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## Lecture Presentation

### Chapter 9

### Chemical Bonding I: The Lewis Model

#### HIV-Protease

- HIV-protease is a protein synthesized by the human immunodeficiency virus (HIV).
- This particular protein is crucial to the virus's ability to multiply and cause AIDS.
- Pharmaceutical companies designed molecules that would disable HIV-protease by sticking to the molecule's active site—**protease inhibitors**.
- To design such a molecule, researchers used **bonding theories** to *simulate* the shape of potential drug molecules and how they would interact with the protease molecule.

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## Bonding Theories

- Explain how and why atoms attach together to form molecules
- Explain why some combinations of atoms are stable and others are not
  - Why is water  $\text{H}_2\text{O}$ , not  $\text{HO}$  or  $\text{H}_3\text{O}$ ?
- Can be used to predict the shapes of molecules
- Can be used to predict the chemical and physical properties of compounds

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## Lewis Model

- One of the simplest bonding theories is called **Lewis theory**.
- Lewis theory emphasizes valence electrons to explain bonding.
- Using Lewis theory, we can draw models, called **Lewis structures**.
  - Also known as electron dot structures
- Lewis structures allow us to predict many properties of molecules.
  - Molecular stability, shape, size, and polarity



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## Why Do Atoms Bond?

- Chemical bonds form because they lower the potential energy between the charged particles that compose atoms.
- A chemical bond forms when the potential energy of the bonded atoms is less than the potential energy of the separate atoms.
- To calculate this potential energy, you need to consider the following interactions:
  - Nucleus-to-nucleus repulsions
  - Electron-to-electron repulsions
  - Nucleus-to-electron attractions

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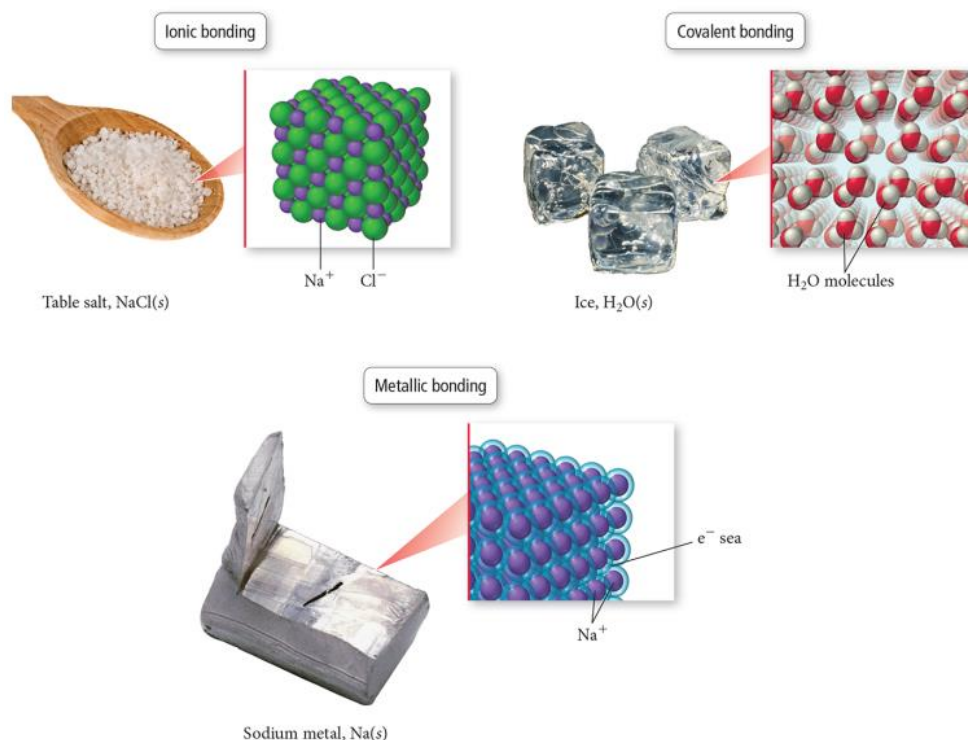
## Types of Bonds

- We can classify bonds based on the kinds of atoms that are bonded together.

| Types of Atoms        | Type of Bond | Characteristic of Bond |
|-----------------------|--------------|------------------------|
| Metal and nonmetal    | Ionic        | Electrons transferred  |
| Nonmetal and nonmetal | Covalent     | Electrons shared       |
| Metal and metal       | Metallic     | Electrons pooled       |

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# Types of Bonding



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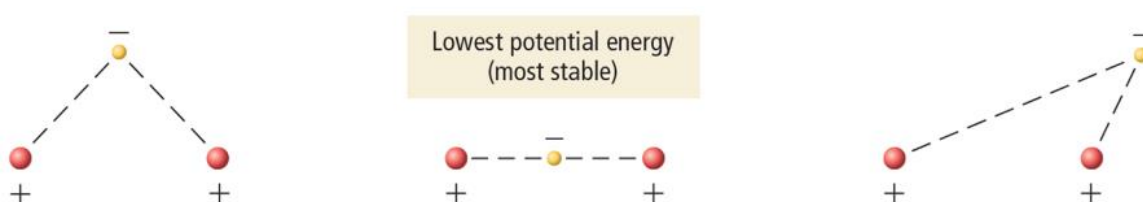
## Ionic Bonds

- When a metal atom loses electrons it becomes a **cation**.
  - Metals have low ionization energy, making it *relatively* easy to remove electrons from them.
- When a nonmetal atom gains electrons it becomes an **anion**.
  - Nonmetals have high electron affinities, making it advantageous to add electrons to these atoms.
- The oppositely charged ions are then attracted to each other, resulting in an **ionic bond**.

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## Covalent Bonds

- Nonmetal atoms have relatively high ionization energies, so it is difficult to remove electrons from them.
- When nonmetals bond together, it is better in terms of potential energy for the atoms to share valence electrons.
  - Potential energy is lowest when the electrons are between the nuclei.
- Shared electrons hold the atoms together by attracting nuclei of both atoms.



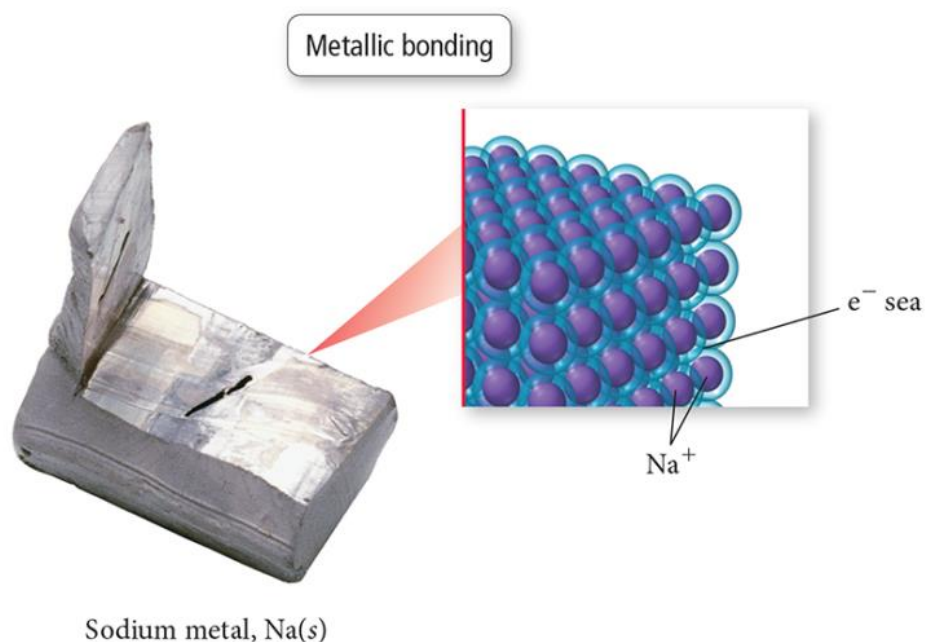
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## Metallic Bonds

- The *relatively* low ionization energy of metals allows them to lose electrons easily.
- The simplest theory of metallic bonding involves the metal atoms releasing their valence electrons to be shared as a pool by all the atoms/ions in the metal.
  - An organization of metal cation islands in a sea of electrons
  - Electrons delocalized throughout the metal structure
- Bonding results from attraction of cation for the delocalized electrons.

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# Metallic Bonding



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## Valence Electrons and Bonding

- Valence electrons are held most loosely.
- Chemical bonding involves the transfer or sharing of electrons between two or more atoms.
- Because of the two previously listed facts, valence electrons are most important in bonding.
- Lewis theory focuses on the behavior of the valence electrons.

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## Determining the Number of Valence Electrons in an Atom

- The column number on the periodic table will tell you how many valence electrons a main group atom has.
  - Transition elements all have two valence electrons. Why?

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## Lewis Structures of Atoms

- In a Lewis structure, we represent the valence electrons of main-group elements as dots surrounding the symbol for the element.
  - Also known as electron dot structures
- We use the symbol of the element to represent the nucleus and inner electrons.



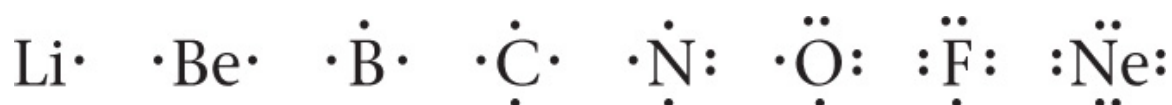
6 dots representing  
valence electrons

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## Lewis Structures of Atoms

- We use dots around the symbol to represent valence electrons.
  - Pair the first two dots for the *s* orbital electrons.
  - Put one dot on each open side for the first three *p* electrons.
  - Then, pair the rest of the dots for the remaining *p* electrons.



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## Lewis Bonding Theory

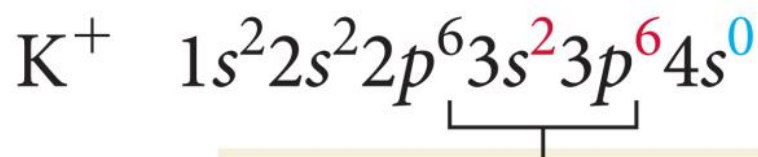
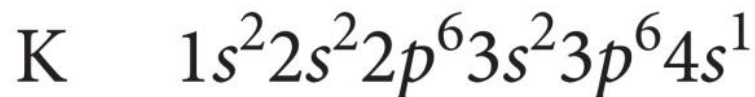
- Atoms bond because bonding results in a more stable electron configuration.
  - More stable = lower potential energy
- Atoms bond together by either transferring or sharing electrons.
- Usually, this results in all atoms obtaining an outer shell with eight electrons.
  - **Octet rule**
  - There are some exceptions to this rule: The key to remember is to try to get an electron configuration like a noble gas.

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## Stable Electron Arrangements and Ion Charge

- Metals form cations by losing valence shell electrons.
- Nonmetals form anions by gaining valence electrons.

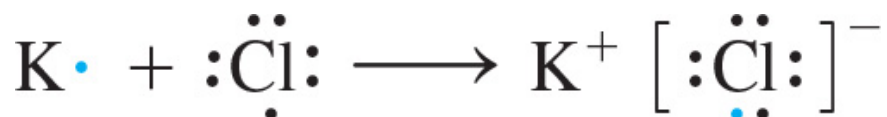


Octet in previous level

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## Lewis Theory and Ionic Bonding

- Lewis symbols can be used to represent the transfer of electrons from a metal atom to a nonmetal atom, resulting in ions that are attracted to each other and, therefore, bond.



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## Lewis Theory Predictions for Ionic Bonding

- Lewis theory predicts the number of electrons a metal atom should lose or a nonmetal atom should gain in order to attain a stable electron arrangement.
  - The octet rule
- This allows us to predict the formulas of ionic compounds that result.
- It also allows us to predict the relative strengths of the resulting ionic bonds from Coulomb's law.

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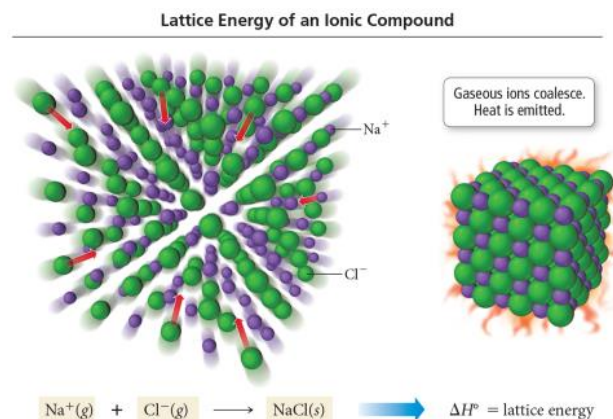
## Ionic Bonding and the Crystal Lattice

- The extra energy that is released comes from the formation of a structure, called a **crystal lattice**, in which every cation is surrounded by anions, and vice versa.
- The crystal lattice is held together by the electrostatic attraction of the cations for all the surrounding anions.
- The crystal lattice maximizes the attractions between cations and anions, leading to the most stable arrangement.

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## Crystal Lattice

- Electrostatic attraction is nondirectional!
  - No direct anion–cation pair
- Therefore, there is *no ionic molecule*.
  - The chemical formula is an empirical formula, simply giving the ratio of ions based on charge balance.



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## Lattice Energy

- The extra stability that accompanies the formation of the crystal lattice is measured as the **lattice energy**.
- The lattice energy is the energy released when the solid crystal forms from separate ions in the gas state.
  - Always exothermic
  - Hard to measure directly, but can be calculated from knowledge of other processes
- Lattice energy depends directly on the size of charges and inversely on distance between ions.

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## Determining Lattice Energy: The Born–Haber Cycle

- The **Born–Haber cycle** is a hypothetical series of reactions that represents the formation of an ionic compound from its constituent elements.
- The reactions are chosen so that the change in enthalpy of each reaction is known except for the last one, which is the lattice energy.

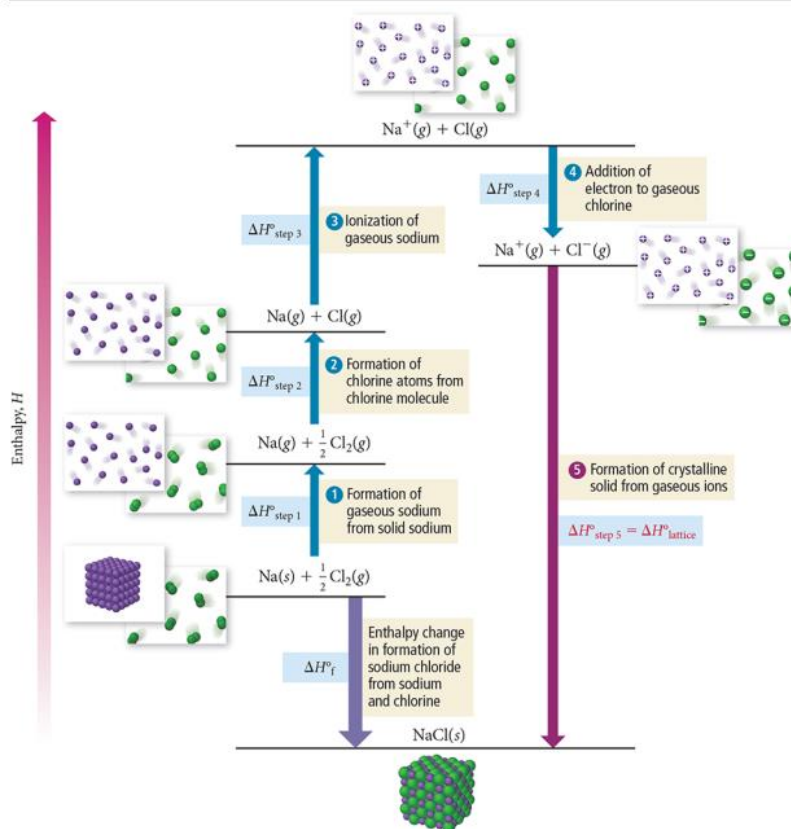
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## Born–Haber Cycle

- Use Hess's law to add up enthalpy changes of other reactions to determine the lattice energy.
- $\Delta H^\circ_f(\text{salt}) = \Delta H^\circ_f(\text{metal atoms, } g) + \Delta H^\circ_f(\text{nonmetal atoms, } g) + \Delta H^\circ_f(\text{cations, } g) + \Delta H^\circ_f(\text{anions, } g) + \Delta H^\circ(\text{crystal lattice})$
- $\Delta H^\circ(\text{crystal lattice}) = \text{lattice energy}$ 
  - For metal atom(*g*) → cation(*g*),  $\Delta H^\circ_f = \text{first ionization energy}$ 
    - Don't forget to add together all the ionization energies to get to the desired cation.
      - $M^{2+} = 1^{\text{st}} \text{ IE} + 2^{\text{nd}} \text{ IE}$
  - For nonmetal atoms (*g*) → anions (*g*),  $\Delta H^\circ_f = \text{electron affinity}$

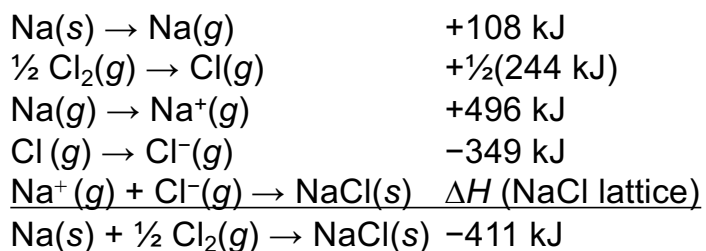
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### Born–Haber Cycle for Production of NaCl from Na(s) and Cl<sub>2</sub>(g)



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## Born–Haber Cycle for NaCl



$$\Delta H^\circ_f(\text{NaCl, s}) = \Delta H^\circ_f(\text{Na atoms, g}) + \Delta H^\circ_f(\text{Cl—Cl bond energy}) + \Delta H^\circ_f(\text{IE}_1 \text{ of Na}) + \Delta H^\circ_f(\text{EA of Cl}) + \Delta H^\circ(\text{NaCl lattice})$$

$$\begin{aligned}
 \text{NaCl lattice energy} &= (-411 \text{ kJ}) - [(+108 \text{ kJ}) + (+122 \text{ kJ}) \\
 &\quad + (+496 \text{ kJ}) + (-349 \text{ kJ})] \\
 &= -788 \text{ kJ}
 \end{aligned}$$

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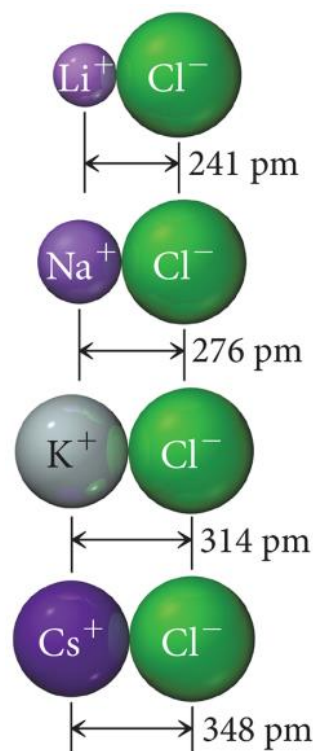
## Trends in Lattice Energy: Ion Size

- The force of attraction between charged particles is inversely proportional to the distance between them.
- Larger ions mean the center of positive charge (nucleus of the cation) is farther away from the negative charge (electrons of the anion).
  - Less exothermic lattice energy with larger ionic radius
  - More exothermic lattice energy with increasing magnitude of ionic charge

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## Lattice Energy versus Ion Size

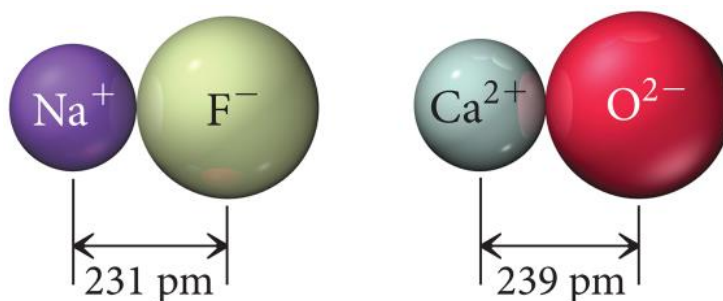
| Metal Chloride | Lattice Energy (kJ/mol) |
|----------------|-------------------------|
| LiCl           | –834                    |
| NaCl           | –788                    |
| KCl            | –701                    |
| CsCl           | –657                    |



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## Trends in Lattice Energy: Ion Charge

- The force of attraction between oppositely charged particles is directly proportional to the product of the charges.
- Larger charge means the ions are more strongly attracted.
  - **Larger charge = stronger attraction**
  - **Stronger attraction = larger lattice energy**
- Of the two factors, ion charge is generally more important.



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## Ionic Bonding Model versus Reality

- Lewis theory implies that the attractions between ions are strong.
- Lewis theory predicts ionic compounds should have high melting points and boiling points because breaking down the crystal should require a lot of energy.
  - The stronger the attraction (larger the lattice energy), the higher the melting point.
- Ionic compounds have high melting points and boiling points.
  - Melting points generally  $> 300\text{ }^{\circ}\text{C}$ .
  - All ionic compounds are solids at room temperature.

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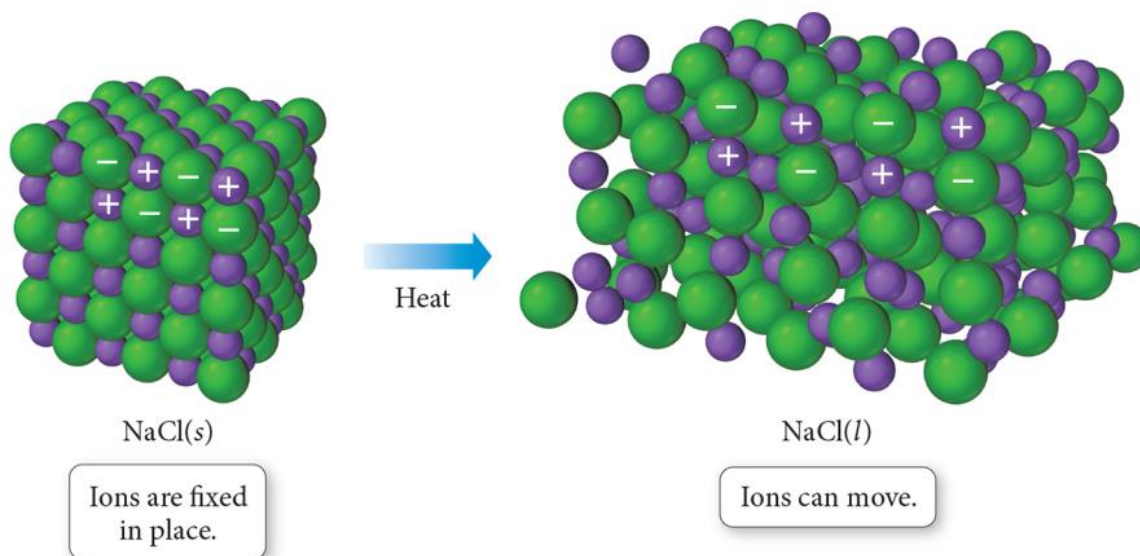


## Properties of Ionic Compounds

- Hard and brittle crystalline solids
  - All are solids at room temperature.
- Melting points generally  $> 300\text{ }^{\circ}\text{C}$
- The liquid state conducts electricity.
  - The solid state does not conduct electricity.
- Many are soluble in water.
  - The solution conducts electricity well.

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## Properties of Ionic Compounds



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## Ionic Bonding Model versus Reality

- Lewis theory implies that the positions of the ions in the crystal lattice are critical to the stability of the structure.
- Lewis theory predicts that moving ions out of position should therefore be difficult and ionic solids should be hard.
  - Hardness is measured by rubbing two materials together and seeing which “streaks” or cuts the other.
  - The harder material is the one that cuts or doesn’t streak.
- Ionic solids are relatively hard.
  - Compared to most molecular solids

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## Ionic Bonding Model versus Reality

- Lewis theory implies that if the ions are displaced from their positions in the crystal lattice, repulsive forces should occur.
- This predicts the crystal will become unstable and break apart. Lewis theory predicts ionic solids will be brittle.
- Ionic solids are brittle. When struck, they shatter.

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## Ionic Bonding Model versus Reality

- To conduct electricity, a material must have charged particles that are able to flow through the material.
- Lewis theory implies that in the ionic solid, the ions are locked in position and cannot move around.
- Lewis theory predicts that ionic solids should not conduct electricity.
- Ionic solids do not conduct electricity.

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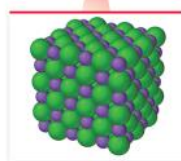
## Ionic Bonding Model versus Reality

- Lewis theory implies that in the liquid state or when dissolved in water, the ions will have the ability to move around.
- Lewis theory predicts that both a liquid ionic compound and an ionic compound dissolved in water should conduct electricity.
- Ionic compounds conduct electricity in the liquid state or when dissolved in water.

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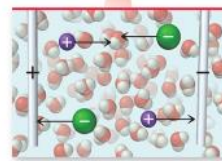
## Conductivity of NaCl

In  $\text{NaCl}(s)$ , the ions are stuck in position and not allowed to move to the charged rods.



$\text{NaCl}(s)$

In  $\text{NaCl}(aq)$ , the ions are separated and allowed to move to the charged rods.



$\text{NaCl}(aq)$

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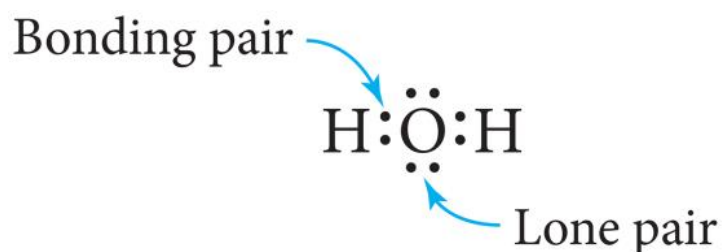
## Lewis Theory of Covalent Bonding

- Lewis theory implies that another way atoms can achieve an octet of valence electrons is to share their valence electrons with other atoms.
- The shared electrons would then count toward each atom's octet.
- The sharing of valence electrons is called **covalent bonding**.

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## Covalent Bonding: Bonding and Lone Pair Electrons

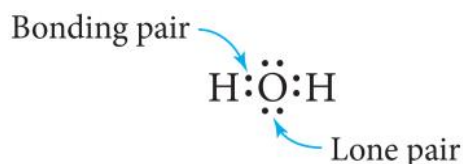
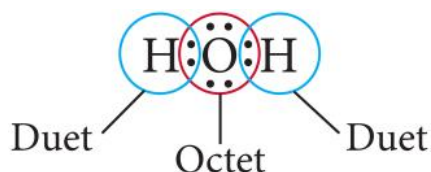
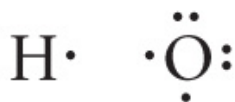
- Electrons that are shared by atoms are called **bonding pairs**.
- Electrons that are not shared by atoms but belong to a particular atom are called **lone pairs**.
  - Also known as **nonbonding pairs**



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## Single Covalent Bonds

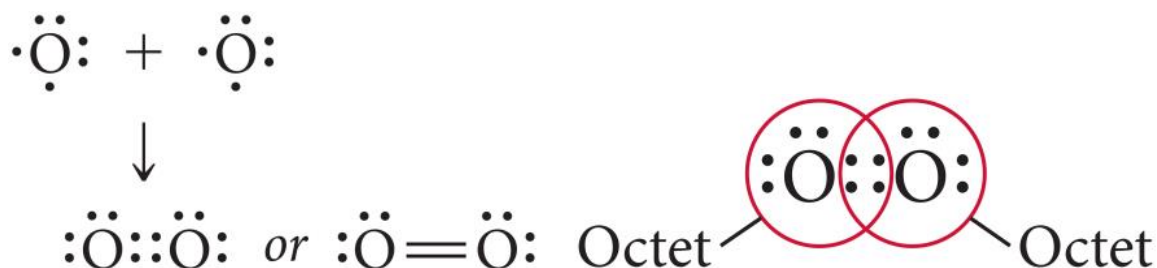
- When two atoms share one pair of electrons, it is called a **single covalent bond**.
  - Two electrons
- One atom may use more than one single bond to fulfill its octet.
  - To different atoms
  - H only duet



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## Double Covalent Bond

- When two atoms share two pairs of electrons the result is called a **double covalent bond**.
  - Four electrons



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## Triple Covalent Bond

- When two atoms share three pairs of electrons the result is called a **triple covalent bond**.
  - Six electrons



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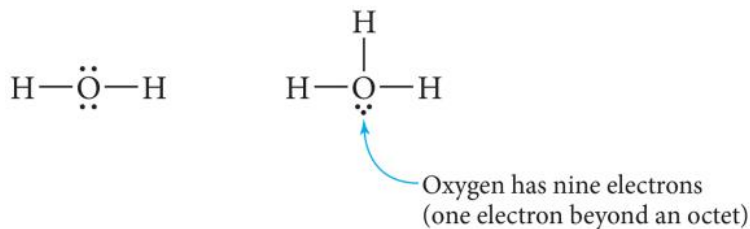
## Covalent Bonding: Model versus Reality

- Lewis theory implies that some combinations should be stable, whereas others should not.
  - Because the stable combinations result in “octets”
- Using these ideas from the Lewis theory allows us to predict the formulas of molecules of covalently bonded substances.
- Hydrogen and the halogens are all diatomic molecular elements, as predicted by Lewis theory.
- Oxygen generally forms either two single bonds or a double bond in its molecular compounds, as predicted by Lewis theory.
  - Though, as we’ll see, there are some stable compounds in which oxygen has one single bond and another where it has a triple bond, but it still has an octet.

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## Predictions of Molecular Formulas by Lewis Theory

Oxygen is more stable when it is singly bonded to two other atoms.



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## Covalent Bonding: Model versus Reality

- Lewis theory of covalent bonding implies that the attractions between atoms are **directional**.
  - The shared electrons are most stable between the bonding atoms.
- Therefore, Lewis theory predicts covalently bonded compounds will be found as individual molecules.
  - Rather than an array like ionic compounds
- Compounds of nonmetals are made of individual molecule units.

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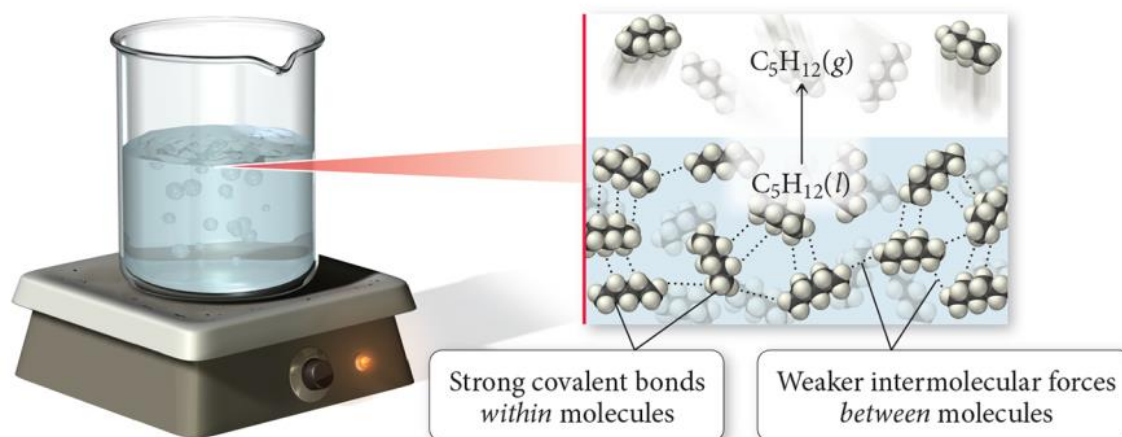
## Covalent Bonding: Model versus Reality

- Lewis theory predicts that the melting and boiling points of molecular compounds should be relatively low.
  - This involves breaking the attractions between the molecules *but not the bonds between the atoms*.
  - The covalent bonds are strong, but the attractions between the molecules are generally weak.
- Molecular compounds have low melting points and boiling points.
  - Melting points generally  $< 300\text{ }^{\circ}\text{C}$
  - Molecular compounds are found in all three states at room temperature.

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# Intermolecular Attractions versus Bonding

## Molecular Compound



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## Covalent Bonding: Model versus Reality

- Lewis theory predicts that the hardness and brittleness of molecular compounds should vary depending on the strength of intermolecular attractive forces.
  - The kind and strength of the intermolecular attractions vary based on many factors.
- Some molecular solids are brittle and hard, but many are soft and waxy.

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## Covalent Bonding: Model versus Reality

- Lewis theory predicts that neither molecular solids nor liquids should conduct electricity.
  - There are no charged particles around to allow the material to conduct.
- Molecular compounds do not conduct electricity in the solid or liquid state.
- Molecular acids conduct electricity when dissolved in water but not in the solid or liquid state, due to them being **ionized** by the water.

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## Covalent Bonding: Model versus Reality

- Lewis theory predicts that the more electrons two atoms share, the stronger the bond should be.
- Bond strength is measured by how much energy must be added into the bond to break it in half.
- In general, triple bonds are stronger than double bonds, and double bonds are stronger than single bonds.
  - However, Lewis theory would predict that double bonds are twice as strong as single bonds; the reality is that they are less than twice as strong.

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## Covalent Bonding: Model versus Reality

- Lewis theory predicts that the more electrons two atoms share, the shorter the bond should be.
  - When comparing bonds to like atoms
- Bond length is determined by measuring the distance between the nuclei of bonded atoms.
- In general, triple bonds are shorter than double bonds, and double bonds are shorter than single bonds.

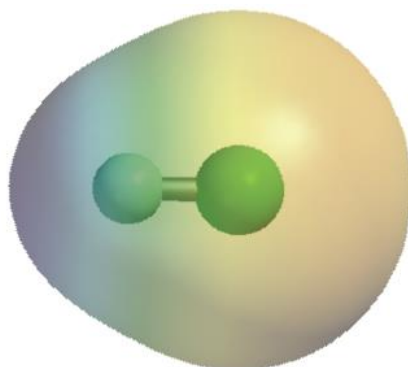
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## Polar Covalent Bonding

- Covalent bonding between unlike atoms results in unequal sharing of the electrons.
  - One atom pulls the electrons in the bond closer to its side.
  - One end of the bond has larger electron density than the other.
- The result is a **polar covalent bond**.
  - Bond polarity
  - The end with the larger electron density gets a partial negative charge.
  - The end that is electron deficient gets a partial positive charge.

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## Unequal Electron Sharing: HF



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## Bond Polarity

- Most bonds have some degree of sharing and some degree of ion formation to them.
- Bonds are classified as covalent if the amount of electron transfer is insufficient for the material to display the classic properties of ionic compounds.
- If the sharing is unequal enough to produce a dipole in the bond, the bond is classified as polar covalent.

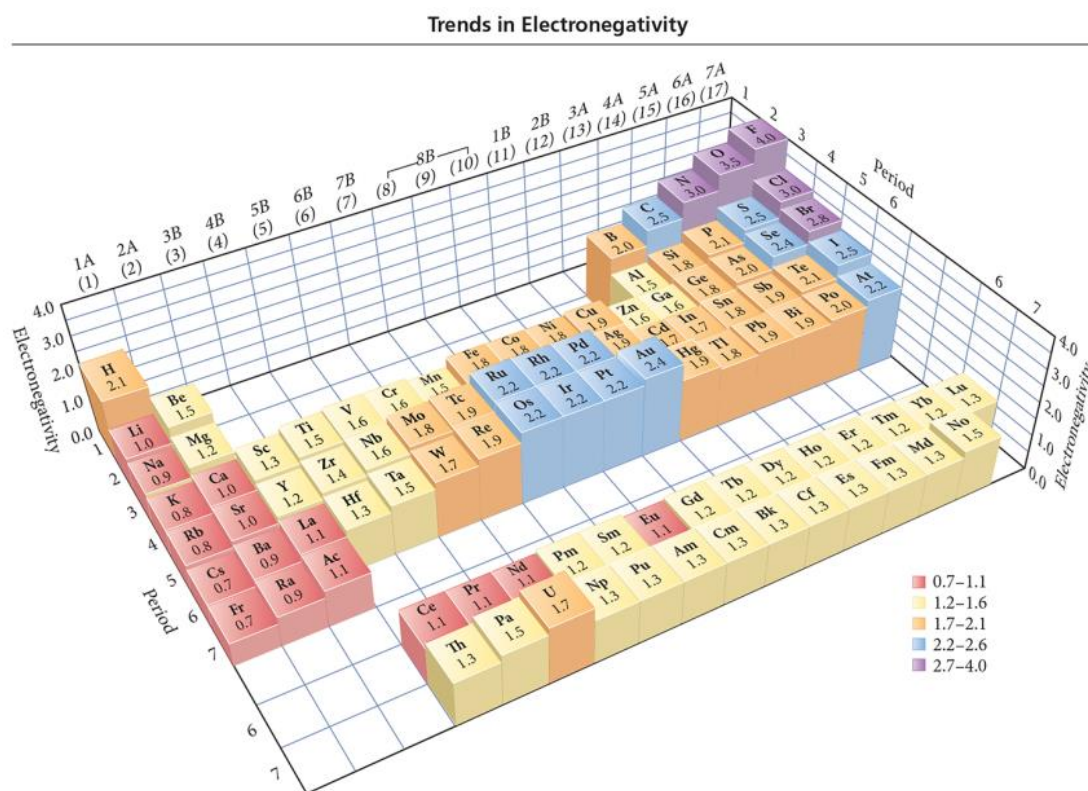
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## Electronegativity

- The ability of an atom to attract bonding electrons to itself is called **electronegativity**.
- Increases across period (left to right) and decreases down group (top to bottom)
  - Fluorine is the most electronegative element.
  - Francium is the least electronegative element.
  - Noble gas atoms are not assigned values.
  - Opposite of atomic size trend
- The larger the difference in electronegativity, the more polar the bond.
  - Negative end toward more electronegative atom

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## Electronegativity Scale



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## Electronegativity Difference and Bond Type

- If the difference in electronegativity between bonded atoms is 0, the bond is **pure covalent**.
  - Equal sharing
- If the difference in electronegativity between bonded atoms is 0.1 to 0.4, the bond is **nonpolar covalent**.
- If the difference in electronegativity between bonded atoms is 0.4 to 1.9, the bond is **polar covalent**.
- If the difference in electronegativity between bonded atoms is larger than or equal to 2.0, the bond is “100%” **ionic**.

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## Effect of Electronegativity Difference on Bond Type

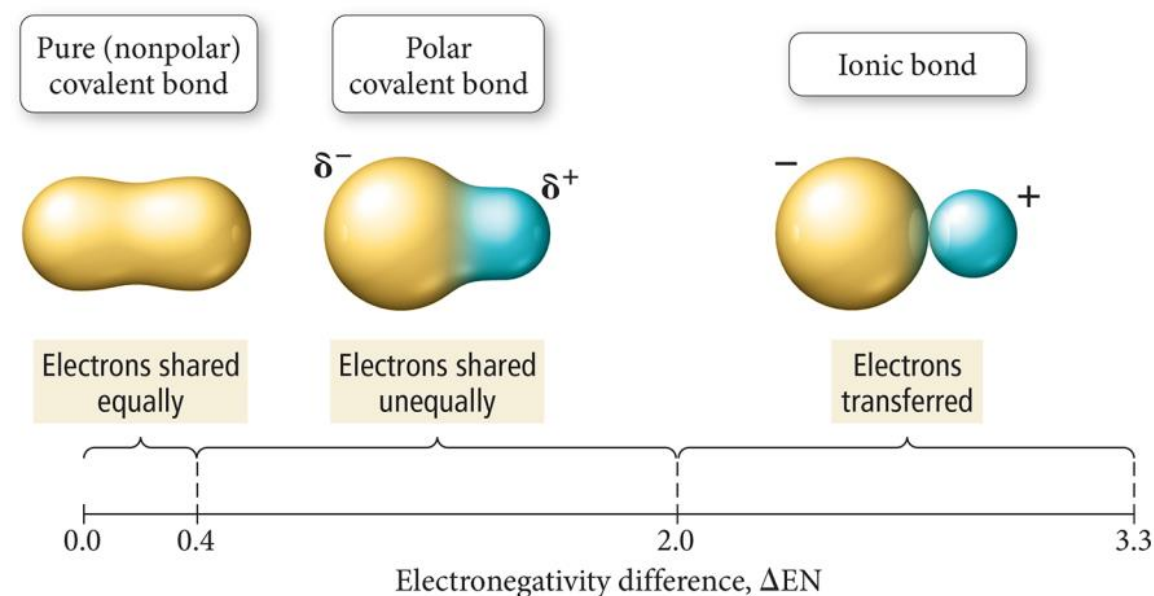
**TABLE 9.1** The Effect of Electronegativity Difference on Bond Type

| Electronegativity Difference ( $\Delta EN$ ) | Bond Type      | Example |
|--|----------------|---------|
| Small (0–0.4)                                | Covalent       | $Cl_2$  |
| Intermediate (0.4–2.0)                       | Polar covalent | HCl     |
| Large (2.0+)                                 | Ionic          | NaCl    |

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## The Continuum of Bond Types



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## Bond Dipole Moments

- **Dipole moment,  $\mu$ ,** is a measure of bond polarity.
  - A dipole is a material with a + and – end.
  - It is directly proportional to the size of the partial charges and *directly* proportional to the distance between them.
    - $\mu = (q)(r)$
    - Not Coulomb's law
    - Measured in Debyes, D
- Generally, the more electrons two atoms share and the larger the atoms are, the larger the dipole moment.

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## Dipole Moments

**TABLE 9.2** Dipole Moments of Several Molecules in the Gas Phase

| Molecule        | $\Delta EN$ | Dipole Moment (D) |
|-----------------|-------------|-------------------|
| Cl <sub>2</sub> | 0           | 0                 |
| ClF             | 1.0         | 0.88              |
| HF              | 1.9         | 1.82              |
| LiF             | 3.0         | 6.33              |

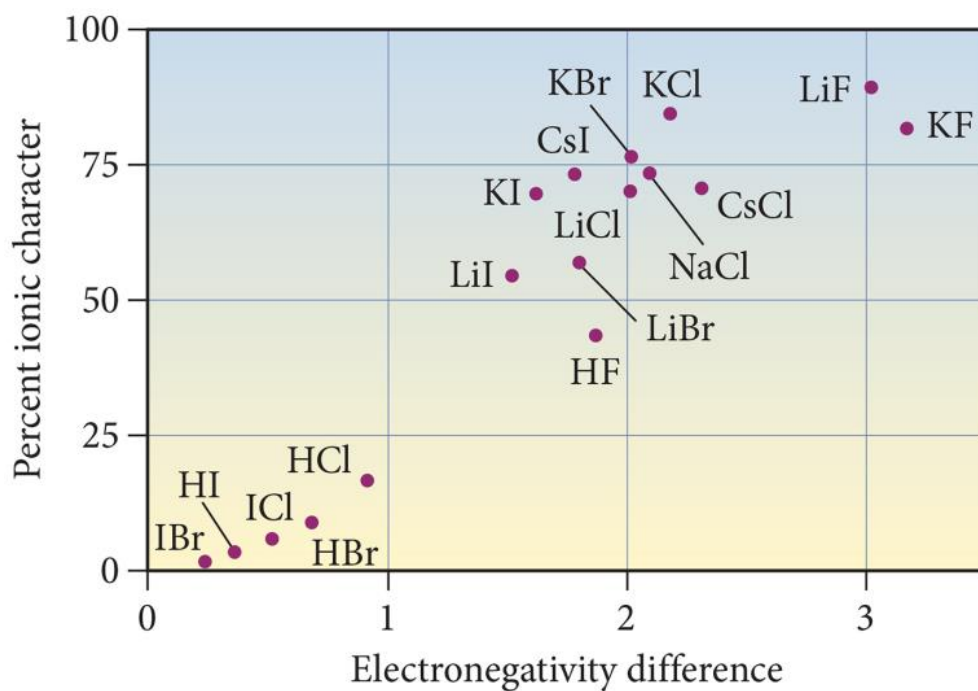
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## Percent Ionic Character

- The **percent ionic character** is the percentage of a bond's measured dipole moment compared to what it would be if the electrons were completely transferred.
- The percent ionic character indicates the degree to which the electron is transferred.

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## Percent Ionic Character



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## Writing Lewis Structures of Molecules

1. Write the correct skeletal structure for the molecule.
  - Hydrogen atoms are always terminal.
  - The more electronegative atoms are placed in terminal positions.
2. Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.
3. Distribute the electrons among the atoms, giving octets (or duets in the case of hydrogen) to as many atoms as possible.
4. If any atoms lack an octet, form double or triple bonds as necessary to give them octets.

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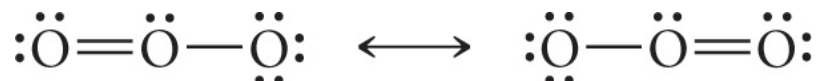
## Resonance

- Lewis theory localizes the electrons between the atoms that are bonding together.
- Extensions of Lewis theory suggest that there is some degree of delocalization of the electrons; we call this concept **resonance**.
- Delocalization of charge helps to stabilize the molecule.

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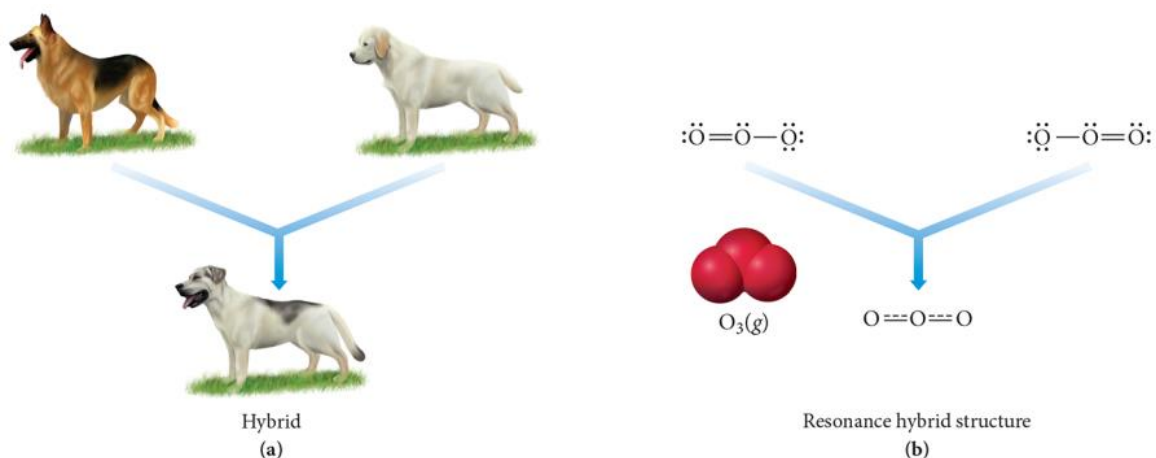
## Resonance Structures

- When there is more than one Lewis structure for a molecule that differ *only* in the position of the electrons, they are called **resonance structures**.
- The actual molecule is a combination of the resonance forms—a **resonance hybrid**.
  - The molecule does *not* resonate between the two forms, though we often draw it that way.
- Look for multiple bonds or lone pairs.



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# Resonance



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## Formal Charge

- **Formal charge** is a fictitious charge assigned to each atom in a Lewis structure that helps us to distinguish among competing Lewis structures.
- In a Lewis structure, we calculate an atom's formal charge, which indicates the charge it would have if all bonding electrons were shared equally between the bonded atom.

$$\text{FC} = \# \text{ valence } e^- - [\text{nonbonding } e^- + \frac{1}{2} \text{ bonding } e^-]$$

- Sum of all the formal charges in a molecule = 0.
  - In an ion, total equals the charge.

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# Formal Charges

|  | Structure A       |                   |                   | Structure B       |                   |                   |
|--|-------------------|-------------------|-------------------|-------------------|-------------------|-------------------|
|  | H                 | — C               | ≡ N:              | H                 | — N               | ≡ C:              |
| number of valence e <sup>-</sup>             | 1                 | 4                 | 5                 | 1                 | 5                 | 4                 |
| -number of nonbonding e <sup>-</sup>         | -0                | -0                | -2                | -0                | -0                | -2                |
| $-\frac{1}{2}(\text{number of bonding e}^-)$ | $-\frac{1}{2}(2)$ | $-\frac{1}{2}(8)$ | $-\frac{1}{2}(6)$ | $-\frac{1}{2}(2)$ | $-\frac{1}{2}(8)$ | $-\frac{1}{2}(6)$ |
| <b>Formal charge</b>                         | <b>0</b>          | <b>0</b>          | <b>0</b>          | <b>0</b>          | <b>+1</b>         | <b>-1</b>         |

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## Evaluating Resonance Structures

- Better structures have fewer formal charges.
- Better structures have smaller formal charges.
- Better structures have the negative formal charge on the more electronegative atom.

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## Exceptions to the Octet Rule

- Odd number electron species (e.g., NO)
  - Will have one unpaired electron
  - Free-radical
  - Very reactive
- Incomplete octets
  - B, Al
- Expanded octets
  - Elements with empty *d* orbitals can have more than eight electrons.

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## Bond Energies

- The amount of energy it takes to break one mole of a bond in a compound is called the **bond energy**, or bond enthalpy.
  - In the gas state
  - Homolytically—each atom gets half of the bonding electrons
  - Always positive, endothermic
- Chemical reactions involve breaking bonds in reactant molecules and making new bonds to create the products.
- The  $\Delta H^\circ_{\text{reaction}}$  can be estimated by comparing the cost of breaking old bonds to the income from making new bonds.

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## Trends in Bond Energies

- In general, the more electrons two atoms share, the stronger the covalent bond.
  - Must be comparing bonds between like atoms
  - $\text{C}\equiv\text{C}$  (837 kJ) >  $\text{C}=\text{C}$  (611 kJ) >  $\text{C}-\text{C}$  (347 kJ)
  - $\text{C}\equiv\text{N}$  (891 kJ) >  $\text{C}=\text{N}$  (615 kJ) >  $\text{C}-\text{N}$  (305 kJ)
- In general, the shorter the covalent bond, the stronger the bond.
  - Must be comparing similar types of bonds
  - $\text{Br}-\text{F}$  (237 kJ) >  $\text{Br}-\text{Cl}$  (218 kJ) >  $\text{Br}-\text{Br}$  (193 kJ)
  - Bonds get weaker down the column.
  - Bonds get stronger across the period.

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## Average Bond Energies

| Bond | Bond Energy (kJ/mol) | Bond  | Bond Energy (kJ/mol) | Bond  | Bond Energy (kJ/mol) |
|------|----------------------|-------|----------------------|-------|----------------------|
| H—H  | 436                  | N—N   | 163                  | Br—F  | 237                  |
| H—C  | 414                  | N=N   | 418                  | Br—Cl | 218                  |
| H—N  | 389                  | N≡N   | 946                  | Br—Br | 193                  |
| H—O  | 464                  | N—O   | 222                  | I—Cl  | 208                  |
| H—S  | 368                  | N=O   | 590                  | I—Br  | 175                  |
| H—F  | 565                  | N—F   | 272                  | I—I   | 151                  |
| H—Cl | 431                  | N—Cl  | 200                  | Si—H  | 323                  |
| H—Br | 364                  | N—Br  | 243                  | Si—Si | 226                  |
| H—I  | 297                  | N—I   | 159                  | Si—C  | 301                  |
| C—C  | 347                  | O—O   | 142                  | S—O   | 265                  |
| C=C  | 611                  | O=O   | 498                  | Si=O  | 368                  |
| C≡C  | 837                  | O—F   | 190                  | S=O   | 523                  |
| C—N  | 305                  | O—Cl  | 203                  | Si—Cl | 464                  |
| C=N  | 615                  | O—I   | 234                  | S=S   | 418                  |
| C≡N  | 891                  | F—F   | 159                  | S—F   | 327                  |
| C—O  | 360                  | Cl—F  | 253                  | S—Cl  | 253                  |
| C=O  | 736*                 | Cl—Cl | 243                  | S—Br  | 218                  |
| C≡O  | 1072                 |       |                      | S—S   | 266                  |
| C—Cl | 339                  |       |                      |       |                      |

\*799 in  $\text{CO}_2$ .

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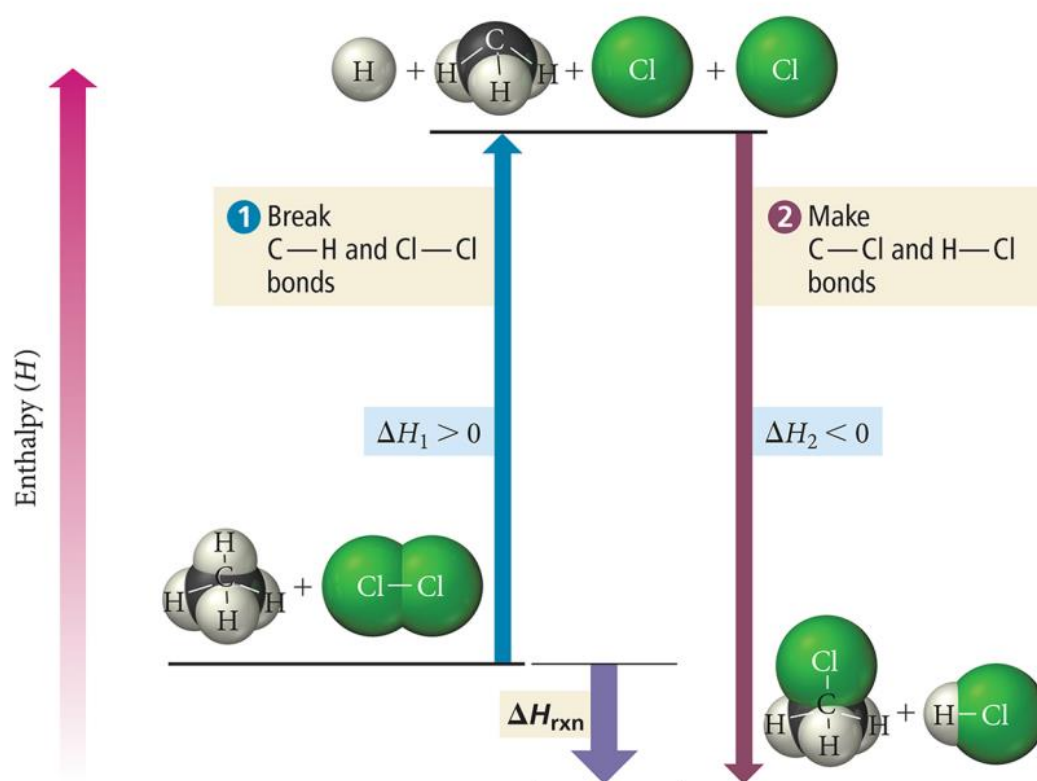
## Using Bond Energies to Estimate $\Delta H^\circ_{\text{rxn}}$

- The actual bond energy depends on the surrounding atoms and other factors.
- We often use **average bond energies** to estimate the  $\Delta H_{\text{rxn}}$ .
  - Works best when all reactants and products in gas state
- Bond breaking is endothermic,  $\Delta H(\text{breaking}) = +$ .
- Bond making is exothermic,  $\Delta H(\text{making}) = -$ .

$$\Delta H_{\text{rxn}} = \sum (\Delta H(\text{bonds broken})) + \sum (\Delta H(\text{bonds formed}))$$

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### Estimating the Enthalpy Change of a Reaction from Bond Energies

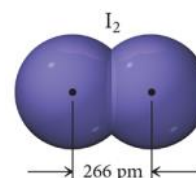
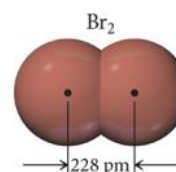
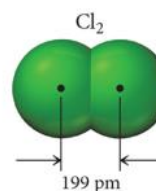
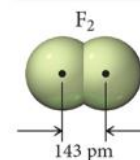


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## Bond Lengths

- The distance between the nuclei of bonded atoms is called the **bond length**.
- Because the actual bond length depends on the other atoms around the bond, we often use the **average bond length**.
  - Averaged for similar bonds from many compounds

Bond Lengths



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## Trends in Bond Lengths

- In general, the more electrons two atoms share, the shorter the covalent bond.
  - Must be comparing bonds between like atoms
  - $C \equiv C$  (120 pm) <  $C = C$  (134 pm) <  $C - C$  (154 pm)
  - $C \equiv N$  (116 pm) <  $C = N$  (128 pm) <  $C - N$  (147 pm)
- Generally, bond length decreases from left to right across period.
  - $C - C$  (154 pm) >  $C - N$  (147 pm) >  $C - O$  (143 pm)
- Generally, bond length increases down the column.
  - $F - F$  (144 pm) >  $Cl - Cl$  (198 pm) >  $Br - Br$  (228 pm)
- In general, as bonds get longer, they also get weaker.

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# Bond Lengths

**TABLE 9.4** Average Bond Lengths

| Bond | Bond Length (pm) | Bond | Bond Length (pm) | Bond  | Bond Length (pm) |
|------|------------------|------|------------------|-------|------------------|
| H—H  | 74               | C—C  | 154              | N—N   | 145              |
| H—C  | 110              | C=C  | 134              | N=N   | 123              |
| H—N  | 100              | C≡C  | 120              | N≡N   | 110              |
| H—O  | 97               | C—N  | 147              | N—O   | 136              |
| H—S  | 132              | C=N  | 128              | N=O   | 120              |
| H—F  | 92               | C≡N  | 116              | O—O   | 145              |
| H—Cl | 127              | C—O  | 143              | O=O   | 121              |
| H—Br | 141              | C=O  | 120              | F—F   | 143              |
| H—I  | 161              | C—Cl | 178              | Cl—Cl | 199              |
|      |                  |      |                  | Br—Br | 228              |
|      |                  |      |                  | I—I   | 266              |

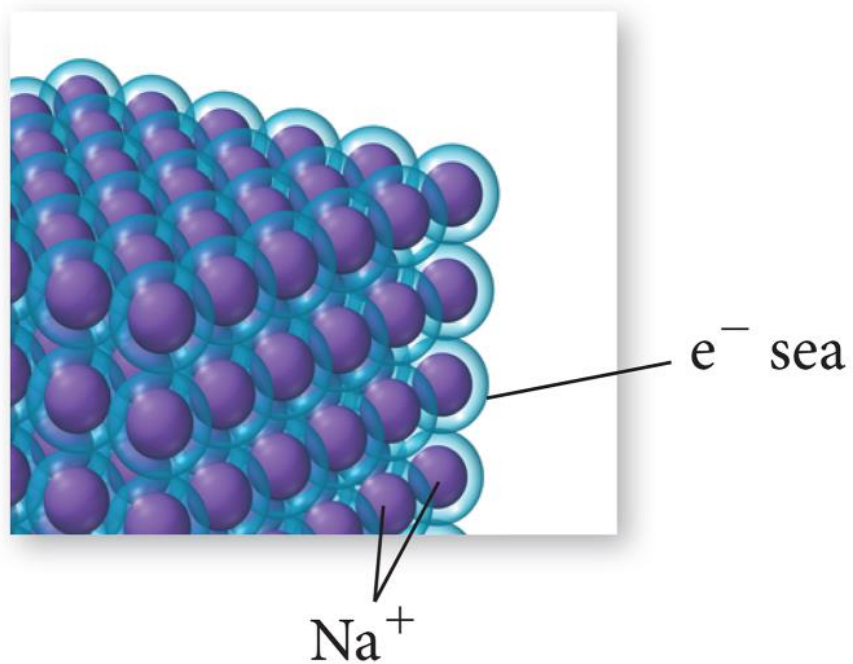
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## Metallic Bonds

- The low ionization energy of metals allows them to lose electrons easily.
- The simplest theory of metallic bonding involves the metal atoms releasing their valence electrons to be shared by all atoms/ions in the metal.
  - An organization of metal cation islands in a sea of electrons
  - Electrons delocalized throughout the metal structure
- Bonding results from attraction of the cations for the delocalized electrons.

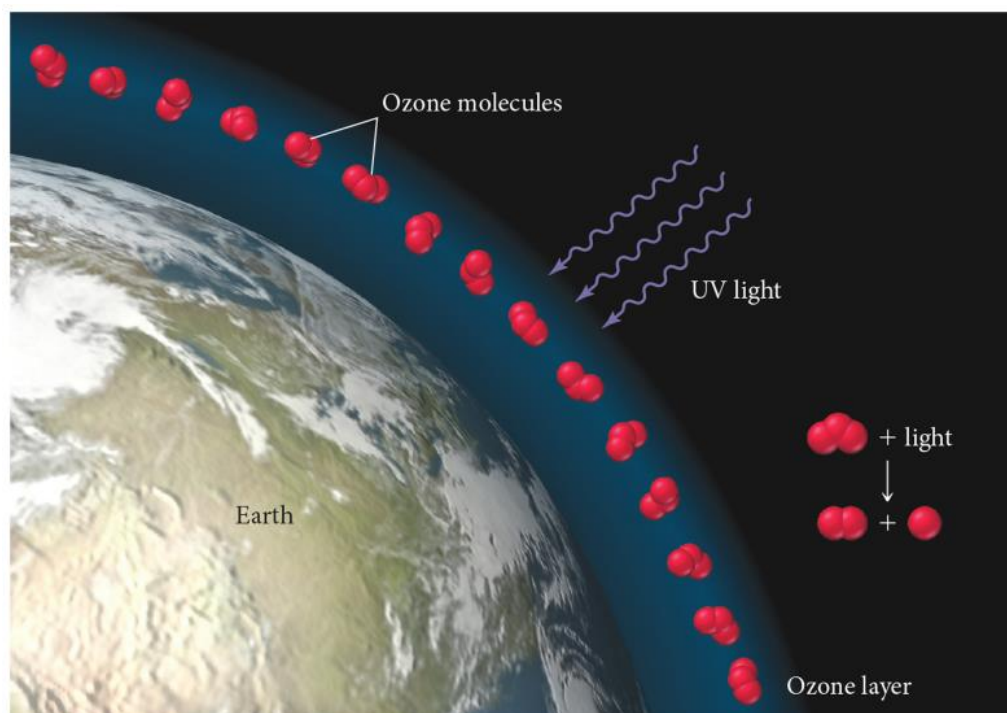
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# Metallic Bonding



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# Ozone Layer



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