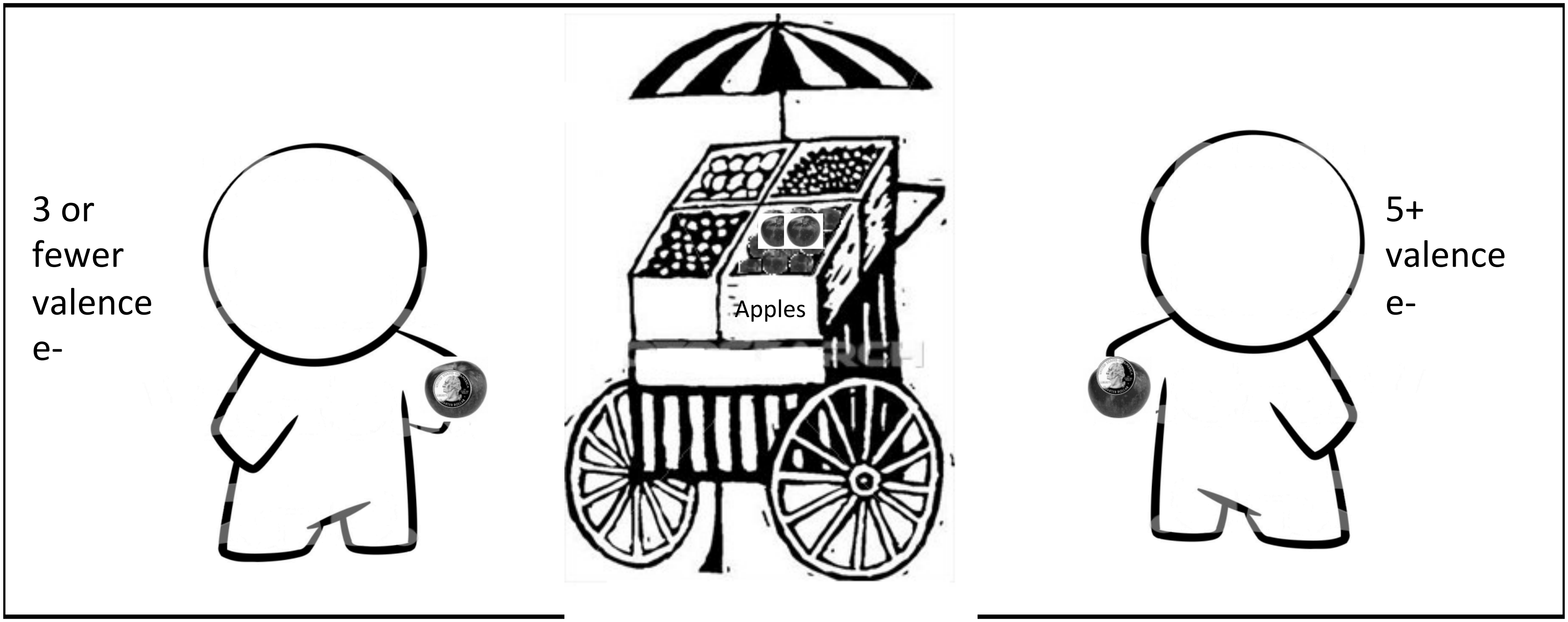
# Chapter 4
# Chemical Bonds

Mrs. Valentine

Physical Science

4th and 6th Period

# Section 1 - Ionic Bonds



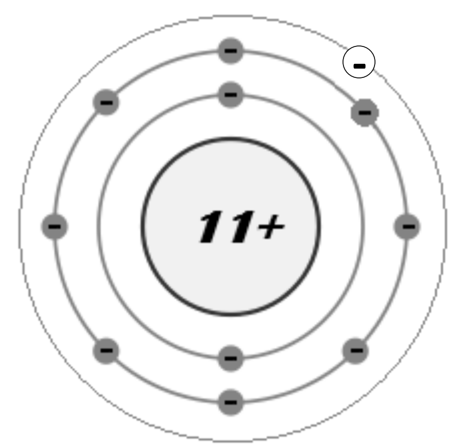

- Elements with fewer than 8 valence electrons behave as these two friends.
- Elements with more than 4 valence electrons are like the friend with not enough money, so they borrow enough electrons to have 8.
- Elements with fewer than 4 valence electrons tend to donate enough to their friend so that they have 0, and their friend has 8.

# What are Ions?

- This way, both atoms are in a lower energy state, and are more stable.

- **Definition:** <u>ion</u> – an atom or group of atoms that has become electrically charged.

- When an atom loses an electron, it loses a negative charge and becomes a positive ion.



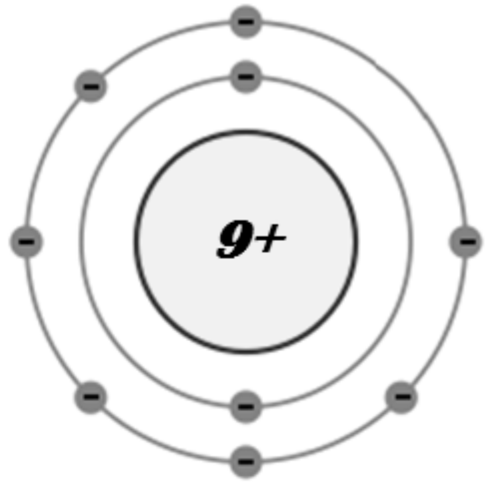

11 protons + 10 electrons = +1 charge
11 protons + 11 electrons = neutral

# Positive or Negative?

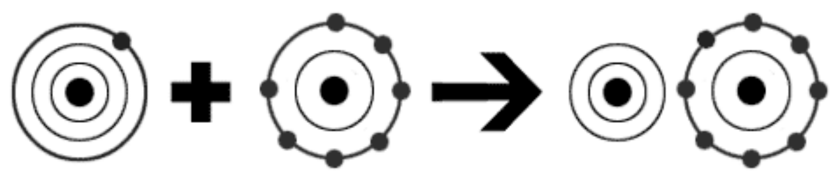
- When an atom gains an electron, it gains a negative charge and becomes a negative ion.

9 protons + 9 electrons = neutral

9 protons + 10 electrons = -1 charge




- Sodium and fluorine form ions and bond to create sodium fluoride.



Na + F → NaF

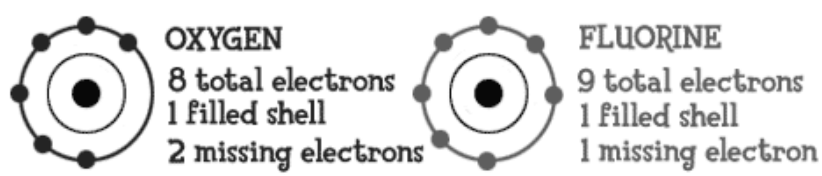
NaF is made of $Na^+$ and $F^-$

# Determining Ionic Charge

- Atoms transfer electrons to become more stable.

- Recall: Atoms are most stable with 0 or 8 **valence** electrons.

- Therefore, they will either give up all of their valence electrons or gain enough to have 8.



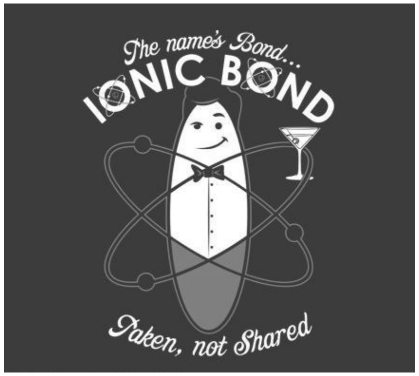

- The resulting ionic charge has the magnitude of the number of electrons transferred. It will be positive if electrons were lost or negative if electrons were gained.

# Forming Ions Practice Problems (p. 1 in packet)

- Using your periodic table only, determine the ions that each of the following elements will form.

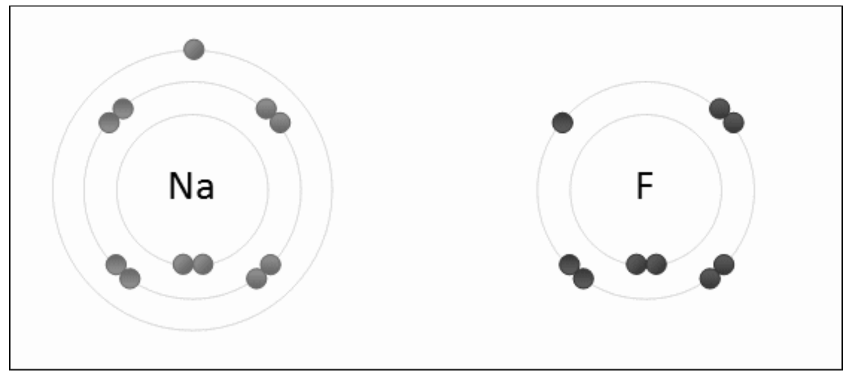
| | | | |
|---|---|---|---|
| Na | $Na^{+}$ | O | $O^{2-}$ |
| Mg | $Mg^{2+}$ | N | $N^{3-}$ |
| Al | $Al^{3+}$ | Cl | $Cl^{-}$ |
| Sr | $Sr^{2+}$ | S | $S^{2-}$ |
| Li | $Li^{+}$ | F | $F^{-}$ |

# Forming an Ionic Bond




- **Definition:** Ionic Bond – the attraction between two oppositely charged ions.

- In the case of sodium fluoride, the positively charged sodium ion ($Na^+$) is attracted to the negatively charged fluoride ion ($F^-$).




- Every sodium ion (positive charge) is balanced out with a fluoride ion (negative charge).

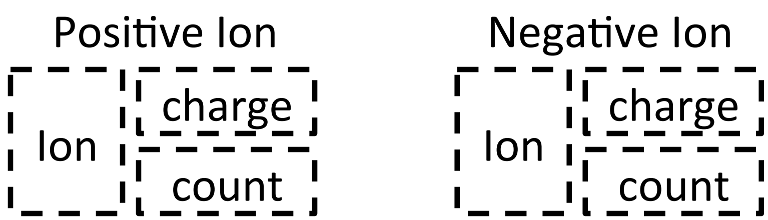
- The formula (NaF) shows that the sodium and fluorine are in a 1:1 ratio.

# Ionic Compounds

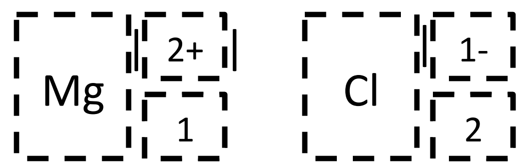
- Compounds are electrically neutral.
- Therefore, the ions must come together in such a way as to cancel out any charge.
- So what if we wanted to combine magnesium and chlorine?
- Magnesium forms an ion with a 2+ charge. Chlorine forms an ion with a 1- charge.
- The cross-over method can be used to determine the formula for this compound.

# Cross-Over Method

- The cross-over method involves crossing over the charges of the ions to determine how many of each is needed.
- The set up:




- Let's set it up for our example:




- Next, take the absolute value of the charges.
- Then, cross the charges into the count boxes.
- As we can see above, there would need to be two chloride ions for each magnesium in order to get a neutral compound.

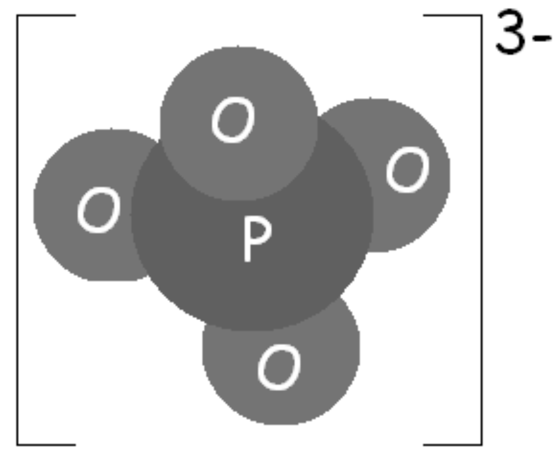
# Ionic Compounds


- Therefore, the compound formed with magnesium and chlorine is $MgCl_2$.
- The subscript 2 indicates that there are two of these ions. Subscripts of 1 are assumed, and therefore are not included in the formula.
- Subscripts are reduced to the simplest whole-number ratio for ionic compounds.
- Typically, ions are formed from the combination of a metal with a nonmetal.
- Try coming up with the formulas for the compounds formed from the following (p. 2 in packet):

| | | | | | |
|---|---|---|---|---|---|
| Ca and O | $CaO$ | Be and Cl | $BeCl_2$ | Ga and O | $Ga_2O_3$ |
| K and Br | $KBr$ | Al and F | $AlF_3$ | Li and S | $Li_2S$ |

# Polyatomic Ions

- **Definition:** Polyatomic Ions – ions made of more than one atom.
- The prefix *poly* means "many," so polyatomic means "many atoms."
- Think of these as groups of atoms acting as one.
- Let's look at the example to the right.

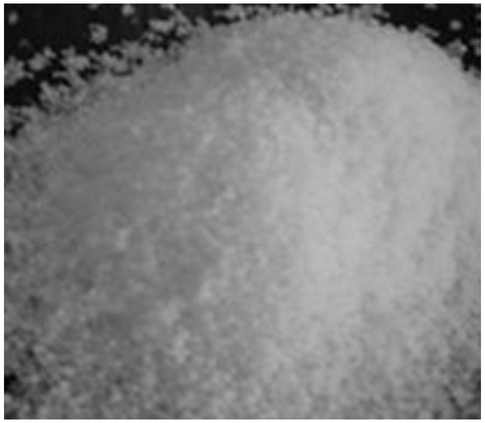



- Other examples can be found in Figure 2.

# Polyatomic Ion Compounds

- When a polyatomic ion reacts with another ion of opposite charge, an ionic compound forms.

- For example, calcium carbonate ($CaCO_3$) is made of calcium ions ($Ca^{2+}$) and carbonate ions ($CO_3^{2-}$). is a main component of limestone.

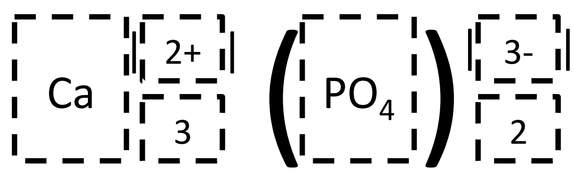


- Another example includes sodium sulfate ($Na_2SO_4$). This is made of sodium ions ($Na^+$) and sulfate ions ($SO_4^{2-}$).

# Cross-Over Method with Polyatomic Ions

- As polyatomic ions are single units, the cross-over method can be used to determine formulas containing them.

- Let's determine the formula for the compound made from calcium ion ($Ca^{2+}$) and phosphate ion ($PO_4^{3-}$). Since phosphate is a single unit, <u>all</u> of the ion's atoms go in the large box:

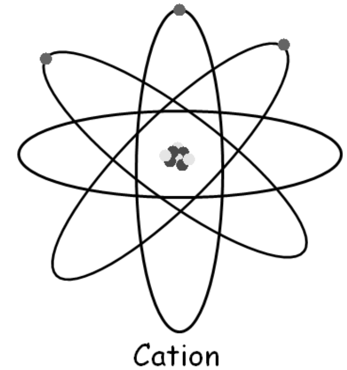



- This time, the large box for the polyatomic ion goes in parentheses. This is to avoid confusion with the subscripts.

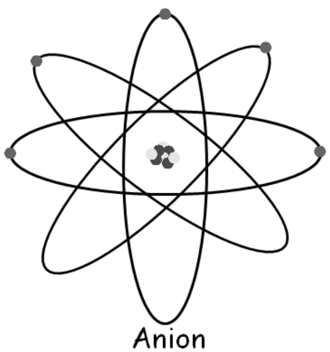
- Then, proceed normally.

- In this case, the formula will be $Ca_3(PO_4)_2$.

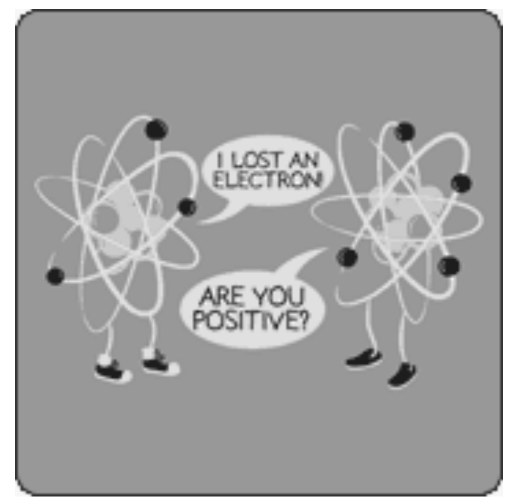
# Ion Types and Naming Ionic Compounds




- **Definition:** Cations – ions with a positive charge. These can be either monatomic or polyatomic ions.




- **Definition:** Anions – ions with a negative charge. These can be either monatomic or polyatomic ions.




- When naming an ionic compound, the cation always comes first, anion second.

- Usually, the name of the positive ion is the name of a metal.

# Naming Ionic Compounds

- The cation may also be a polyatomic ion, such as ammonium.
- If the anion is an element, the end of the name changes to *–ide*.
- For example, MgO is called "magnesium oxide."

- Try naming the following compounds (p.2 in packet):

| | | | | | |
|---|---|---|---|---|---|
| NaCl | sodium chloride | LiI | lithium iodide | CaO | calcium oxide |
| MgS | magnesium sulfide | $BeF_2$ | beryllium fluoride | $Na_2O$ | sodium oxide |

# Naming Ionic Compounds

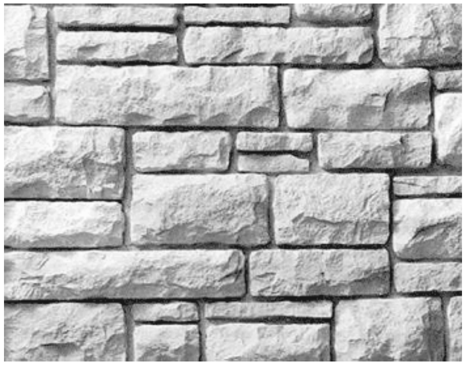
- If the anion is polyatomic, its name is unchanged.

- For example, $Na_2CO_3$ is called "sodium carbonate."

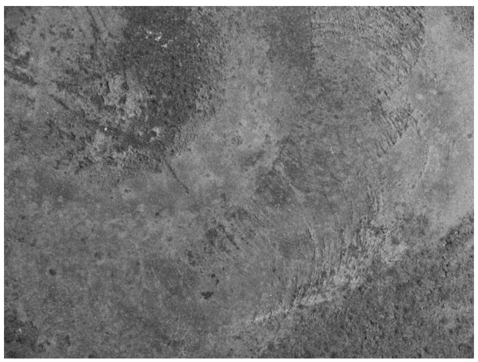
- Try naming these other compounds (p.3 in packet):

| | | |
|---|---|---|
| $Al_2(SO_4)_3$ | $NH_4Cl$ | $MgCO_3$ |
| aluminum sulfate | ammonium chloride | magnesium carbonate |
| $K_3PO_4$ | $CaHCO_3$ | $LiNO_3$ |
| potassium phosphate | calcium bicarbonate | lithium nitrate |

# Properties of Ionic Compounds



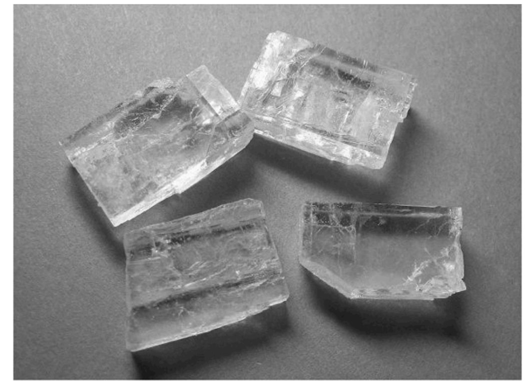
- Even though some ionic compounds may not appear to be anything alike, they may have some similarities.



- For example, limestone, rust, and table salt.
- You might not be able to use them for the same things, and they may not look the same, but they still share some properties.

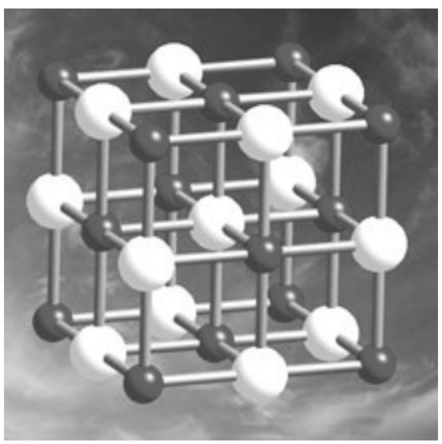
- **The characteristic properties of ionic compounds include crystal shape, high melting points, and electrical conductivity.**

# Crystal Shape



- All halite (table salt) samples have sharp edges, corners, and flat surfaces, whether the halite is in bricks or small crystals.

- The sodium chloride packs together in an alternating pattern.

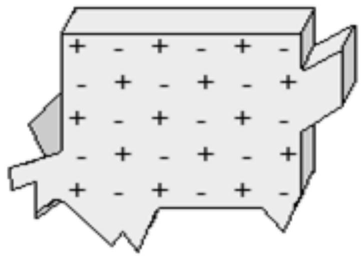


- **Definition:** <u>Crystal</u> – an orderly, three-dimensional arrangement of atoms or ions in a solid.

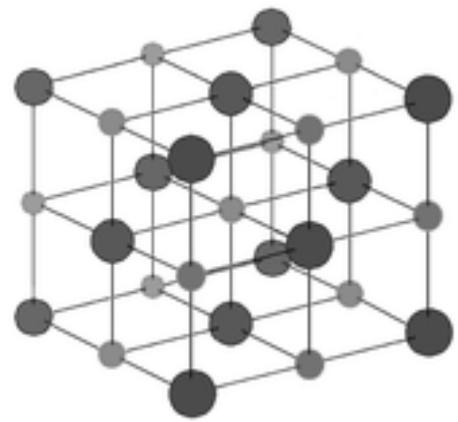
- Crystals of a single compound tend to have a similar shape, while crystals of different compounds differ in shape.

# Crystal Shape

- In an ionic compound, every ion is attracted to ions near it that have an opposite charge. Like charges repel each other.



ionic crystal



- One sodium ion does not just bond to one chloride ion. They bond to all those around them (in all directions) to form the crystal.

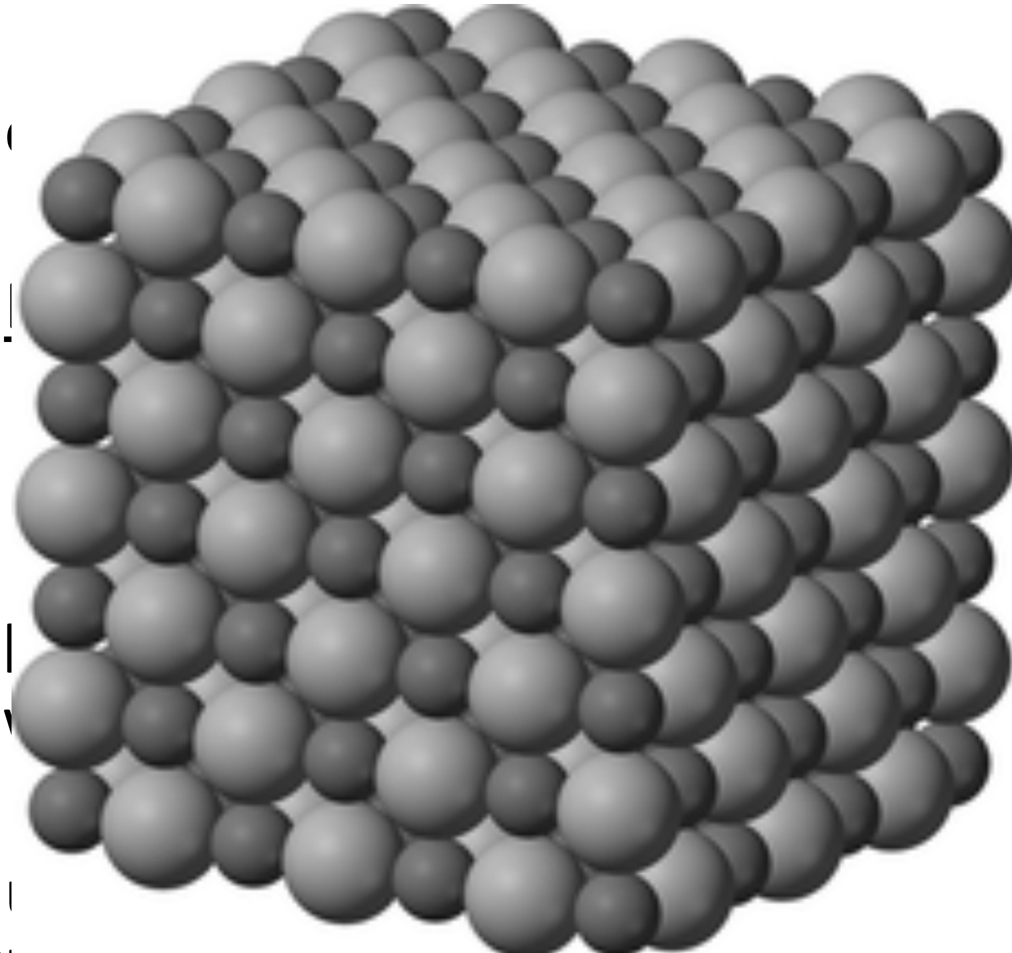
- The number of sodium ions and chloride ions exist in a 1:1 ratio. The formula, NaCl, represents this ratio.

# High Melting Points

- In order to melt the crystal of salt, the crystal structure must be broken apart.
- Since the structure is held together by ionic bonds, these bonds must be broken.
- For table salt, this takes a temperature of 801°C.

- Ionic bonds are very strong, strong enough to cause all sorts of compounds to be solid at room temperature.

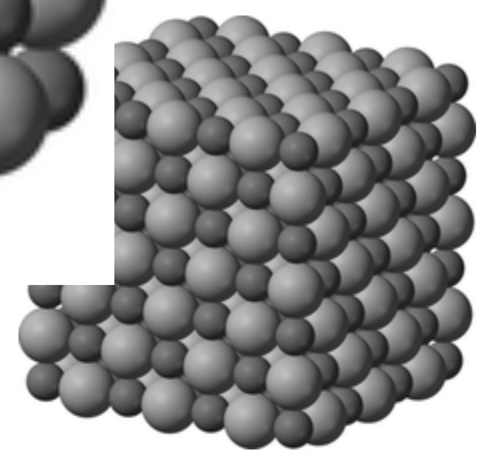
# Electrical Conductivity



- When ionic … solution conducts el…
- **Definition:** … rge.

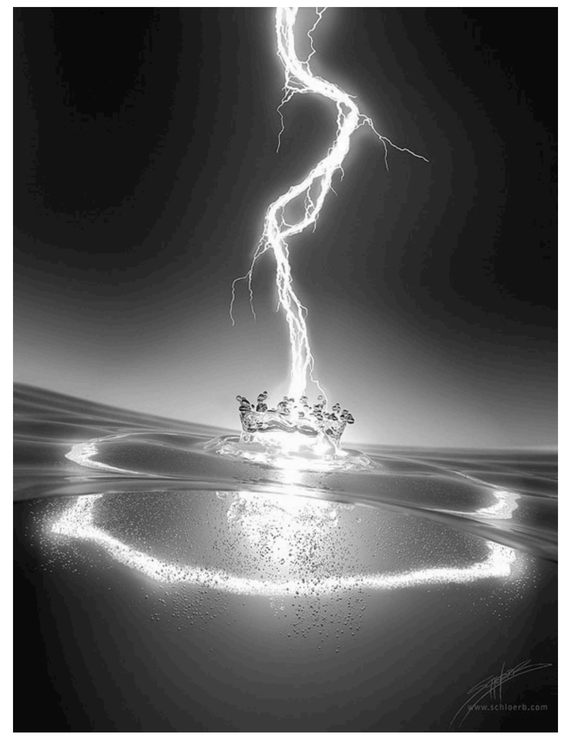
∘⟶

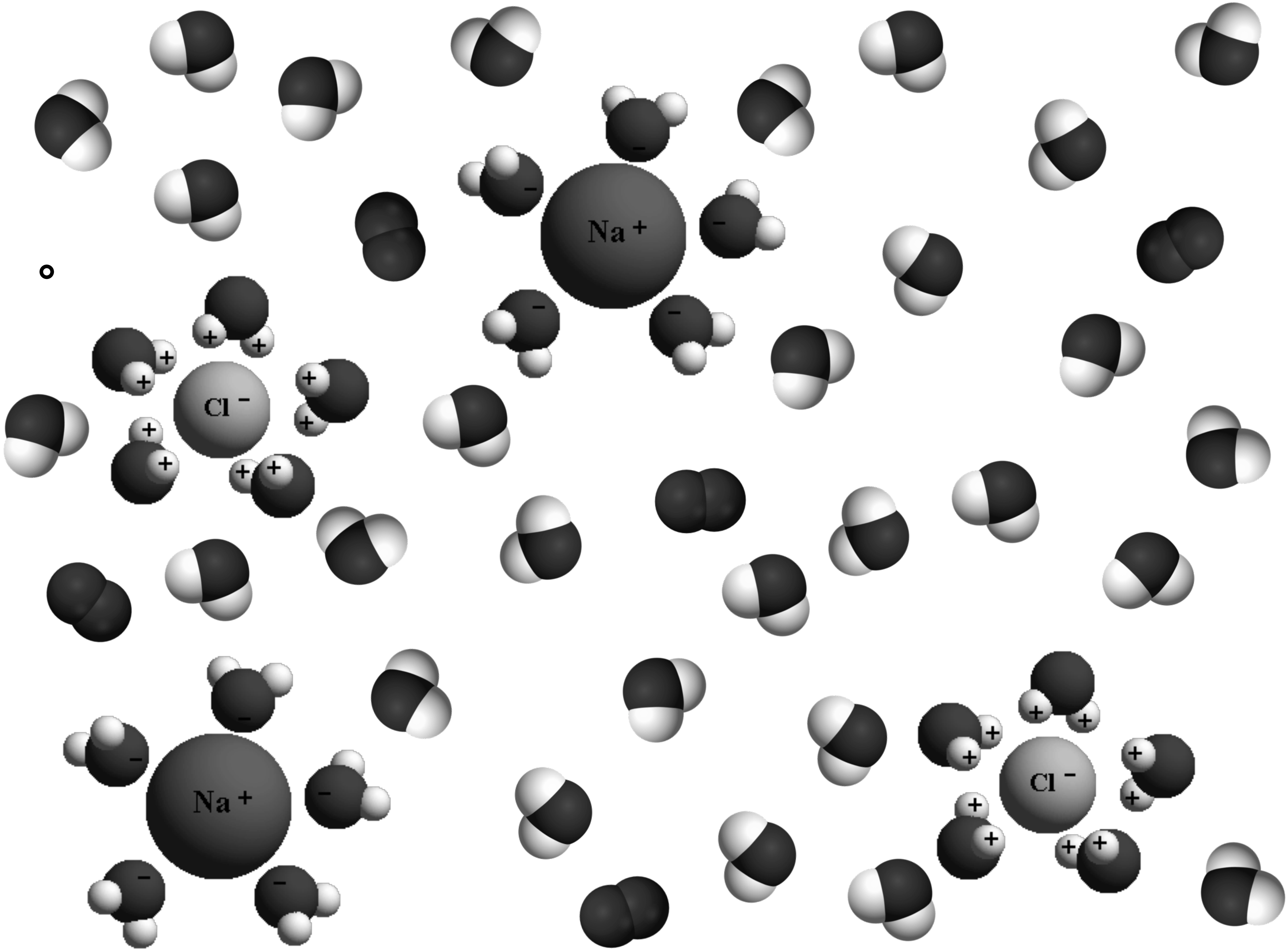
- Ions have el… nic compound …
- This is becau… bound together when they are together as a solid.



# Electrical Conductivity

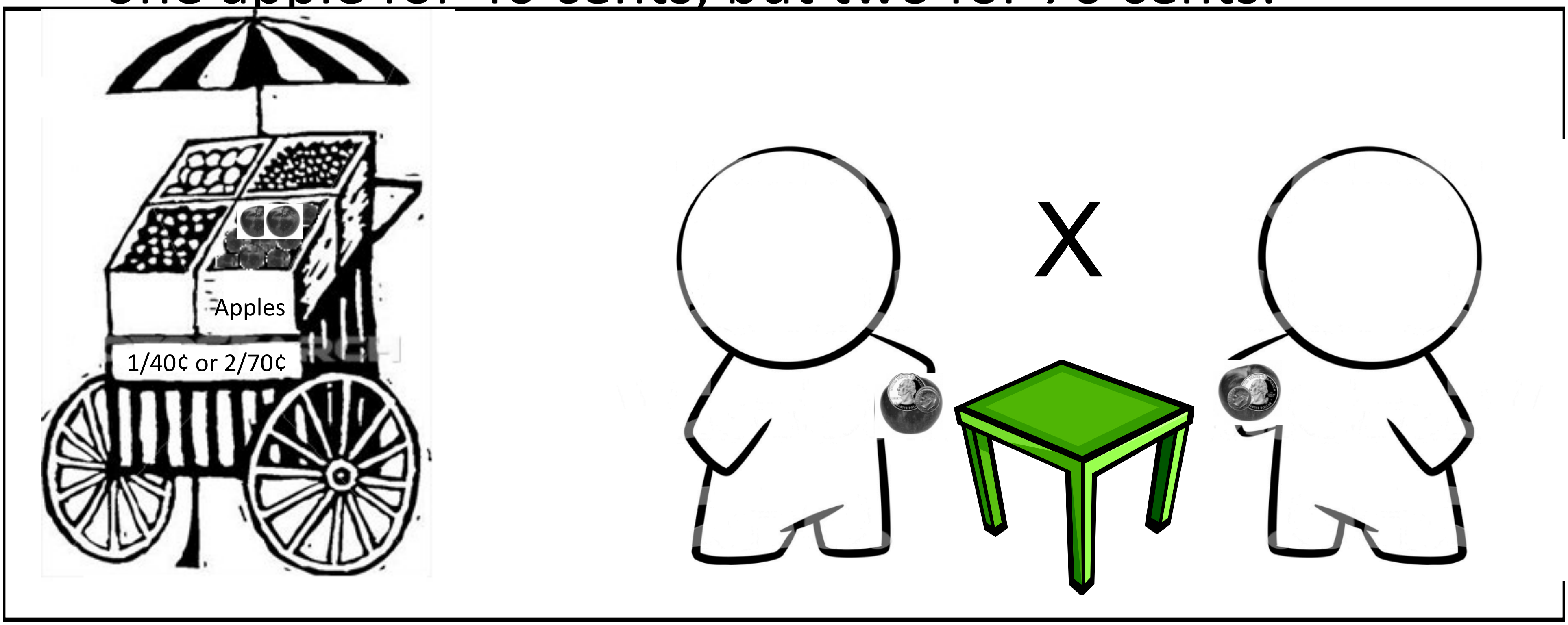
- When ionic compounds are dissolved in water, they break apart and separate.
- These free-moving ions in the solution are what enables it to conduct electricity.
- Melting also allows them to conduct electricity.




Na +
Cl −
Na +
Cl −

# Section 2 - Covalent Bonds

- Back to the apple analogy.  Let's say that you can get one apple for 40 cents, but two for 70 cents.

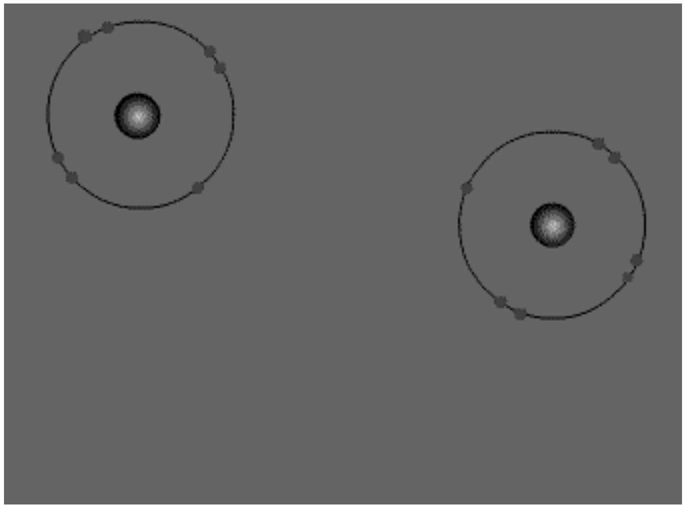



- They could pool their money together so that they could each get an apple.  This is like covalent bonds.

# Electron Sharing

- **Definition:** Covalent Bond – a chemical bond formed when two atoms share electrons.

  For example, when these fluorine atoms approach each other, the atoms share electrons and stay together.

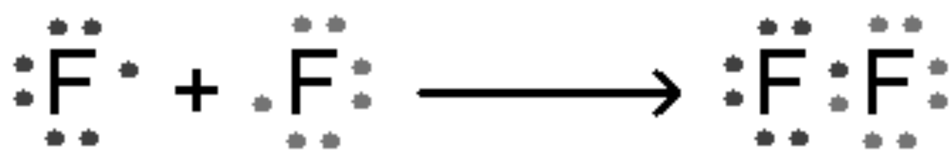


- Unlike ionic bonds, which form between metals and nonmetals, covalent bonds form between two nonmetals.
- Oxygen, carbon, nitrogen, and the halogens are all examples of nonmetals that frequently bond with other nonmetals.

# Electron Sharing

- Recall that fluorine forms a diatomic molecule. Each of the fluorine atoms has 7 valence electrons.

- Each one shares one of these with the other.




- When you count the number of electrons on one atom in the molecule, you count both of the shared electrons each time.

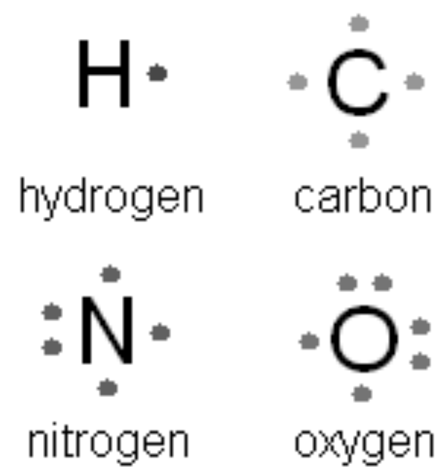
- In a covalent bond, both atoms attract the two shared electrons at the same time.

# How Many Bonds?

- Let’s look at the electron dot diagrams for hydrogen, oxygen, nitrogen, and carbon atoms.

- **The number of bonds each of these can form equals the number of electrons needed to make a total of eight.**

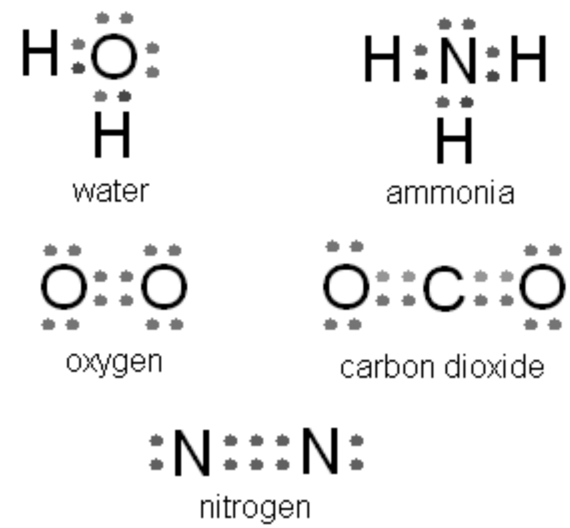
- For example, oxygen has six valence electrons, so it can form two covalent bonds.




- Since nitrogen has 5 valence electrons, it can form three bonds.

# Multiple Bonds

- Some of the molecules formed by these elements can be seen in the figure below.



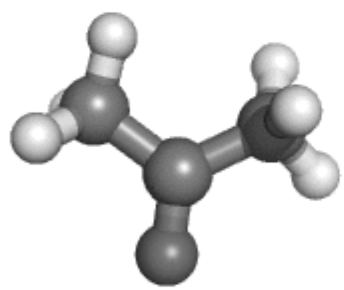

- **Definition:** <u>Double Bond</u> – two pairs of electrons are shared between two atoms.

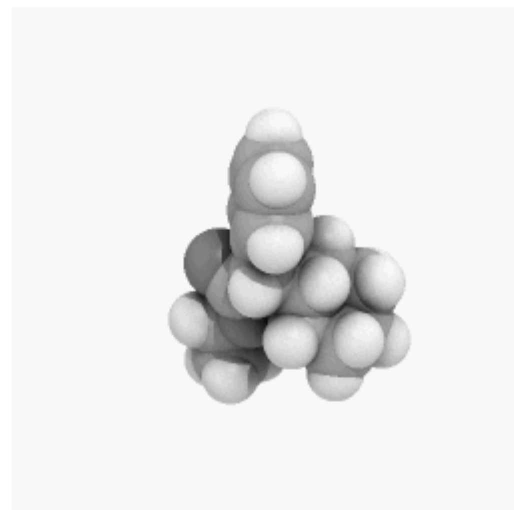
- Sometimes, two atoms can even share three pairs of electrons. This is called a <u>triple bond</u>.

- Remember that shared pairs of electrons count towards both atoms.

# Properties of Molecular Compounds

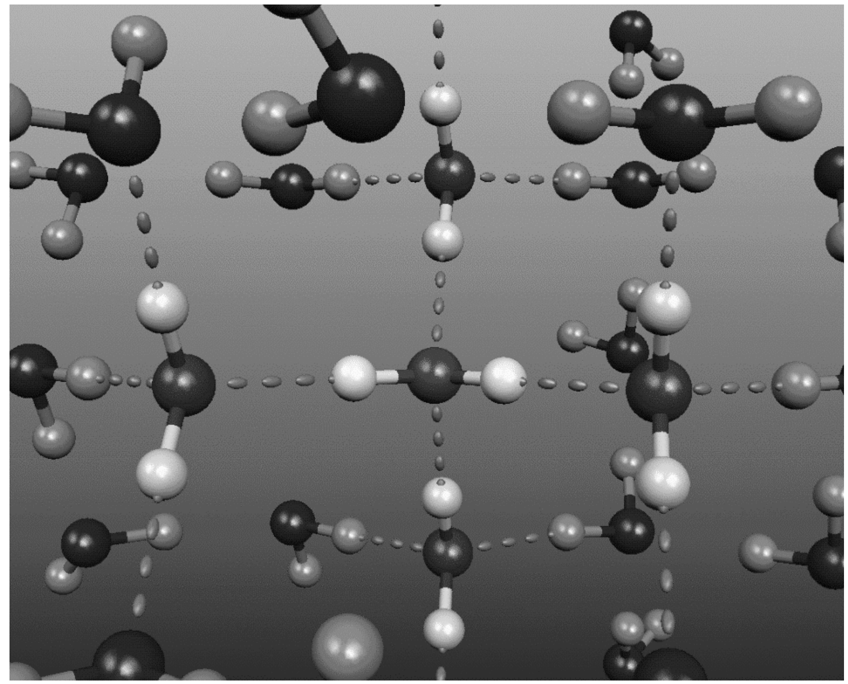
- **Definition:** Molecular Compounds – molecules having covalently bonded atoms.





- The properties of molecular compounds are very different than ionic compounds.
- These are very different than the very high temperatures that ionic compounds require.

# Properties of Molecular Compounds

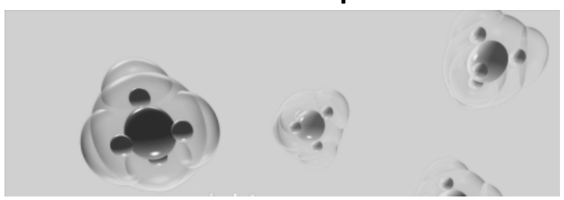


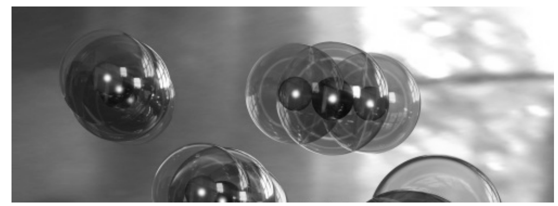
- The forces between any two molecules are much weaker than those holding ions together.
- Some molecular compounds do form crystals, such as water and sugar, but they melt and boil at much lower temperatures than ionic crystals.
- Most molecular compounds are poor conductors. This is why some molecular compounds, such as rubber, are used as insulators for electrical wires.

# Sharpen Your Skills

- Hypothetically, you have two samples of colorless, odorless gas.

- One is methane ($CH_4$) and the other is carbon dioxide ($CO_2$).





- Use Figure 8 on page 122 to find a way to determine which gas is which.

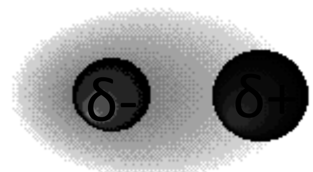
- What conditions would you change? What conditions would you control?

- What result would you look for to get an answer?

# Unequal Sharing of Electrons

- The sharing of electrons can be like a game of tug-of-war.

Note:
δ = partial




- If the pull is not the same on both sides, the electrons will move closer to the atom with the stronger pull.

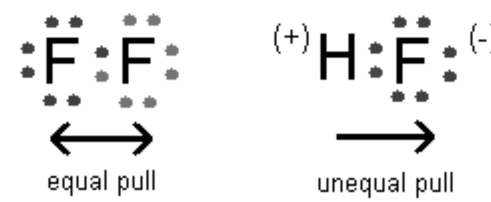
- This causes the atoms to have slight electrical charges.

- These charges are not as strong as those in ions.

# Unequal Sharing of Electrons

- **Definition:** Polar – describes a covalent bond in which electrons are shared unequally.

- If two atoms pull equally on the electrons, then neither atom becomes charged.

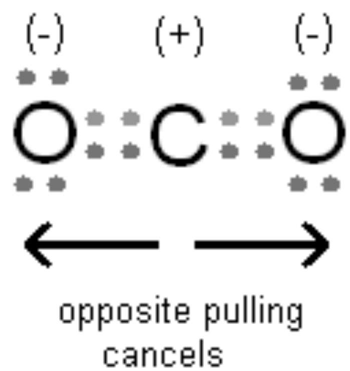



- This is the case when the two atoms are identical, as in fluorine gas ($F_2$).

- **Definition:** Nonpolar – describes a covalent bond in which electrons are shared equally.

# Nonpolar Molecules

- Let's look at carbon dioxide.

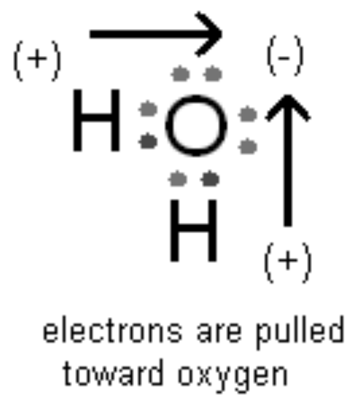



- Since the two oxygen atoms are pulling in equal and opposite directions, they, in a sense, cancel each other out.

- **NOTE:** A molecule can be nonpolar, even though it has polar bonds.

# Polar Bonds

- Let’s look at the water molecule.




- The oxygen atom is pulling the electrons from the hydrogen atoms closer to it.

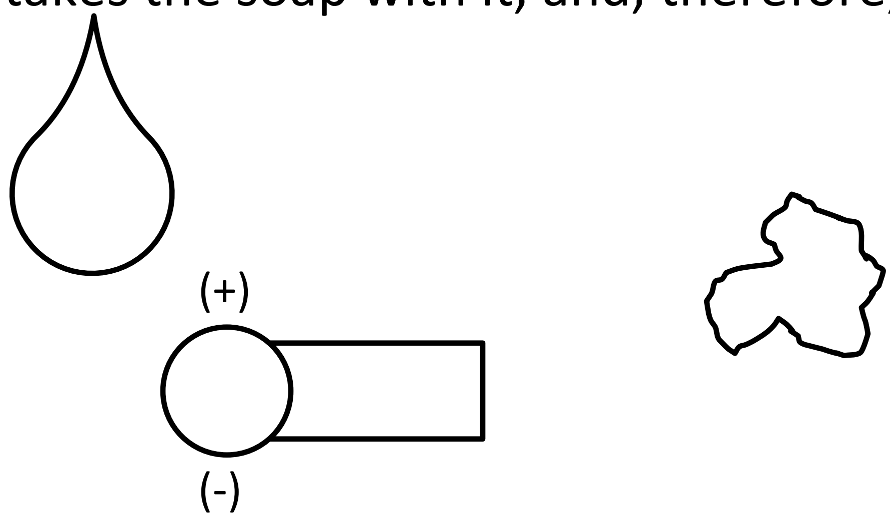
- Overall, the molecule is polar, since the two pulls are not opposite, though they are equal in magnitude.

# Attraction Between Molecules

- The slightly positive and slightly negative ends of a water molecule act like the poles of a bar magnet.
- They attract the opposite ends of other water molecules.
- **Differences in the attractions between molecules lead to different properties in polar and nonpolar compounds.**
- For example, water and vegetable oil don't mix. Oil is nonpolar, and nonpolar compounds do not dissolve in water.

# How Soap Works

- Soap has a polar end and a nonpolar end.
- The nonpolar dirt or oil is attracted to the nonpolar end of the soap.
- The polar end of the soap is then attracted to the polar water.
- As the water flows, it takes the soap with it, and, therefore, the dirt.



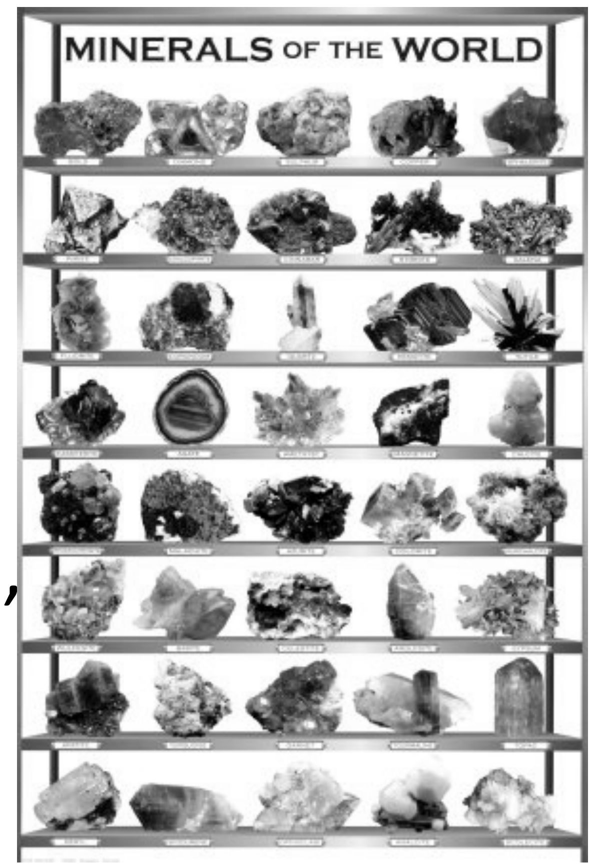

☐ Water
☐ Soap – polar end
☐ Soap – nonpolar end
☐ Dirt

# Section 3
# Crystal Chemistry

- **Definition:** Mineral – a naturally occurring solid that has a crystal structure and a definite chemical composition.

- There are a few minerals that are elements, such as sulfur and gold.

- Minerals can be identified by their properties, such as color, shininess, density, crystal shape, hardness, and magnetism.

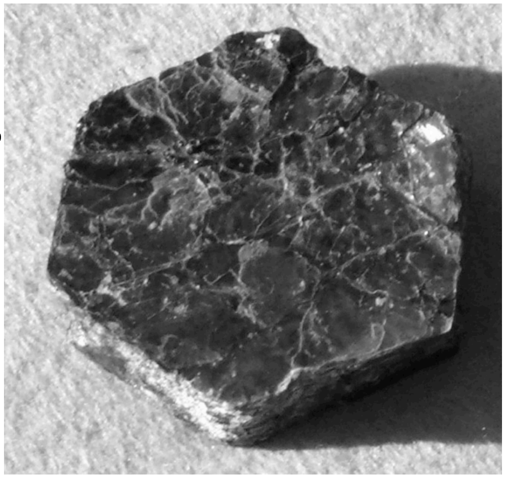


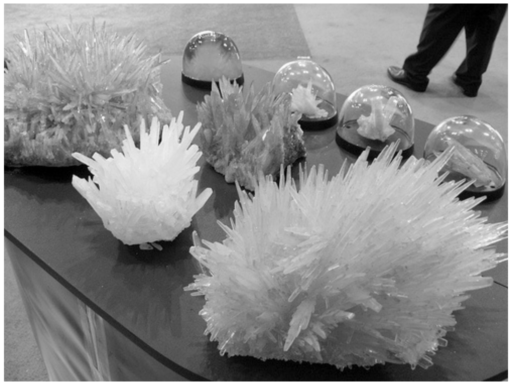
# Mineral Properties

- Some of these properties are easily observed, such as color.
- Others required measurements or testing, such as density.
- Talc, the softest mineral, can easily be scratched with your fingernail.
- Diamond is the hardest mineral.
- Others are in between.

# Mineral Properties

- Another key property is the way a mineral breaks apart.
- Some break apart into regular shapes.
- Mica, for example, breaks apart along flat surfaces at sharp edges.





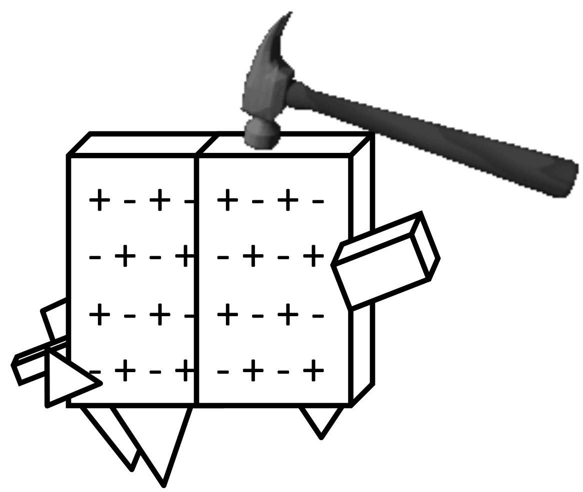
- Crystals also grow in characteristic shapes.

# Bonding in Mineral Crystals

- The structure of a crystal is a characteristic property of a mineral.
- Minerals can be made of either ionic compounds or molecular compounds.
- **The arrangement of particles in a mineral and the kind of bonds holding them together determine properties, such as crystal shape, hardness, and the way the crystal breaks apart.**

# Ionic Crystals

- Halite, for example, is an ionic crystal.
- The opposite charge of the sodium and chloride ions are laid out like a 3D checkerboard.
- Let’s look at how ionic crystals break.



# A Molecular Crystal

- Quartz ($SiO_2$) is an example of a molecular mineral.
- The covalent bonds in quartz are much stronger than the ionic bonds in halite.
- It cannot be crushed into predictable shapes with a hammer.
- Instead, it breaks into smaller, irregularly shaped pieces with shell-like ridges similar to broken glass.

# Comparing Crystals

- Not all ionic crystals are similar in properties to halite, and not all molecular crystals are similar in properties to quartz.
- Some molecular crystals are stronger than quartz. The stronger the bonds are, the stronger the crystals are.
- Many crystals are identifiable simply by sight. However, if there is doubt, further tests maybe performed.

# Chapter 4 Review

Answer questions

1-10, 12, 14, 16, 20-23

on pages 132-133.

Frank, D. V., Little, J. G., Miller, S., Pasachoff, J. M., & Wainwright, C. L. (2001). *Physical science*. Needham, Mass.: Prentice Hall.