared pairs of electrons is a **double bond** (e.g., O\(_2\));
  • Three shared pairs of electrons is a **triple bond** (e.g., N\(_2\)).

• **Bond length** is the distance between the nuclei of the atoms in a bond.

  Generally, bond distances decrease as we move from single through double to triple bonds.

**FORWARD REFERENCES**

• Coordinate covalent bonds between Lewis acids and bases will be discussed in section 16.11.
• Covalent bonding between carbon atoms (and other nonmetals) will be highlighted throughout Chapter 24 on organic chemistry.

**8.4 Bond Polarity and Electronegativity**\(^{21,22,23}\)

• The electron pairs shared between two different atoms are usually unequally shared.
• **Bond polarity** describes the sharing of the electrons in a covalent bond.
  • Two extremes:
    • In a **nonpolar covalent bond**, the electrons are shared equally.
    • An example is bonding between identical atoms (example: Cl\(_2\)).
    • In a **polar covalent bond**, one of the atoms exerts a greater attraction for bonding electrons than the other (example: HCl).
    • If the difference is large enough, an ionic bond forms (example: NaCl).

**Electronegativity**\(^{24,25,26}\)

• The ability of an atom *in a molecule* to attract electrons to itself is its **electronegativity**.
• The electronegativity of an element is related to its ionization energy and electron affinity.
• Pauling electronegativity scale: from 0.7 (Cs) to 4.0 (F).
• Electronegativity increases across a period and decreases down a group.

**Electronegativity and Bond Polarity**\(^{27,28,29}\)

• Electronegativity differences close to zero result in nonpolar covalent bonds.
  • The electrons are equally or almost equally shared.
• The greater the difference in electronegativity between two atoms, the more polar the bond (polar covalent bonds).
• There is no sharp distinction between bonding types.

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\(^{21}\) “Electronegativity and Bond Type: Predicting Bond Type” from Further Readings
\(^{22}\) “Electronegativity from Avogadro to Pauling Part II: Late Nineteenth- and Early Twentieth-Century Developments” from Further Readings
\(^{23}\) “Bending a Stream of Water” from Live Demonstrations
\(^{24}\) “Periodic Trends: Electronegativity” Animation from Instructor’s Resource CD/DVD
\(^{25}\) Figure 8.7 from Transparency Pack
\(^{26}\) “Electronegativity from Avogadro to Pauling Part I: Origins of the Electronegativity Concept” from Further Readings
\(^{27}\) Figure 8.8 from Transparency Pack
\(^{29}\) “Electron Densities, Atomic Charges, and Ionic, Covalent, and Polar Bonds” from Further Readings
Basic Concepts of Chemical Bonding

Dipole Moments

- Molecules like HF have centers of positive and negative charge that do not coincide.
- These are polar molecules.
- We indicate the polarity of molecules in two ways:
  - The positive end (or pole) in a polar bond may be represented with a “δ+” and the negative pole with a “δ−”.
- We can also place an arrow over the line representing the bond.
  - The arrow points toward the more electronegative element and shows the shift in electron density toward that atom.
- We can quantify the polarity of the molecule.
  - When charges are separated by a distance, a dipole is produced.
  - The dipole moment is the quantitative measure of the magnitude of the dipole (µ).
    \[ \mu = Qr \]
  - The magnitude of the dipole moment is given in debyes (D).

Differentiating Ionic and Covalent Bonding

- Interactions of metals and nonmetals often yield ionic compounds.
  - When ionic bonding is dominant, we expect compounds to exhibit properties associated with ionic substances (high-melting solids, strong electrolyte behavior when dissolved in water, etc.).
- Interactions of nonmetals with other nonmetals often yield compounds that are covalent.
  - When covalent bonding is dominant, we expect compounds to exist as molecules and exhibit properties associated with molecular substances (low melting and boiling points, nonelectrolyte behavior when dissolved in water, etc.).
- Assigning the labels “ionic” and “covalent” to compounds is not necessarily straightforward.
  - There is a continuum between the extremes of ionic and covalent bonding.

FORWARD REFERENCES

- Bond polarities combined with molecular shapes (geometries) will be used to determine molecular polarity in Chapter 9 and physical properties of substances (Chapters 11, 13, 24).
- Polar covalent bonds between atoms of F, O, N, and H will be implicated in hydrogen bonding in Chapter 11 (section 11.2).
- The electronegativities of atoms in acids will be linked to acid strength in Chapter 16 (section 16.10).
- The electronegativities of nonmetals will be addressed again in sections 22.1 and 22.4.

8.5 Drawing Lewis Structures

30 “Molecular Polarity” Activity from Instructor’s Resource CD/DVD
31 “Drawing Lewis Structures from Lewis Symbols: A Direct Electron Pairing Approach” from Further Readings
32 “Writing Lewis Structures I” Activity from Instructor’s Resource CD/DVD
33 “Writing Lewis Structures II” Activity from Instructor’s Resource CD/DVD
34 “Lewis Structures Are Models for Predicting Molecular Structure, Not Electronic Structure” from Further Readings
35 “Teaching a Model for Writing Lewis Structures” from Further Readings
36 “Drawing Lewis Structures without Anticipating Octets” from Further Readings
37 “The ‘6N + 2 Rule’ for Writing Lewis Octet Structures” from Further Readings
38 “Another Procedure for Writing Lewis Structures” from Further Readings
39 “Lost in Lewis Structures: An Investigation of Student Difficulties in Developing Representational Competence” from Further Readings
40 “Methane” 3-D Model from Instructor’s Resource CD/DVD

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Some simple guidelines for drawing Lewis structures:

- **Sum the valence electrons from all atoms.**
  - For an anion, add electrons equal to the negative charge.
  - For a cation, subtract electrons equal to the positive charge.
- **Write the symbols for the atoms, show which atoms are attached to which, and connect them with a single bond (a dash, representing two electrons).**
  - When a central atom has other atoms bound to it, the central atom is usually written first.
  - Example: In CO$_3^{2-}$ the central atom is carbon.
  - Place the central atom in the center of the molecule and add all other atoms around it.
  - Place one bond (two electrons) between each pair of atoms.
- **Complete the octets for all atoms connected to the central atom** (exception: hydrogen can only have two electrons).
- **Place any leftover electrons on the central atom.**
- **If there are not enough electrons to give the central atom an octet, try multiple bonds.**

**Formal Charge and Alternative Lewis Structures**

- Sometimes it is possible to draw more than one Lewis structure with the octet rule obeyed for all the atoms.
- To determine which structure is most reasonable, we use formal charge.
- The *formal charge* of an atom is the charge that an atom (in a molecule) would have if all of the atoms had the same electronegativity.
- To calculate formal charge, electrons are assigned as follows:
  - All nonbonding (unshared) electrons are assigned to the atom on which they are found.
  - Half of the bonding electrons are assigned to each atom in a bond.
  - Formal charge is the number of valence electrons in the isolated atom, minus the number of electrons assigned to the atom in the Lewis structure.
- For example: consider CN$^-$ (cyanide ion):
  - For carbon:
    - There are four valence electrons (from periodic table).
    - In the Lewis structure there are two nonbonding electrons and three electrons from the triple bond.
    - There are five electrons from the Lewis structure.
    - Formal charge: $4 - 5 = -1$.
  - For nitrogen:
    - There are five valence electrons.
    - In the Lewis structure there are two nonbonding electrons and three from the triple bond.
    - There are five electrons from the Lewis structure.
    - Formal charge $= 5 - 5 = 0$.
- Using formal charge calculations to distinguish between alternative Lewis structures:
  - the most stable structure has the smallest formal charge on each atom and

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41 “Using Formal Charges in Teaching Descriptive Inorganic Chemistry” from Further Readings
42 “Formal Charges” Animation from Instructor’s Resource CD/DVD
43 “Lewis Structures, Formal Charge, and Oxidation Numbers: A More User-Friendly Approach” from Further Readings
44 Figure 8.11 from Transparency Pack
45 “Valence, Oxidation Number, and Formal Charge: Three Related but Fundamentally Different Concepts” from Further Readings
46 “Hydrogen Chloride” 3-D Model from Instructor’s Resource CD/DVD
47 “Carbon Dioxide” 3-D Model from Instructor’s Resource CD/DVD
• the most negative formal charge on the most electronegative atoms.
• It is important to keep in mind that formal charges do NOT represent REAL charges on atoms!

**FORWARD REFERENCES**
- Lewis structures of various species (neutral and charged) will appear throughout the textbook.
- Lewis structures can be considered starting points for the application of the VSEPR model or the Valence Bond theory in Chapter 9.

### 8.6 Resonance Structures

- Some molecules are not adequately described by a single Lewis structure.
- Typically, structures with multiple bonds can have similar structures with the multiple bonds between different pairs of atoms.
- Example: Experimentally, ozone has two identical bonds, whereas the Lewis structure requires one single (longer) and one double bond (shorter).
- **Resonance structures** are attempts to represent a real structure that is a mix between several extreme possibilities.
  - Resonance structures are Lewis structures that differ only with respect to placement of the electrons.
  - The “true” arrangement is a blend or hybrid of the resonance structures.
  - Example: In ozone the extreme possibilities have one double and one single bond.
    - The resonance structure has two identical bonds of intermediate character.
  - We use a double-headed arrows (↔) to indicate resonance.
  - Common examples: O$_3$, NO$_3^-$, SO$_3$, NO$_2$, and benzene.

**Resonance in Benzene**
- Benzene belongs to an important category of organic molecules called aromatic compounds.
- Benzene (C$_6$H$_6$) is a cyclic structure.
  - It consists of six carbon atoms in a hexagon.
  - Each carbon atom is attached to two other carbon atoms and one hydrogen atom.
  - There are alternating double and single bonds between the carbon atoms.
  - Experimentally, the C−C bonds in benzene are all the same length and benzene is planar.
- To emphasize the resonance between the two Lewis structures (hexagons with alternating single and double bonds), we often represent benzene as a hexagon with a circle in it.

**FORWARD REFERENCES**
- Ozone will be discussed in Chapter 18 (section 18.1).
- Benzene rings will be brought up in Chapter 11 (section 11.7) for liquid crystals and in Chapter 12 (section 12.8) for graphite and condensation polymers.
- Structure and reactivity of aromatic compounds will be further discussed in Chapter 24.

### 8.7 Exceptions to the Octet Rule

- There are three classes of exceptions to the octet rule:
  - molecules with an odd number of electrons,

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48 “If It’s Resonance, What Is Resonating?” from Further Readings
49 “Aromatic Bagels; An Edible Resonance Analogy” from Further Readings
50 “The Concept of Resonance” from Further Readings
51 “Ozone” 3-D Model from Instructor’s Resource CD/DVD
52 “The Origin of the Circle Symbol for Aromaticity” from Further Readings
53 “Explaining Resonance—A Colorful Approach” from Further Readings
54 “A Visual Aid for Teaching the Resonance Concept” from Further Readings
55 “Benzene” 3-D Model from Instructor’s Resource CD/DVD

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molecules in which one atom has less than an octet,
• molecules in which one atom has more than an octet.

**Odd Number of Electrons**

• Most molecules have an even number of electrons and complete pairing of electrons occurs. although some molecules have an odd number of electrons.
• Examples: ClO₂, NO, and NO₂.

**Less than an Octet of Valence Electrons**

• Molecules with less than an octet are also relatively rare.
• Most often encountered in compounds of boron or beryllium.
• A typical example is BF₃.

**More than an Octet of Valence Electrons**

• This is the largest class of exceptions.
• Molecules and ions with more than an octet of electrons around the central atom are often called hypervalent.
• Atoms from the third period and beyond can accommodate more than an octet.
• Examples: PCl₅, SF₄, AsF₆⁻, and ICl₄⁻.
• Elements from the third period and beyond have unfilled d orbitals that can be used to accommodate the additional electrons.
• Size also plays a role.
• The larger the central atom, the larger the number of atoms that can surround it.
• The size of the surrounding atoms is also important.
• Expanded octets occur often when the atoms bound to the central atom are the smallest and most electronegative (e.g., F, Cl, O).

**FORWARD REFERENCES**

• Chemistry of species with an odd number of electrons (radicals) will be discussed in Chapter 18 (section 18.1).
• Electron deficiency of boron will be further discussed in section 22.11.
• Valence shell expansion will be used to explain the formation of noble-gas compounds in section 22.3.

**8.8 Strengths of Covalent Bonds**

• The energy required to break a particular covalent bond in one mole of a gaseous substance is called the bond enthalpy, \( D \).
• That is, for the Cl₂ molecule, \( D(\text{Cl–Cl}) \) is given by \( \Delta H \) for the reaction:
  \[
  \text{Cl}_2(g) \rightarrow 2\text{Cl}(g).
  \]
• When more than one bond is broken:
  \[
  \text{CH}_4(g) \rightarrow \text{C}(g) + 4\text{H}(g) \quad \Delta H = 1660 \text{ kJ}
  \]
• The bond enthalpy is a fraction of \( \Delta H \) for the atomization reaction:
  \[
  D(\text{C–H}) = \frac{1}{4} \Delta H = \frac{1}{4} (1660 \text{ kJ}) = 415 \text{ kJ}.
  \]
• Bond enthalpy is always a positive quantity.

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56 “Nitric Oxide—Some Old and New Perspectives” from Further Readings
57 “Biological Roles of Nitric Oxide” from Further Readings
58 “Phosphorus Pentachloride” 3-D Model from Instructor’s Resource CD/DVD
59 Table 8.4 from Transparency Pack
Bond Enthalpies and the Enthalpies of Reactions

- We can use bond enthalpies to calculate the enthalpy for a chemical reaction.
- We recognize that in any chemical reaction bonds need to be broken and then new bonds form.
- The enthalpy of the reaction, \( \Delta H_{\text{rxn}} \), is given by the sum of bond enthalpies for bonds broken (in reactants) less the sum of bond enthalpies for bonds formed (in products):
  \[
  \Delta H_{\text{rxn}} = \sum D(\text{bonds broken}) - \sum D(\text{bonds formed})
  \]
- We illustrate the concept with the reaction between methane, CH\(_4\), and chlorine:
  \[
  \text{CH}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{CH}_3\text{Cl}(\text{g}) + \text{HCl}(\text{g})
  \]
  - In this reaction, one C–H bond and one Cl–Cl bond are broken, while one C–Cl bond and one H–Cl bond are formed.
  - So, \( \Delta H_{\text{rxn}} = [D(\text{C–H}) + D(\text{Cl–Cl})] - [D(\text{C–Cl}) + D(\text{H–Cl})] = -104 \text{ kJ} \).
  - The overall reaction is exothermic, which means that the bonds formed are stronger than the bonds broken.
  - The above result is consistent with Hess’s law.

Bond Enthalpy and Bond Length

- The distance between the nuclei of the atoms involved in a bond is called the bond length.
- Multiple bonds are shorter than single bonds.
  - We can show that multiple bonds are stronger than single bonds.
  - As the number of bonds between atoms increases, the atoms are held closer and more tightly together.

FORWARD REFERENCES

- Bond enthalpies and catalysis will be discussed in Chapter 14 (section 14.7).
- Relative length and strengths of X-H bonds (X = F, Cl, Br, I) will be important in determining strength of binary acids (e.g., HF vs. HCl) in Chapter 16 (section 16.10).
- Bond dissociation energies in N\(_2\) vs. O\(_2\) will be contrasted and linked to their reactivities in Chapter 18 (section 18.1).

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60 Figure 8.15 from Transparency Pack
61 “Bond Enthalpy” Activity from Instructor’s Resource CD/DVD
62 “The Relative Explosive Power of Some Explosives” from Further Readings
63 “Exothermic Bond Breaking: A Persistent Misconception” from Further Readings
64 “Chloromethane” 3-D Model from Instructor’s Resource CD/DVD
65 Table 8.5 from Transparency Pack
Further Readings:


**Live Demonstrations:**