## 2.1 Types of Chemical Bonds

Essential knowledge statements from the AP Chemistry CED:

- Electronegativity values for the representative elements increase going from left to right across a period and decrease going down a group. These trends can be understood qualitatively through the electronic structure of the atoms, the shell model, and Coulomb's law.
- Valence electrons shared between atoms of similar electronegativity constitute a nonpolar covalent bond. For example, bonds between carbon and hydrogen are effectively nonpolar even though carbon is slightly more electronegative than hydrogen.
- Valence electrons shared between atoms of unequal electronegativity constitute a polar covalent bond.
  - The atom with a higher electronegativity will develop a partial negative charge relative to the other atom in the bond.
  - In single bonds, greater differences in electronegativity lead to greater bond dipoles.
  - All polar bonds have some ionic character, and the difference between ionic and covalent bonding is not distinct but rather a continuum.
- The difference in electronegativity is not the only factor in determining if a bond should be designated as ionic or covalent. Generally, bonds between a metal and nonmetal are ionic, and bonds between two nonmetals are covalent. Examination of the properties of a compound is the best way to characterize the type of bonding.
- In a metallic solid, the valence electrons from the metal atoms are considered to be delocalized and not associated with any individual atom.

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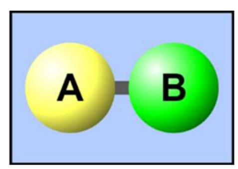
Visit the following website: bit.ly/PhET_polarity

Or you can search online for "PhET polarity"

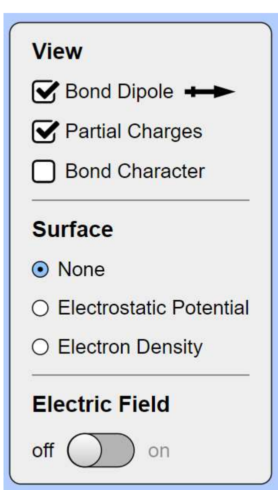
Once the website loads, you should have the following three options.

"Two Atoms" "Three Atoms" "Real Molecules"

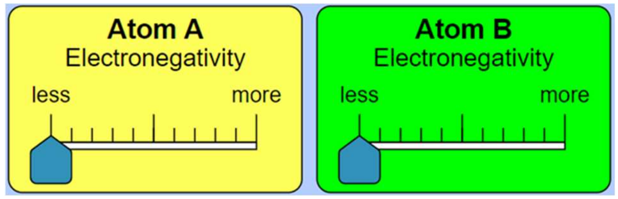
Choose the "Two Atoms" icon located on the left side of the screen.




On the right side of the screen, select the following settings.



At the bottom of the screen, adjust the slider for the Electronegativity of Atom A and Atom B so that each atom looks like the following.




Now slowly move the slider for Atom B from left to right, so that the Electronegativity of Atom B changes from "less" to "more. Observe the changes that occur in the bond between A and B.

- The symbol $\delta+$ represents a partial positive charge.
- The symbol $\delta-$ represents a partial negative charge.
- The atom with the lower electronegativity value has a partial positive charge ($\delta+$).
- The atom with the higher electronegativity value has a partial negative charge ($\delta-$).
- We use the term **dipole** or **dipole moment** when two opposite charges are separated by a small distance. A dipole is created in a polar covalent bond. The arrow in the dipole points toward the atom with the higher electronegativity value.
- In the PhET simulation, you can observe the following.
  - When the atoms have the same electronegativity value, a nonpolar bond is formed. The partial charges and the dipole arrow disappear.
  - When the atoms have different electronegativity values, a polar bond is formed. The length of the dipole arrow increases as the difference in electronegativity increases. The longer the arrow is, the greater the dipole moment is and the more polar the bond is.

1. Each of the following bonds is polar. Based on the periodic trends in electronegativity, label each of the following bonds as follows.
   - Place the symbols $\delta+$ and $\delta-$ on each atom based on their relative electronegativity values.
   - Draw an arrow to represent the direction of the dipole in each polar bond.

C—N F—Br

Cl—P S—O

2. Which of the following bonds is classified as nonpolar (circle one)? Si–P F–F Br–Cl

| Electronegativity Values | | | | | | |
|---|---|---|---|---|---|---|
| H 2.20 | | | | | | |
| Li 0.98 | Be 1.57 | B 2.04 | C 2.55 | N 3.04 | O 3.50 | F 3.98 |
| Na 0.93 | Mg 1.31 | Al 1.61 | Si 1.90 | P 2.19 | S 2.58 | Cl 3.16 |
| K 0.82 | Ca 1.00 | Ga 1.81 | Ge 2.01 | As 2.18 | Se 2.55 | Br 2.96 |
| Rb 0.82 | Sr 0.95 | In 1.78 | Sn 1.96 | Sb 2.05 | Te 2.10 | I 2.66 |

3. Use the information in the data table above to answer the following questions.

(a) Arrange each of the following bonds in order of increasing polarity: C–O, N–O, and B–F

| least polar bond | ----------> | most polar bond |
|---|---|---|
| | | |

(b) Arrange each of the following bonds in order of increasing polarity: C–P, P–F, and C–Cl

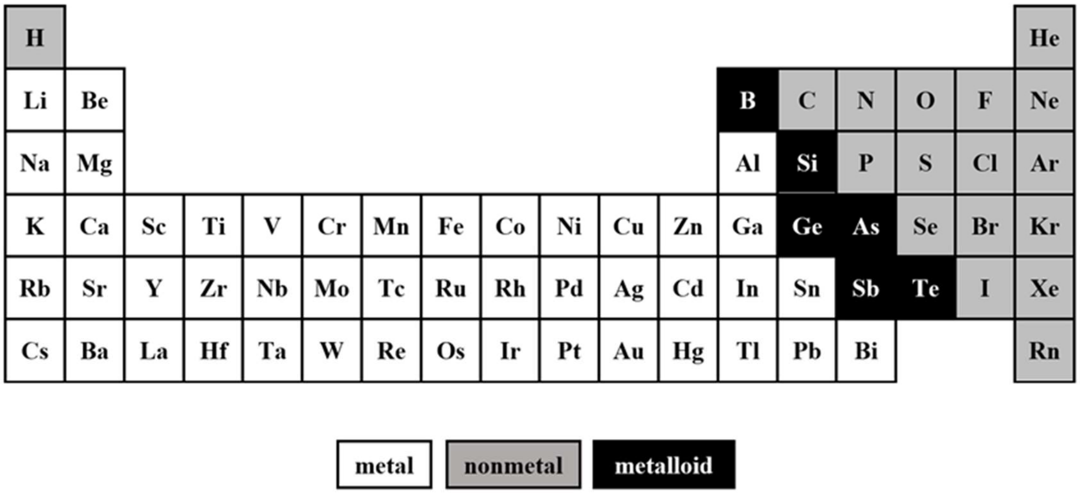
| least polar bond | ----------> | most polar bond |
|---|---|---|
| | | |

4. (a) Which of the following bonds, H–X or H–Y, is the more polar bond? Justify your choice based on the information in the table below.

| Bond | Dipole Moment (D) |
|---|---|
| H–X | 1.82 |
| H–Y | 0.44 |

(b) The table below gives information about the dipole moment for the H–Br bond. Do you predict that the dipole moment for the H–Cl bond should be less than 0.82 D or greater than 0.82 D? Justify your answer based on periodic trends in electronegativity.

| Bond | Dipole Moment (D) |
|---|---|
| H–Cl | ? |
| H–Br | 0.82 |

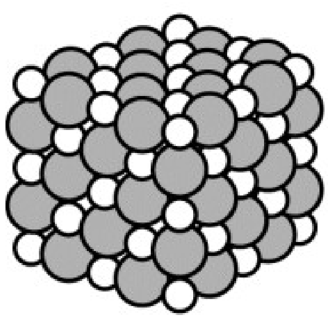



There is an important difference between a covalent bond and an ionic bond. The usual method used to classify a bond as either ionic or covalent is the following.

metal + nonmetal = ionic bond two nonmetals = covalent bond

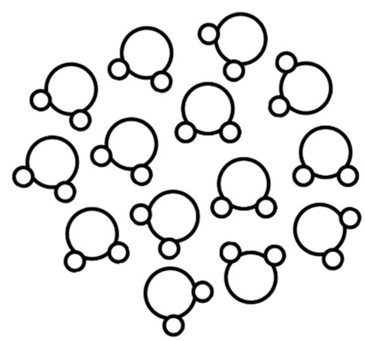
This is a simple criterion that works well for many examples. However, the nature of a chemical bond is more complicated than this. In reality, **all polar bonds have some ionic character. There is a continuum between the extremes of covalent bonding and ionic bonding.**

If we consider a particle-level view, we can think of the difference between an ionic substance and a covalent (molecular) substance as follows.



An ionic substance that consists of positive and negative ions, with attractive forces between the oppositely charged ions

- it is a solid at room temperature
- it is a hard, brittle solid with a relatively high melting point
- it does not conduct electricity as a solid
- it does conduct electricity when molten or aqueous



A covalent (molecular) substance that consists of individual molecules, with attractive forces between the molecules

- it can exist as a solid, a liquid, or a gas at room temperature
- if it is a solid, it is a soft solid with a relatively low melting point
- it does not conduct electricity, whether solid, molten, or aqueous

5. Here is a situation in which examination of the properties of substance is the best way to characterize the type of bonding present in that substance.

(a) Tin (Sn) is classified as a (circle one) metal nonmetal

(b) Chlorine (Cl) is classified as a (circle one) metal nonmetal

(c) What type of bonding, ionic or covalent, would you **predict** for a compound that contains

tin (Sn) and chlorine (Cl)? ______________________________________________

(d) Consider the properties of the substances $SnCl_2$ and $SnCl_4$ listed in the table below. This information should help you to predict the type of bonding that is present in each substance.

| | $SnCl_2$ | $SnCl_4$ |
|---|---|---|
| Appearance | white crystalline solid | colorless liquid |
| Melting Point | 247°C | –34°C |
| Boiling Point | 623°C | 114°C |
| Type of Bonding | | |

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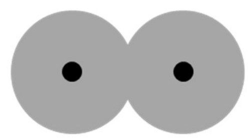
**2.2 Intramolecular Force and Potential Energy**

Essential knowledge statements from the AP Chemistry CED:

- A graph of potential energy versus the distance between atoms is a useful representation for describing the interactions between atoms. Such graphs illustrate both the equilibrium bond length (the separation between atoms at which the potential energy is lowest) and the bond energy (the energy required to separate the atoms).
- In a covalent bond, the bond length is influenced by both the size of the atom's core and the bond order (i.e., single, double, triple). Bonds with a higher order are shorter and have larger bond energies.
- Coulomb's law can be used to understand the strength of interactions between cations and anions.
  - Because the interaction strength is proportional to the charge on each ion, larger charges lead to stronger interactions.
  - Because the interaction strength increases as the distance between the centers of the ions (nuclei) decreases, smaller ions lead to stronger interactions.

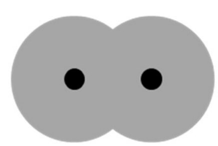
A chemical bond involves a balance between attractions and repulsions.
Consider the following arrangements of two hydrogen atoms.

This is not a covalent bond.
The atoms are too far apart, and they do not interact with each other.
The potential energy between the atoms is equal to zero.

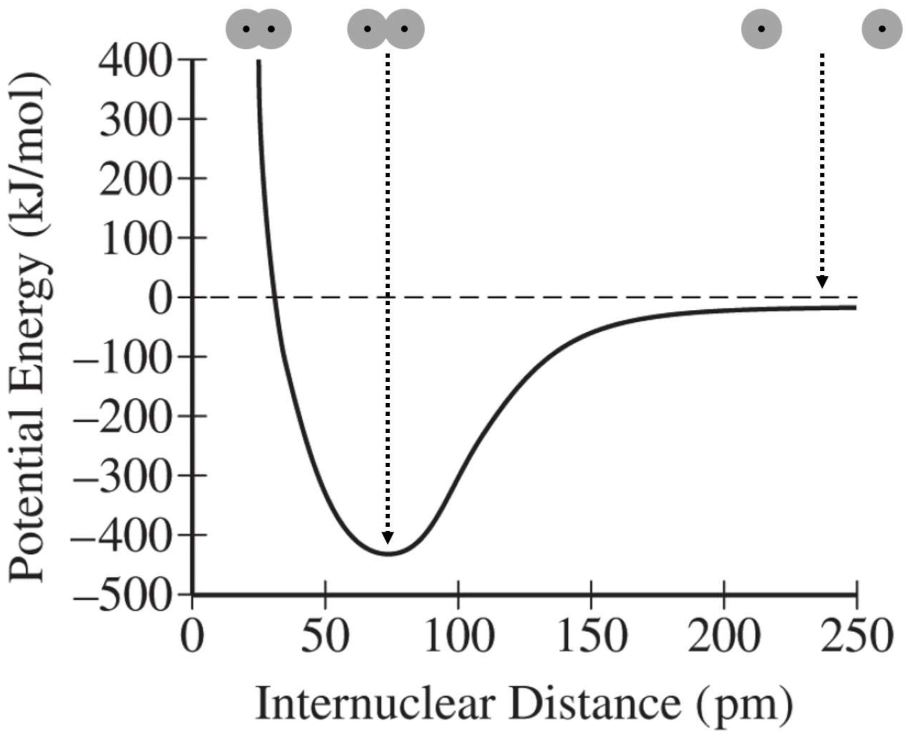


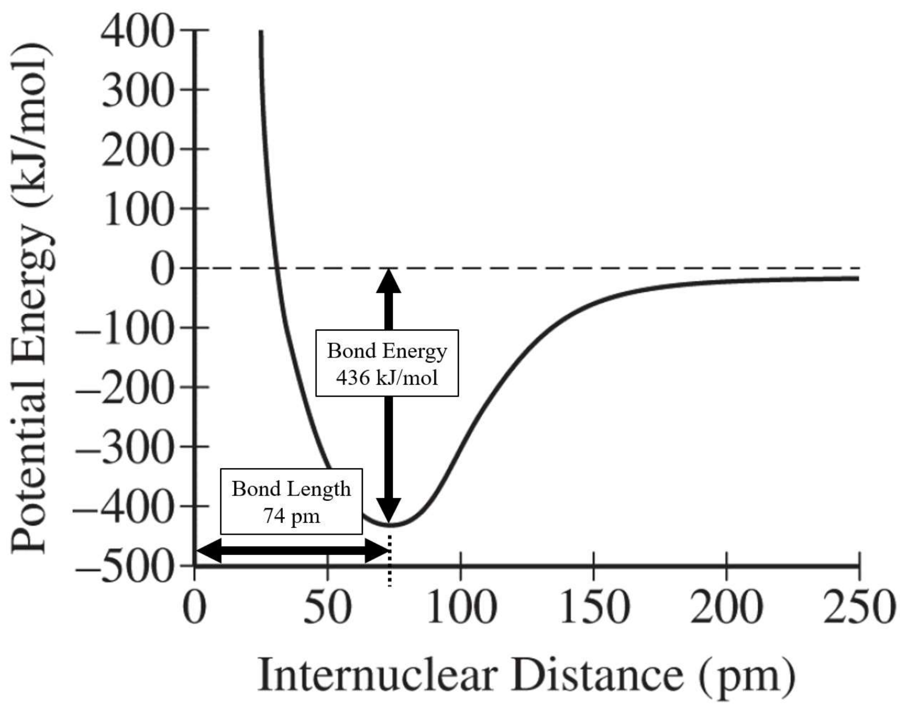
This is a stable covalent bond between two hydrogen atoms. The internuclear distance (bond length) is equal to 74 pm.
The potential energy between the atoms is equal to –436 kJ/mol.
There is a balance between the following:

- **attractions** between the proton of one atom and the electron of the other atom
- **repulsions** between protons
- **repulsions** between electrons

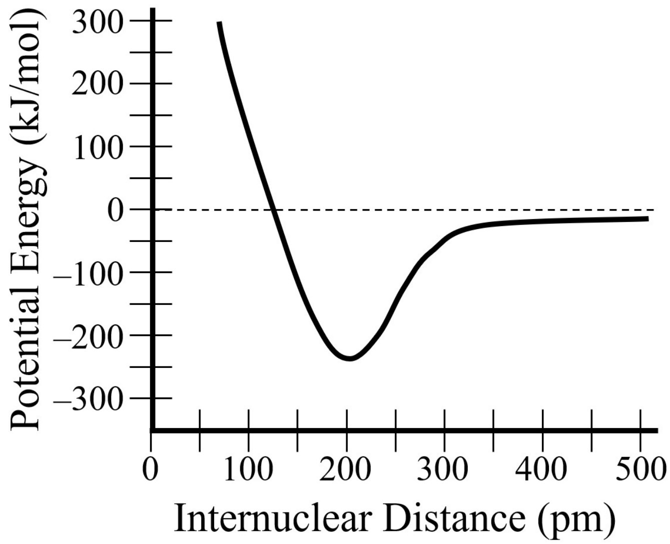


This is an unstable arrangement.
The atoms are too close to each other.
There is too much repulsion between the protons of each atom.
The potential energy between the atoms is very high.




- The bond length of the H–H bond is equal to 74 pm.
- The bond energy of the H–H bond is equal to 436 kJ/mol.
- 436 kJ/mol represents the following.
  - The energy that is released when the H–H bond is formed.
  - The energy that is absorbed when the H–H bond is broken

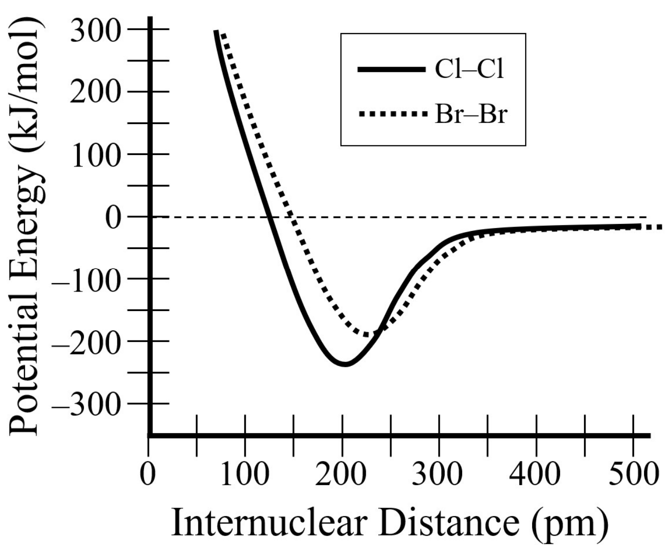



6. The diagram above represents a graph of potential energy versus internuclear distance for two atoms of chlorine (Cl). Use the information in the diagram to answer the following questions.

   (a) The length of the Cl–Cl bond is approximately _____________ pm.

   (b) The bond energy of the Cl–Cl bond is approximately ________________ kJ/mol.

7. Based on the periodic trend in atomic radius, the Br atom has a ___________________________ atomic radius than the Cl atom. The length of the Br–Br bond should be _______________________ than the length of the Cl–Cl bond. A shorter bond is harder to break, and a longer bond is easier to break. Therefore the Br–Br bond should have a _________________________ bond energy than the Cl–Cl bond. You can see a comparison of the bond length and the bond energy for Cl–Cl and Br–Br in the diagram below.



| Bond | Bond Length (pm) | Bond Energy (kJ/mol) |
|---|---|---|
| Br–Br | 228 | 193 |

8. Information for the bond length and the bond energy of the Br–Br bond is shown in the table above.

   The bond length of the I–I bond should be ________________________ than 228 pm.

   The bond energy of the I–I bond should be ________________________ than 193 kJ/mol.

Bond length and bond energy are also affected by the number of bonds that are formed between two atoms. The **bond order** refers to the number of electron pairs that are shared between two atoms.

| Example | Number of Shared Electron Pairs | Bond Order | Bond Length (pm) | Bond Energy (kJ/mol) |
|---|---|---|---|---|
| C–C | 1 | 1 | 154 | 348 |
| C=C | 2 | 2 | 134 | 614 |
| C≡C | 3 | 3 | 120 | 839 |

9. When comparing bonds between the same atom, we can make the following general statements.

   A double bond is ( shorter longer ) and ( weaker stronger ) than a single bond.

   A triple bond is ( shorter longer ) and ( weaker stronger ) than a double bond.

   As the bond order increases, the bond length __________________________ and

   the bond energy ___________________________.

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Ionic compounds have relatively high melting points because the attractions between the ions in the solid crystal are very strong. Another measure of the strength of the attractions between ions is known as the **lattice energy**. Lattice energy can be defined as the energy required to separate one mole of a solid ionic compound into its gaseous ions. For example, the lattice energy of NaCl is equal to 788 kJ/mol.

$$788\text{ kJ} + \text{NaCl}(s) \rightarrow \text{Na}^{+}(g) + \text{Cl}^{-}(g)$$

Lattice energy can also be defined as the energy released when gaseous ions combine to form one mole of a solid ionic compound.

$$\text{Na}^{+}(g) + \text{Cl}^{-}(g) \rightarrow \text{NaCl}(s) + 788\text{ kJ}$$

The greater the magnitude of the lattice energy, the stronger the attraction between the ions in the crystal lattice of an ionic compound.

When you were studying Topic 1.7 (Periodic Trends), you learned about Coulomb's law. This law describes the attractive (or repulsive) force between two charged particles.

$$F_{coulombic} \propto \frac{q_1 q_2}{r^2}$$

The magnitude of the lattice energy of an ionic solid depends on two factors.

- the magnitude of the charges on the ions
- the distance between the ions in the crystal lattice

10. The magnitude of the lattice energy of an ionic compound tends to **increase**

    as the magnitude of the charges on the ions ______________________________, and as

    the distance between the ions in the crystal lattice ____________________________.

| Substance | Lattice Energy (kJ/mol) |
|---|---|
| sodium fluoride, NaF | 930 |

11. The lattice energy of sodium fluoride, NaF, is given in the table above.

    (a) Do you predict that the lattice energy of sodium chloride (NaCl) should be less than or greater than 930 kJ/mol? Justify your choice in terms of periodic trends and Coulomb's law.

    (b) Do you predict that the lattice energy of magnesium oxide (MgO) should be less than or greater than 930 kJ/mol? Justify your choice in terms of periodic trends and Coulomb's law.

**2.3 Structure of Ionic Solids**

Essential knowledge statement from the AP Chemistry CED:

- The cations and anions in an ionic crystal are arranged in a systematic, periodic 3-D array that maximizes the attractive forces among cations and anions while minimizing the repulsive forces.

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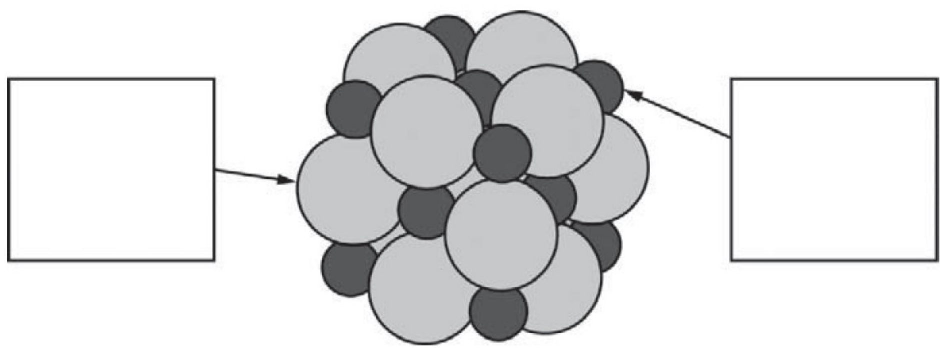
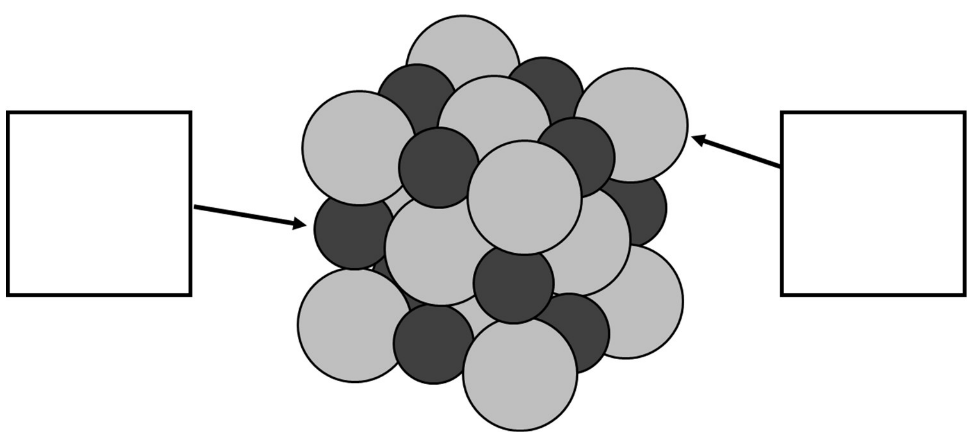
12. The diagram below is a particulate representation of a portion of a crystal of lithium chloride (LiCl). Identify the ions in the representation by writing the symbol and charge of each ion in the boxes provided.



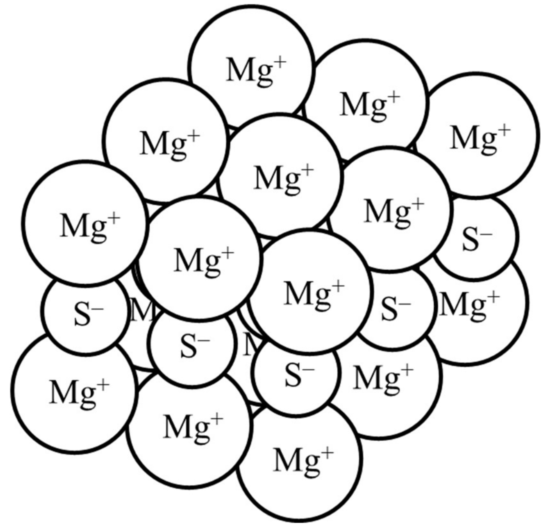
13. The diagram below is a particulate representation of a portion of a crystal of cesium fluoride (CsF). Identify the ions in the representation by writing the symbol and charge of each ion in the boxes provided.



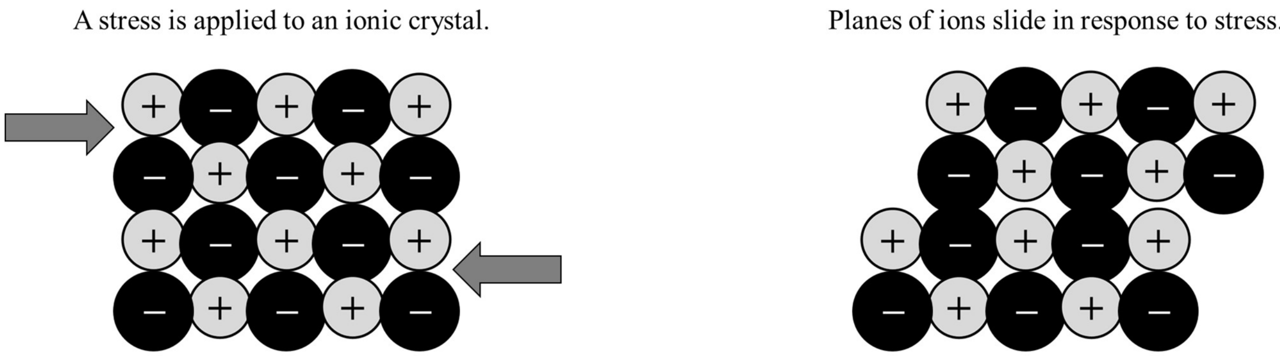
14. Sodium chloride (NaCl) contains positive and negative ions. Explain why solid NaCl does not conduct electricity.

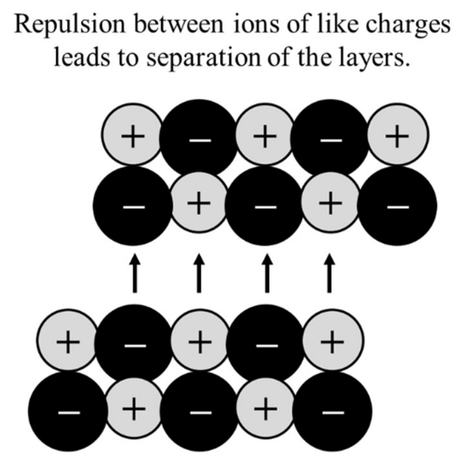
15. What are two changes that can be done to an ionic solid such as NaCl that will enable it to conduct electricity?

16. A particle diagram for solid magnesium sulfide (MgS) is shown below. This particle diagram is incorrect, and contains several different errors. Identify all of the errors in the diagram.




Ionic solids tend to be brittle, and they can be cleaved along well-defined planes. The following diagrams can be used to explain the brittleness of ionic solids.



## 2.4 Structure of Metals and Alloys

Essential knowledge statements from the AP Chemistry CED:

- Metallic bonding can be represented as an array of positive metal ions surrounded by delocalized valence electrons (i.e., a "sea of electrons").
- Interstitial alloys form between atoms of different radii, where the smaller atoms fill the interstitial spaces between the larger atoms (e.g., with steel in which carbon occupies the interstices in iron).
- Substitutional alloys form between atoms of comparable radius, where one atom substitutes for the other in the lattice. (In certain brass alloys, other elements, usually zinc, substitute for copper.)

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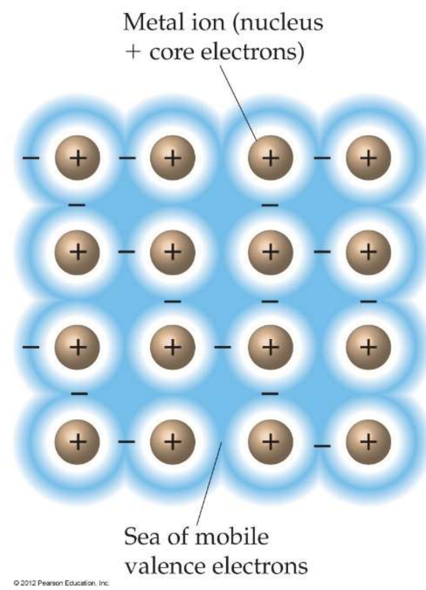
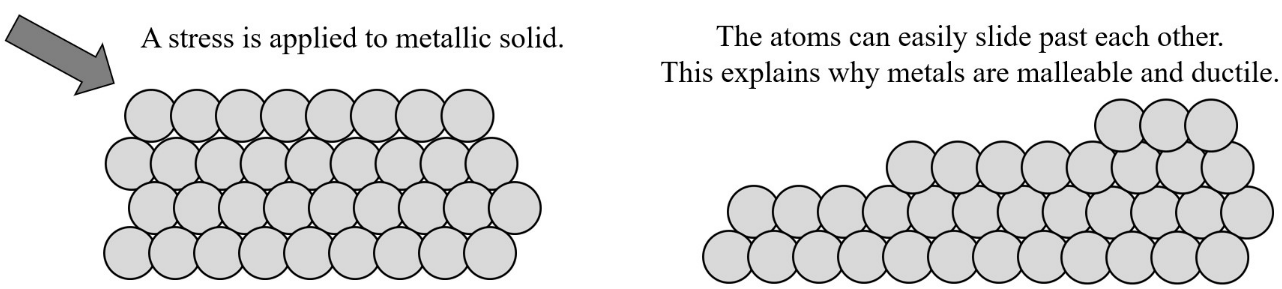
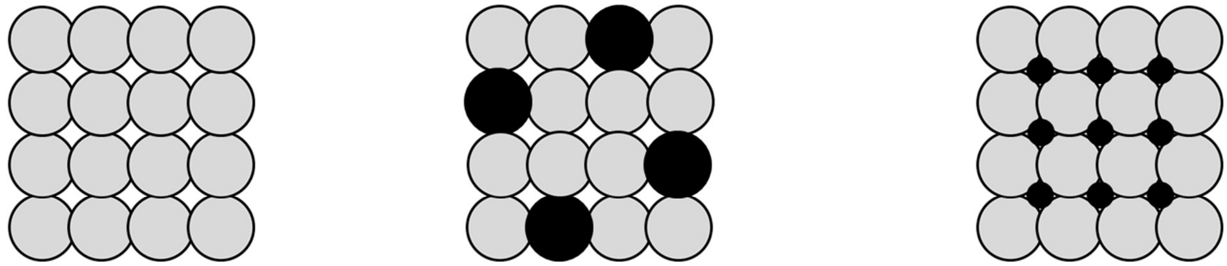
17. Explain why a metallic solid can conduct electricity.




In addition to being good conductors of electricity, metals are malleable, which means that they can be pressed or hammered into flat, thin sheets. They are also ductile, which means that they can be pulled or stretched out into long, thin wires.




The ability of metals to be deformed and change their shape can be explained by the fact that the metal atoms form bonds to many other neighboring atoms, within the sea of mobile valence electrons. Changes in the positions of the atoms happen relatively easily, and the electrons are redistributed around the collection of atoms.



pure metal substitutional alloy interstitial alloy

A metal alloy is a mixture of two different metals, or possibly a mixture of a metal and a nonmetal atom. The two types of alloys are substitutional and interstitial.

A **substitutional alloy** is formed between two different atoms whose atomic radius values are similar to each other.

One example of a substitutional alloy is brass, which contains a mixture of copper (Cu) and zinc (Zn).

An **interstitial alloy** is formed between two different atoms in which the atomic radius of one atom is much smaller than that of the other. The smaller atoms can fit in the empty spaces in between the larger atoms.

One example of an interstitial alloy is steel, which contains iron (Fe) and and carbon (C).

An interstitial alloy is often stronger than the pure metal, because the presence of the smaller atoms in the empty spaces makes it more difficult for the atoms of the alloy to move and slide past each other.