

## Chapter 3 Lecture Notes: Compounds

### Educational Goals

1. Understand where electrons are located in atoms and how the locations of electrons affect the energy of the atom.
2. Define the term **valence electron** and draw the **electron dot structure** of an atom or ion.
3. Define the term **ion** and explain how the electron dot structure of a s- or p-block element can be used to predict the charge of the monoatomic ion.
4. Given the symbol, be able to name monoatomic **cations** and **anions** (and vice versa).
5. Explain the difference between an **ionic bond** and a **covalent bond**.
6. Understand the structural difference between ionic and covalent compounds.
7. Given the name, be able to write the formulas of ionic compounds and binary covalent compounds (and vice versa).
8. Define the terms **molar mass**, **formula mass**, and **molecular mass** and use these values in unit conversions involving moles and mass.
9. Given the formula, draw the **line bond structures** of diatomic molecules.

### The Arrangement of Electrons

Before we learn about compound, we must build on our understanding of atoms and electrons.

Specifically, in the beginning of chapter 3 you will learn:

- 1) **Where electrons are located in atoms.**
- 2) **How the location of electrons effect the energy of the atom.**

Scientists used **light** to study how electrons are arranged around the nucleus.

Energy, in the form of light or heat, can be \_\_\_\_\_ by atoms.

Energy is **absorbed** by \_\_\_\_\_ an electron to a **new** area.

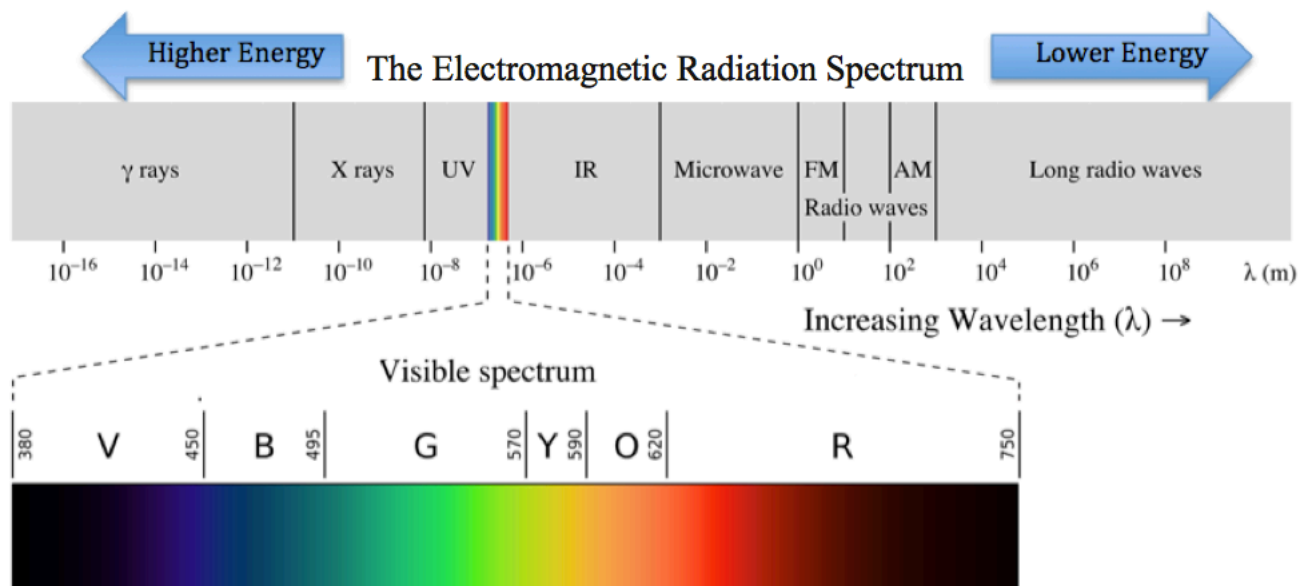
Atoms release energy when electrons move *back* to \_\_\_\_\_ areas.

- This can happen when an atom collides with another particle.
- Another way this can happen is by an atom emitting \_\_\_\_\_.

## The Modern Model of the Atom

New scientific laws and models of nature were needed to explain the pattern of light that was emitted by atoms.

Another word for *light* is *electromagnetic radiation*.



Visible light, the part of the electromagnetic spectrum that can be detected with the human eye, is a small part of the electromagnetic radiation spectrum (see the textbook for colored spectrum).

Short wavelengths correspond to higher energy; longer wavelengths correspond to lower energy light.

If **all energies** of light could be released from excited atoms, then we would expect the pattern of emitted light to look like this:



However, only light with **discrete** (distinct) energies is emitted. For example, the pattern of light emitted from excited hydrogen atoms is:



Our understanding of nature was dramatically changed when Max Planck and Albert Einstein introduced “\_\_\_\_\_.”

They proposed that energy is absorbed and emitted by atoms **only** in \_\_\_\_\_ amounts called **quanta**.

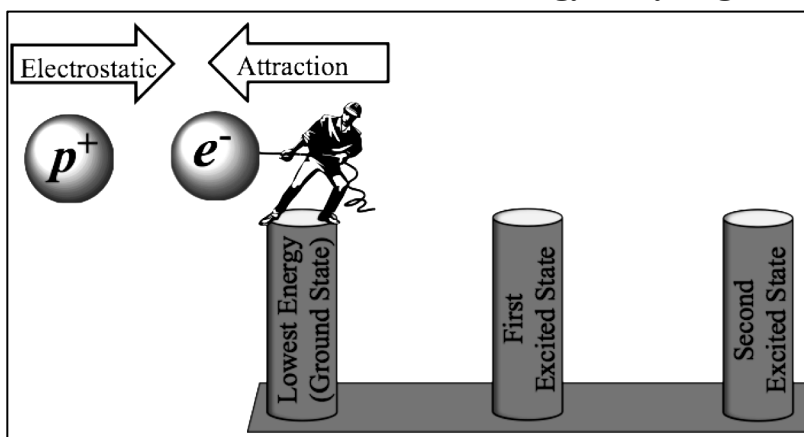
- Another word for “**discrete**” is “**distinct**.”

Recall that the light emitted from excited atoms is generated by electrons losing energy as they move from areas further from the nucleus (high energy) to areas nearer the nucleus (low energy). To lose the energy in this process, atoms emit light.

The observation that only discrete energies are emitted from excited atoms is explained using an atomic model that says that the **electrons can only exist in certain areas and therefore atoms have discrete energies**.

- We say that the energy of atoms is “\_\_\_\_\_.”
- The first scientist to propose a model of the atom where electrons existed in specific regions that had discrete energies was Niels Bohr.

## Illustration of the “Quantized” Energy of Hydrogen



When an atom's electron(s) are in the lowest possible energy area, we call this the \_\_\_\_\_.

- At room temperature, all atoms will exist in their ground state unless *temporarily* excited to a higher energy area by absorbing light.

Absorption of a discrete amount of energy corresponds to the worker **only** being able to move to **particular areas** (represented by posts in the illustration above).

When hydrogen's electron is in any other region than the ground state (lowest energy), we call that an \_\_\_\_\_ of hydrogen.

The excited atom will soon lose energy as the electron moves back to the ground state position. When the energy lost is in the form of light, that light will be the color (wavelength) corresponding to the energy difference between the initial “excited” region and the final, lower energy region.

### The Modern Model of the Atom: The Quantum Mechanical Model

You can **avoid getting lost** in the detail (and wonder) of nature by focusing on the following two educational goals:

- 1) Understand where electrons are \_\_\_\_\_ in atoms.
- 2) Understand how the location of electrons affect the \_\_\_\_\_ of the atom.

### The Hydrogen Atom

Hydrogen is unique because it has only \_\_\_\_\_ electron.

Electrons exist in certain **three-dimensional regions** called \_\_\_\_\_.

Orbitals can be described by these properties:

1. The \_\_\_\_\_ an electron in a particular orbital is from the nucleus.
  - Since orbitals are three-dimensional and the electrons move (very quickly) within the orbitals, the distance an electron is from the nucleus is not constant (as it would be in a circular two-dimensional path). Therefore, we talk about the electron's **average distance** from the nucleus.
- As hydrogen's orbitals get larger, the average distance of an electron from the nucleus increases, therefore the \_\_\_\_\_ **the orbital** occupied by an electron, the \_\_\_\_\_ **the energy**. (As described in the illustration at the top of this page).

## 2. The three-dimensional \_\_\_\_\_ of the orbital.

- Not only do the sizes of orbitals vary, the shapes of orbitals vary as well.
- When the shapes of orbitals are shown as three-dimensional representations, the shapes represent the region that would contain the electron \_\_\_\_\_ of the time. The remaining 10% of the time, the electron would be outside of the shape that is shown in the graphic representation.

### The Language of Quantum Mechanics

The orbitals are *centered on the* \_\_\_\_\_, and are labeled by a \_\_\_\_\_.

In a hydrogen atom:

- This number is related to the orbital **size** and the **energy** of an electron in the orbital.
- The orbitals are numbered from **lowest** energy (smallest size) to **higher** energy (larger size).

These numbers are referred to as “energy level,” or “quantum number,” or “quantum level,” or “shell.”

- We will use the term “\_\_\_\_\_” or “\_\_\_\_\_” and abbreviate it by using “\_\_\_\_\_.”

In the lowest energy state of a hydrogen atom (the *ground state*), the electron occupies the **n=1** quantum level.

The **n=1** quantum level has \_\_\_\_\_ orbital.

- It is called an \_\_\_\_\_ orbital.
- “**s**” represents the **shape of the orbital**, we use **1s** because **n=1**).
- **s** orbitals are \_\_\_\_\_ in shape.

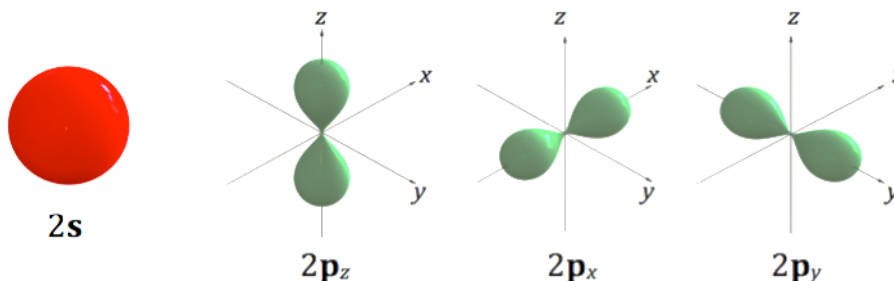
#### An Illustration of a 1s Orbital



- The \_\_\_\_\_ is in the **center** of the orbital.

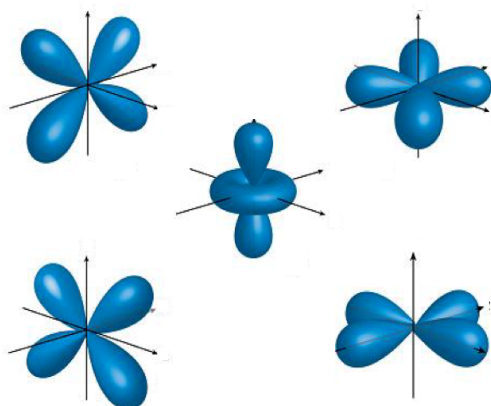
The **n=2** quantum level has \_\_\_\_\_ orbitals.

- There is **one 2s** orbital
  - **All s** orbitals are spherically shaped.
  - We use **2s** because **n=2**.
  - The major difference between the **2s** orbital and the **1s** orbital is that the **2s** orbital is **larger**.
- There are **three 2p** orbitals.
  - **p** represents the shape; we use **2p** because **n=2**.
  - The **p** orbitals all have the **same shape** and only differ in *how they are* \_\_\_\_\_ around the nucleus.



The **n=3** quantum level has \_\_\_\_\_ orbitals.

- There is **one 3s** orbital, **three 3p** orbitals, and **five 3d** orbitals.
- The shapes of the **3s** and **3p** orbitals are similar to those of the **2s** and **2p** orbitals, respectively, but they are *larger*.
- The **five 3d** orbitals are illustrated below:



As is the case for all orbitals, the **d** orbitals are centered on the nucleus.

Image Source: Wikimedia Commons, CK-12 Foundation, CC-BY-SA, <http://creativecommons.org/licenses/by-sa/3.0/legalcode>

The **n=4** quantum level has \_\_\_\_\_ orbitals.

- There is **one 4s** orbital, **three 4p** orbitals, **five 4d** orbitals, and **seven 4f** orbitals.
- The **f** orbitals have shapes that are even more complicated than the **d** orbitals.
- The shapes of the **4s**, **4p**, and **4d** orbitals are similar to those of the **3s**, **3p**, and **3d** orbitals, respectively, but they are *larger*.

The **n=5** level has **twenty-five** orbitals.

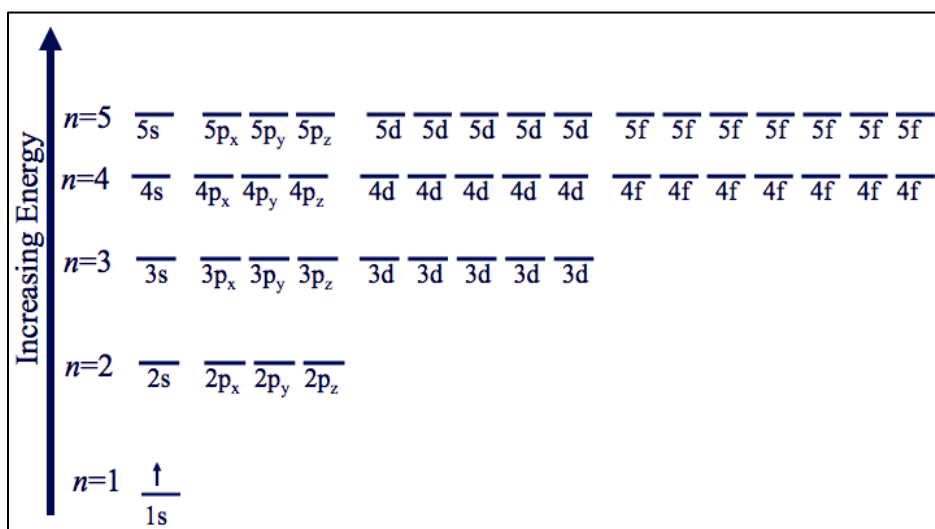
This just keeps going, **n= 6, 7, 8**, etc.

Although quantum levels with **n > 4** contain orbitals other than **s**, **p**, **d**, and **f**, these other orbitals are *never occupied by electrons of any element in its ground state*.

- The only time an electron can occupy any of those orbitals will be if the atom absorbs energy.

### Energy Level Diagram for Hydrogen

We can organize the various atomic orbitals according to their energy in an illustration called an **energy level diagram**. The energy level diagram for the first five quantum levels (**n = 1-5**) of a **hydrogen atom** is shown below.

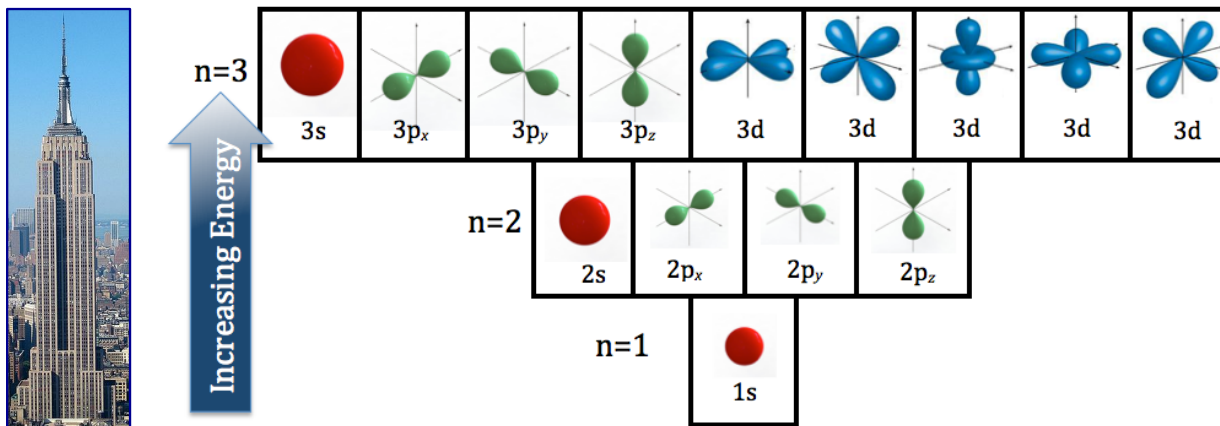


In an **energy level diagram**, we draw a short horizontal line that is labeled for each \_\_\_\_\_.

The orbitals are arranged, from bottom to top, in order of increasing \_\_\_\_\_.

An electron is depicted as an \_\_\_\_\_ above the line that represents the orbital occupied by it.

Let's compare the *energy level diagram* to a *skyscraper*, we will call this our **skyscraper model**.



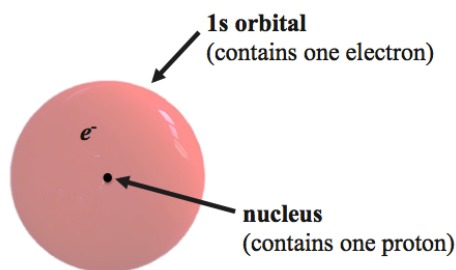
Skyscraper photo source: Wikimedia Commons, Author: Avala, CC-BY-SA, <http://creativecommons.org/licenses/by-sa/3.0/legalcode>

The different \_\_\_\_\_ of the skyscraper represent the quantum levels (**n**).

- The first floor is the lowest energy floor so it would correlate with **n=1**, the ground state.

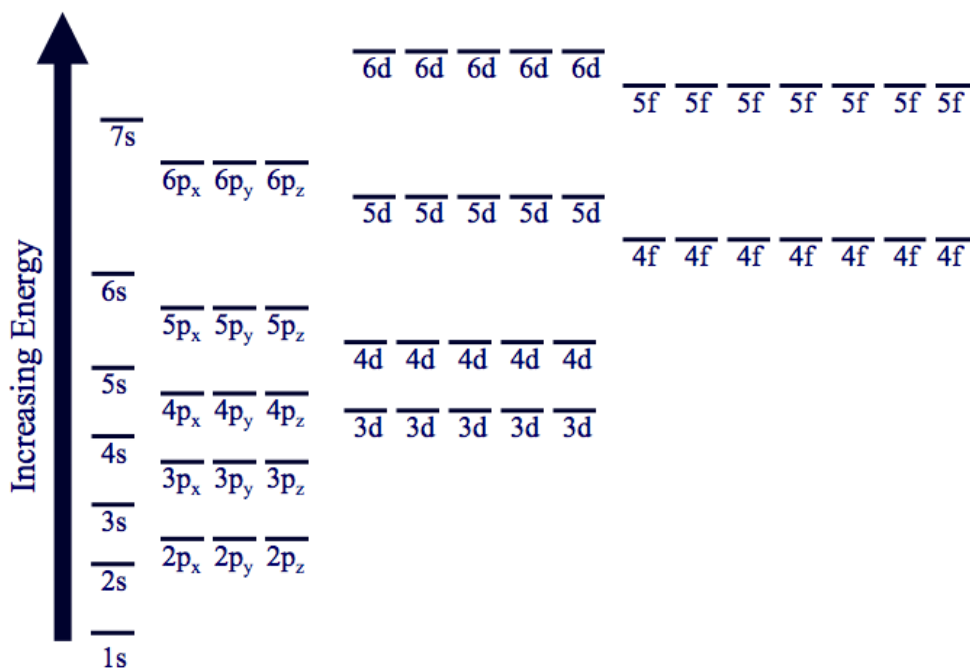
\_\_\_\_\_ on a particular floor are analogous to the various orbitals in a particular *quantum level*.

### Model of a Hydrogen Atom

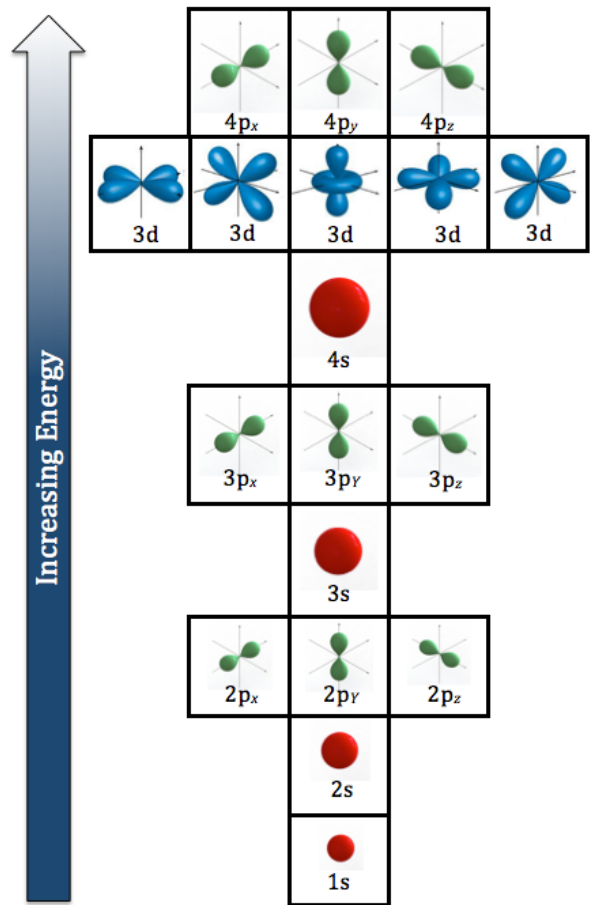


### Atomic Model for Multi-Electron Atoms

#### Energy Level Diagram for Multi-Electron Atoms



## Skyscraper Model for Multi-Electron Atoms



We live in a universe where matter tends to exist in the *lowest possible energy state*.

An informal way to state this is:

Nature wants everything to be at the \_\_\_\_\_ possible energy.

Electrons are arranged (configured) into the orbitals of multi-electron atoms in the way that results in the **lowest** possible energy.

Nature does this by obeying the following principles:

### 1) The Aufbau Principle

"Aufbau" (German) means *build-up* or *construct*.

The *aufbau principle* states that an electron occupies the **lowest energy orbital that can receive it**.

### 2) The Pauli Exclusion Principle

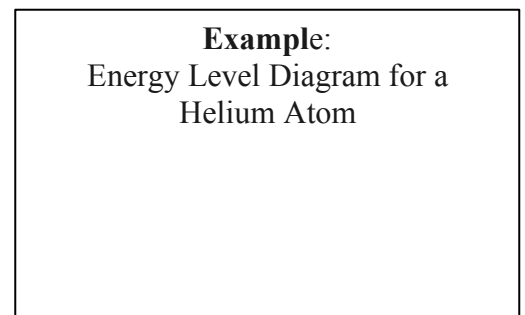
An orbital can hold a maximum of \_\_\_\_\_ electrons.

Electrons have a quantum mechanical property called **spin**.

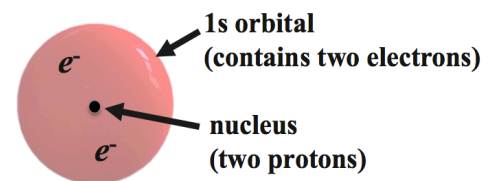
We call the spin states "**up**" or "**down**."

- *When two electrons occupy the same orbital, one electron has spin "**up**" the other has spin "**down**."*

This is all that you need to know about spin to understand all of the concepts covered in this textbook. You may find it interesting that spin is responsible for magnetic properties of matter. In fact, spin is the reason that electrons behave as tiny magnets!



**Model of a Helium Atom**

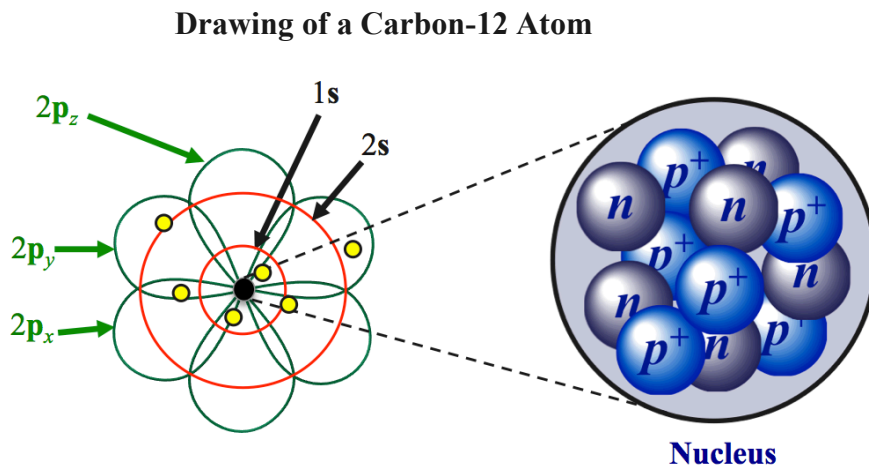


### 3) Hund's Rules

When electrons are configured into orbitals that all have the *same energy*, a **single electron** is placed into **each** of the equal-energy orbitals before a second electron is added to an occupied orbital.

When electrons are configured into a *set of orbitals* that all have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all “up”).

**Example:** Energy Level Diagram for a Carbon Atom



**Understanding Check:**  
Energy Level Diagram for a Neon (Ne) Atom

**Understanding Check:**  
Energy Level Diagram for an Iodine (I) Atom



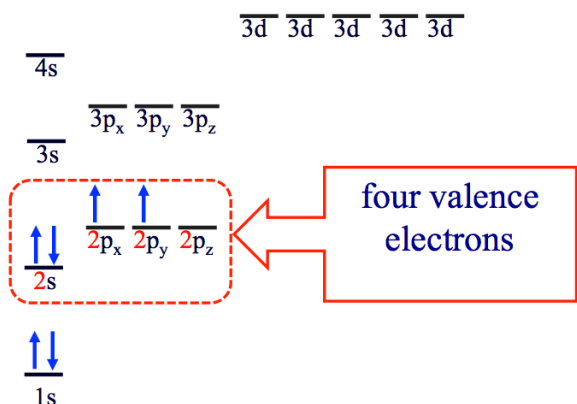
## Valence Electrons

Valence electrons are the electrons held in the \_\_\_\_\_ *shell* (largest "n").

**Language Reminder:** “shell” = “quantum level” = “energy level”

- Valence electrons are furthest away from the \_\_\_\_\_.
- It is important to know how many valence electrons are in an atom because:
  - These are the electrons that are involved in \_\_\_\_\_ to other elements to form \_\_\_\_\_.
  - These are the electrons that elements lose to become \_\_\_\_\_.

**Example:** How many valence electrons do **carbon** (C) atoms have?



**Understanding Check:** How many valence electrons do **oxygen** (O) atoms have?

## Short-Cut for Determining the Number of Valence Electrons

*Elements are arranged in the periodic table according to the number of valence electrons.*

For **s-** and **p-block** elements, all elements in the same periodic \_\_\_\_\_ (group) have the *same number of valence electrons* as all others in that column.

The group numbers for the columns *represent* the number of valence electrons contained in those atoms.

|    |    |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      |    |
|----|----|---------------------------|-----|-----|-----|-----|-----|-----|----|----|----|-----|-----|-----|-----|----|----|--|--|------|----|
| I  |    |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  | VIII |    |
| 1  |    |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      | 2  |
| H  | II |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      | He |
| 3  | 4  |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      | 10 |
| Li | Be |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      | Ne |
| 11 | 12 | Transition Metals         |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      |    |
| Na | Mg |                           |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      | 18 |
| 19 | 20 | 21                        | 22  | 23  | 24  | 25  | 26  | 27  | 28 | 29 | 30 | 31  | 32  | 33  | 34  | 35 | 36 |  |  |      |    |
| K  | Ca | Sc                        | Ti  | V   | Cr  | Mn  | Fe  | Co  | Ni | Cu | Zn | Ga  | Ge  | As  | Se  | Br | Kr |  |  |      |    |
| 37 | 38 | 39                        | 40  | 41  | 42  | 43  | 44  | 45  | 46 | 47 | 48 | 49  | 50  | 51  | 52  | 53 | 54 |  |  |      |    |
| Rb | Sr | Y                         | Zr  | Nb  | Mo  | Tc  | Ru  | Rh  | Pd | Ag | Cd | In  | Sn  | Sb  | Te  | I  | Xe |  |  |      |    |
| 55 | 56 | 57                        | 72  | 73  | 74  | 75  | 76  | 77  | 78 | 79 | 80 | 81  | 82  | 83  | 84  | 85 | 86 |  |  |      |    |
| Cs | Ba | La                        | Hf  | Ta  | W   | Re  | Os  | Ir  | Pt | Au | Hg | Tl  | Pb  | Bi  | Po  | At | Rn |  |  |      |    |
| 87 | 88 | 89                        | 104 | 105 | 106 | 107 | 108 | 109 |    |    |    |     |     |     |     |    |    |  |  |      |    |
| Fr | Ra | Ac                        | Rf  | Db  | Sg  | Bh  | Hs  | Mt  |    |    |    |     |     |     |     |    |    |  |  |      |    |
|    |    | (Inner) Transition Metals |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      |    |
|    |    | 58                        | 59  | 60  | 61  | 62  | 63  | 64  | 65 | 66 | 67 | 68  | 69  | 70  | 71  |    |    |  |  |      |    |
|    |    | Ce                        | Pr  | Nd  | Pm  | Sm  | Eu  | Gd  | Tb | Dy | Ho | Er  | Tm  | Yb  | Lu  |    |    |  |  |      |    |
|    |    | Lanthanides               |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      |    |
|    |    | 90                        | 91  | 92  | 93  | 94  | 95  | 96  | 97 | 98 | 99 | 100 | 101 | 102 | 103 |    |    |  |  |      |    |
|    |    | Th                        | Pa  | U   | Np  | Pu  | Am  | Cm  | Bk | Cf | Es | Fm  | Md  | No  | Lr  |    |    |  |  |      |    |
|    |    | Actinides                 |     |     |     |     |     |     |    |    |    |     |     |     |     |    |    |  |  |      |    |

Different elements with the same number of valence electrons are said to be \_\_\_\_\_.

Example of isoelectric elements: oxygen and sulfur

*Isoelectric* atoms often behave in similar ways. For example, oxygen atoms often chemically “bond” to two hydrogen atoms to form water (H<sub>2</sub>O); sulfur atoms, also often “bond” with two hydrogen atoms to form hydrogen sulfide (H<sub>2</sub>S).

**Understanding Check:** Use the periodic table to determine the number of valence electrons in each of these types of atoms:

- hydrogen (H)
- nitrogen (N)
- bromine (Br)
- krypton (Kr)

## Electron Dot Structures

*Electron dot structures* show the number of *valence electrons* that an atom carries.

- In these structures, *valence electrons* are represented by \_\_\_\_\_ drawn next to an element's symbol.

| I   |     |  |  |  |  |  |  |  |  |  |  | VIII |
|-----|-----|--|--|--|--|--|--|--|--|--|--|------|
| H•  | II  |  |  |  |  |  |  |  |  |  |  | He•  |
| Li• | Be• |  |  |  |  |  |  |  |  |  |  | •B•  |
| Na• | Mg• |  |  |  |  |  |  |  |  |  |  | •C•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •N•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •O•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •F•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •Ne• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Al• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Si• |
|     |     |  |  |  |  |  |  |  |  |  |  | •P•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •S•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •Cl• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Ar• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Ga• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Ge• |
|     |     |  |  |  |  |  |  |  |  |  |  | •As• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Se• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Br• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Kr• |
|     |     |  |  |  |  |  |  |  |  |  |  | •In• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Sn• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Sb• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Te• |
|     |     |  |  |  |  |  |  |  |  |  |  | •I•  |
|     |     |  |  |  |  |  |  |  |  |  |  | •Xe• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Tl• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Pb• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Bi• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Po• |
|     |     |  |  |  |  |  |  |  |  |  |  | •At• |
|     |     |  |  |  |  |  |  |  |  |  |  | •Rn• |
| Fr• | Ra• |  |  |  |  |  |  |  |  |  |  |      |

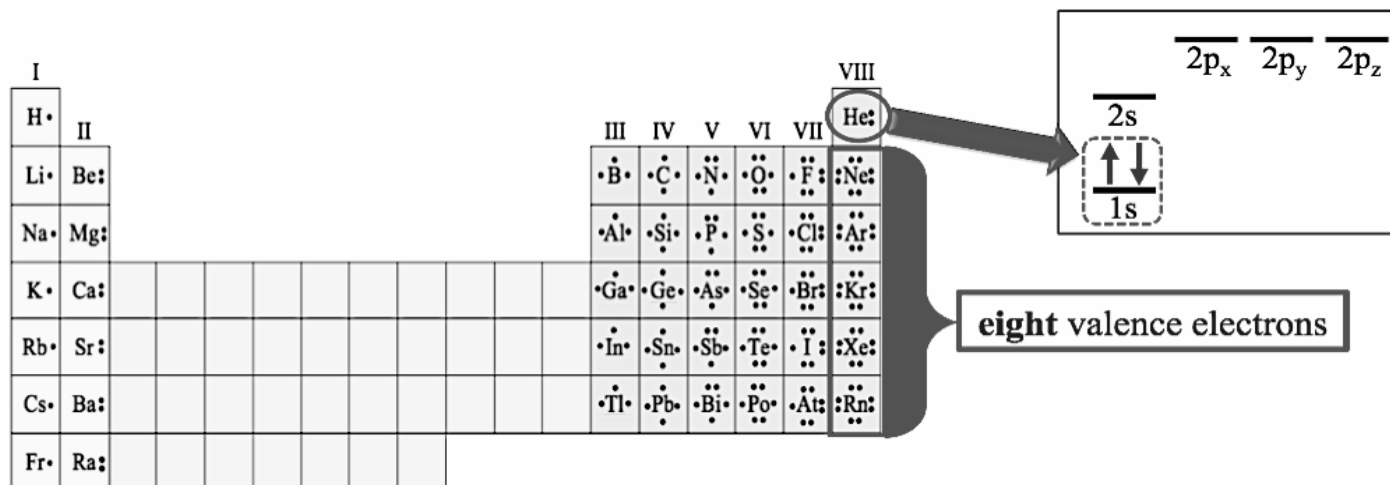
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## Noble Gases and the Octet Rule

The group VIII elements (He, Ne, Ar, Kr, Xe, and Rn) are called \_\_\_\_\_ *gases*.

He, Ne, Ar, Kr, Xe, and Rn belong to the **noble gas family**, which gets its name from the fact that these elements are resistant to change and, with few exceptions, do not lose or gain electrons.

The resistance to change (stability) of the noble gases is related to having their outermost quantum level (*shell*) completely \_\_\_\_\_ with electrons



*Helium's* outermost shell (the  $n=1$  quantum level) is completely filled with its *two* electrons.

*All of the other noble gas elements* have completely filled outermost shells with \_\_\_\_\_ electrons.

This stability of the noble gas elements that have *eight electrons* in their outermost shell led to what chemists call the \_\_\_\_\_.

Most substances around us do not exist as individual atoms. Atoms will “bond” with other atoms to form compounds such as water (H<sub>2</sub>O), carbon dioxide (CO<sub>2</sub>), and table salt (sodium chloride). In the remainder of chapter 3, we will discuss the nature of this “bonding” of atoms to other atoms.

**The Octet Rule is quite useful in predicting and understanding bonding patterns in chemical compounds.**

### The Octet Rule

Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an *octet* (eight) of electrons in its outermost shell.

There are exceptions to the octet rule. An important exception that we will always use is for *hydrogen* and *helium*.

- Hydrogen and helium have filled outer shells (are stable) with just *two* electrons because their outermost level ( $n=1$ ) has only one orbital.

# Ions

**Atoms** have the same number of electrons as protons and are therefore *electrically neutral*.

An \_\_\_\_\_ is a small particle that has an *electrical charge*.

Atoms can *gain or lose* \_\_\_\_\_ to become **ions**.

**Metal** atoms can \_\_\_\_\_ electrons to form *positive ions*.

If an atom *loses* one or more electrons, it will then have more protons than electrons and have an overall *positive charge*.

- *Positive ions* are called \_\_\_\_\_.

**Nonmetal** atoms can \_\_\_\_\_ electrons to form *negative ions*.

If an atom *gains* one or more electrons, it will then have more electrons than protons and have an overall *negative charge*.

- *Negative ions* are called \_\_\_\_\_.

## The Octet Rule in the Formation of Ions

The *Octet Rule* can be used to predict the formation of ions.

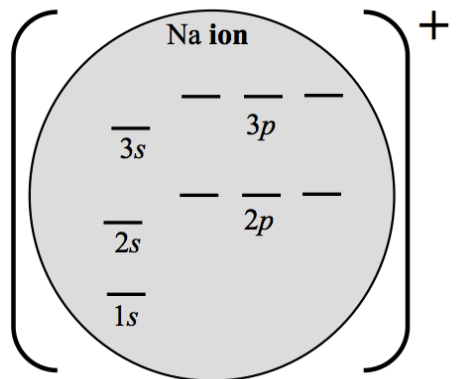
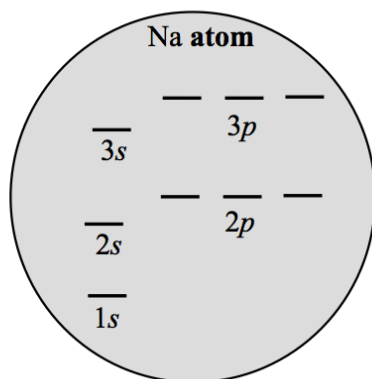
Very often, ions are formed such that the *ion* has an \_\_\_\_\_ in its outermost shell.

This tendency will allow us to predict the \_\_\_\_\_ of the ion that is formed for particular elements.

### Example: Let's do a Cation - Sodium (Na)

- A sodium atom has \_\_\_\_\_ protons and \_\_\_\_\_ electrons.
- How many valence electrons does the sodium atom have? \_\_\_\_\_
- How many valence electrons does sodium want? \_\_\_\_\_

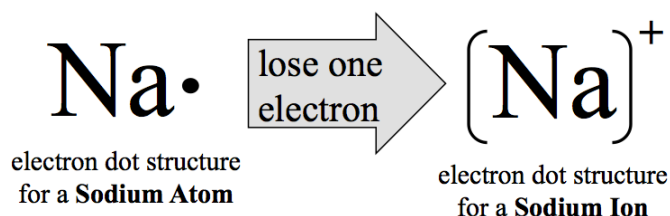
Fill the energy level diagrams with electrons:



When sodium loses an electron, it has an octet of electrons in its outer shell.

Sodium will lose \_\_\_\_\_ electron to become a sodium ion ( $\text{Na}^+$ ).

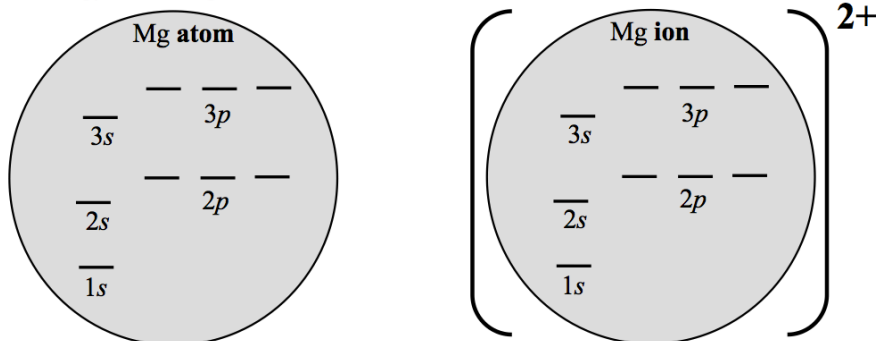
- Sodium has one valence electron
- There are two ways to have a filled octet:
  - 1) Add 7 electrons
  - 2) Remove one electron
- It is easier to remove one electron!



**Example: Let's do Another Cation - Magnesium (Mg)**

- A magnesium atom has \_\_\_\_\_ protons and \_\_\_\_\_ electrons.
- How many valence electrons does the magnesium atom have? \_\_\_\_\_
- How many valence electrons does magnesium “want?” \_\_\_\_\_

Fill the energy level diagrams with electrons:



When magnesium loses two electrons, it has an octet of electrons in its outer shell.

Magnesium will lose \_\_\_\_\_ electrons to become a magnesium ion ( $Mg^{2+}$ ).

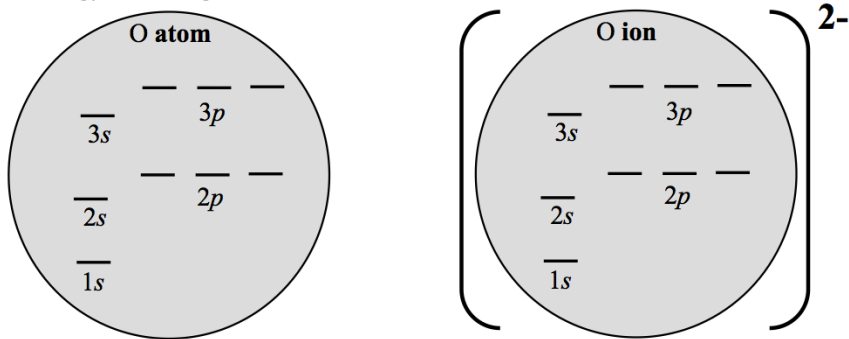
**Understanding Check:** Based on the octet rule, what would be the charge of an aluminum ion?

**HINT:** Begin with the energy level diagram (or the number of valence electrons) for an aluminum atom.

**Example: Let's do an Anion - Oxygen (O)**

- An oxygen atom has \_\_\_\_\_ protons and \_\_\_\_\_ electrons.
- How many valence electrons does the oxygen atom have? \_\_\_\_\_
- How many valence electrons does oxygen want? \_\_\_\_\_

Fill the energy level diagrams with electrons:



When oxygen gains two electrons, it has an octet of electrons in its outer shell.

Oxygen will gain \_\_\_\_\_ electrons to become an oxide ion ( $O^{2-}$ ).

The *electron dot structure* can give us the same conclusion!

Draw an electron dot structure for an **Oxygen Atom**:

gain two electrons

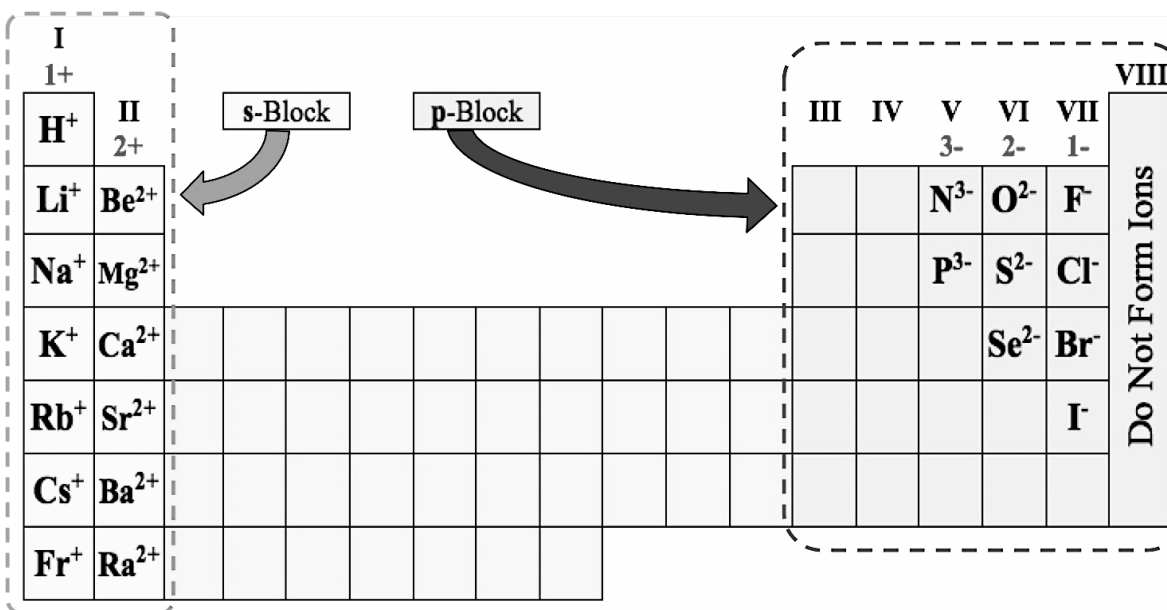
Draw an electron dot structure for an **Oxide Ion**:

Oxygen has 6 valence electrons, if we add two electrons, its outer shell will have an octet.

**Understanding Check:** What would be the charge of an **ion** formed from a **chlorine** atom?

**HINT:** Begin with the electron dot structure for a chlorine atom.

We can determine the charge of an ion formed from *s-block elements* and *p-block nonmetals* from the number of valence electrons in those elements, and therefore by their location on the periodic table.



| Periodic Group                   | Number of Valence Electrons of the Element                    | Number of Electrons Gained or Lost in Ion Formation               | Charge of Ion Formed |
|----------------------------------|---|---|----------------------|
| <b>s-Block Elements</b>          |   |   |                      |
| Group I                          | 1   | Lose 1 electron   | 1+                   |
| Group II                         | 2   | Lose 2 electrons  | 2+                   |
| <b>p-Block Nonmetal Elements</b> |   |   |                      |
| Group III                        | There are no Group III nonmetals (only metals and metalloids) |   |                      |
| Group IV                         | 4   | Do not form ions, high energy to gain <b>or</b> lose 4 electrons! |                      |
| Group V                          | 5   | Gain 3 electrons  | 3-                   |
| Group VI                         | 6   | Gain 2 electrons  | 2-                   |
| Group VII                        | 7   | Gain 1 electron   | 1-                   |
| Group VIII                       | 8   | Do not form ions, noble gas atoms have filled outer shells.       |                      |

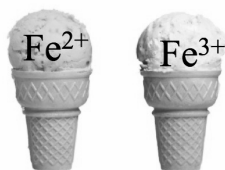
The charge of the ions formed from the **transition metals** and **p-block metals** cannot always be predicted by their position in the periodic table.

| I<br>1+         |                  |                  |   |  |  |  |  |  |  |  |  | III |  | IV | V<br>3-         | VI<br>2-        | VII<br>1-      | VIII             |  |  |  |  |                 |                 |                  |                 |
|-----------------|------------------|------------------|---|--|--|--|--|--|--|--|--|-----|--|----|-----------------|-----------------|----------------|------------------|--|--|--|--|-----------------|-----------------|------------------|-----------------|
| H <sup>+</sup>  | Li <sup>+</sup>  | Be <sup>2+</sup> | <div style="border: 1px dashed black; padding: 5px; display: inline-block;">                     transition metals                 </div> |  |  |  |  |  |  |  |  |     |  |    | N <sup>3-</sup> | O <sup>2-</sup> | F <sup>-</sup> | Do Not Form Ions |  |  |  |  |                 |                 |                  |                 |
| Na <sup>+</sup> | Mg <sup>2+</sup> |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  | P <sup>3-</sup> | S <sup>2-</sup> | Cl <sup>-</sup>  |                 |
| K <sup>+</sup>  | Ca <sup>2+</sup> |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  |                 |                 | Se <sup>2-</sup> | Br <sup>-</sup> |
| Rb <sup>+</sup> | Sr <sup>2+</sup> |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  |                 |                 |                  | I <sup>-</sup>  |
| Cs <sup>+</sup> | Ba <sup>2+</sup> |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  |                 |                 |                  |                 |
| Fr <sup>+</sup> | Ra <sup>2+</sup> |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  |                 |                 |                  |                 |
|                 |                  |                  |   |  |  |  |  |  |  |  |  |     |  |    |                 |                 |                |                  |  |  |  |  |                 |                 |                  |                 |

Many of these elements can form *more than one* type (charge) of ion.

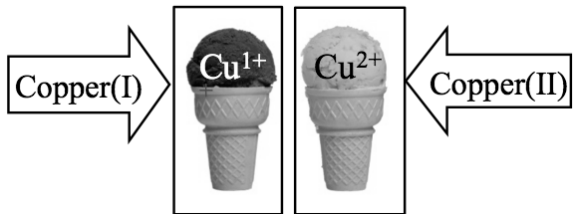
**Example: Iron (Fe):**

Iron (Fe) ions can come as Fe<sup>2+</sup> or Fe<sup>3+</sup>



**Example: Copper (Cu):**

Copper (Cu) ions can come as Cu<sup>1+</sup> or Cu<sup>2+</sup>



To differentiate the various charge states of ions when reading or writing their names, we use \_\_\_\_\_ numerals corresponding to the **charge** after the element name.

- When saying the ion's name, one would say "copper one" for Cu<sup>1+</sup> and "copper two" for Cu<sup>2+</sup>.

We *only* use the Roman numeral for ions that can *exist in more than one charge state*.

Some of the **transition metals** and **p-block metals** only exist in **one** charge state.

- For example, **cadmium ions** only exist as Cd<sup>2+</sup>.



Roman numerals *are not used* when the metal cations *have just one charge state*.

Since the charges of many of the transition metal and p-block metal ions *cannot* be easily predicted from their positions on the periodic table, and many can have more than one charge, we must refer to tabulated list for the charges.

The table below lists the charges for some transition metals and p-block ions. **You do not need to memorize the metal names and charges in this table**; I will give you this table for with your exams.

### Charges for Some Transition Metal and p-Block Metal Ions

| Ions that occur with only <u>one</u> charge  |                  |                 |                  |
|--|------------------|-----------------|------------------|
| Name   | Charge           | Name            | Charge           |
| aluminum ion                                 | Al <sup>3+</sup> | cadmium ion     | Cd <sup>2+</sup> |
| silver ion                                   | Ag <sup>+</sup>  | zinc ion        | Zn <sup>2+</sup> |
| Ions that occur with <u>multiple</u> charges |                  |                 |                  |
| Name   | Charge           | Name            | Charge           |
| copper(I) ion                                | Cu <sup>+</sup>  | tin(II) ion     | Sn <sup>2+</sup> |
| copper(II) ion                               | Cu <sup>2+</sup> | tin(IV) ion     | Sn <sup>4+</sup> |
| iron(II) ion                                 | Fe <sup>2+</sup> | lead(II) ion    | Pb <sup>2+</sup> |
| iron(III) ion                                | Fe <sup>3+</sup> | lead(IV) ion    | Pb <sup>4+</sup> |
| cobalt(II) ion                               | Co <sup>2+</sup> | mercury(I) ion  | Hg <sup>+</sup>  |
| cobalt(III) ion                              | Co <sup>3+</sup> | mercury(II) ion | Hg <sup>2+</sup> |

This table does not contain data for the all ions formed by *all* of the transition and p-block metal cations, however it includes the ions that you will need in order to solve and understand any of the examples and problems in this course.

### Naming Monatomic Ions

A *monatomic ion* is an ion that is made when a \_\_\_\_\_ atom gains or loses electron(s).

#### Naming Monatomic Cations

Cations use the name of the element, followed by the word “ion.”

- **Examples:**

Na<sup>+</sup> is referred to as a sodium ion.

Mg<sup>2+</sup> is referred to as a magnesium ion.

For monatomic cations that can occur *with multiple charges*, indicate the charge using Roman numerals *after* the element’s name.

- **Examples:**

Fe<sup>2+</sup> is referred to as an iron(II) ion

Fe<sup>3+</sup> is referred to as an iron(III) ion

#### Naming Monatomic Anions

Anions are named by changing the *suffix* (ending) of the name to “-\_\_\_\_\_.”

- **Examples:**

F<sup>-</sup> is referred to as a fluoride ion.

O<sup>2-</sup> is referred to as an oxide ion.



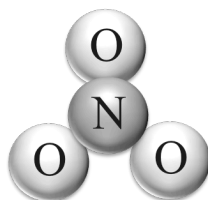
## Polyatomic Ions

Several atoms often “*stick*” (bond) together to form a small particle.

If the resulting particle has the *same number of protons as electrons*, then it will be **electrically neutral**, and we call the particle a \_\_\_\_\_.

If, on the other hand, there is an *excess of protons or an excess of electrons* in the particle, then it will have an **overall electrical charge**, and we call the particle a \_\_\_\_\_ **ion**.

### Example of a Polyatomic Ion: Nitrate Ion



|          |                          |   |                                |
|----------|--------------------------|---|--------------------------------|
| Nitrogen | (7 electrons, 7 protons) | } | Nitrate Ion<br>$\text{NO}_3^-$ |
| Oxygen   | (8 electrons, 8 protons) |   |                                |
| Oxygen   | (8 electrons, 8 protons) |   |                                |
| Oxygen   | (8 electrons, 8 protons) |   |                                |
|          | + one extra electron     |   |                                |

The table below lists the names and charges for some polyatomic ions. **You do not need to memorize this table**; I will give you this table for with your exams.

**Some Polyatomic Ion Names and Charges**

| POLYATOMIC CATIONS  |   |
|---|---|
| $\text{H}_3\text{O}^+$ hydronium ion                              | $\text{NH}_4^+$ ammonium ion  |
| POLYATOMIC ANIONS   |   |
| $\text{OH}^-$ hydroxide ion                                       | $\text{HSO}_4^-$ hydrogen sulfate (or bisulfate) ion  |
| $\text{CO}_3^{2-}$ carbonate ion                                  | $\text{PO}_4^{3-}$ phosphate ion  |
| $\text{HCO}_3^-$ bicarbonate (also called hydrogen carbonate) ion | $\text{HPO}_4^{2-}$ hydrogen phosphate ion  |
| $\text{NO}_2^-$ nitrite ion                                       | $\text{H}_2\text{PO}_4^-$ dihydrogen phosphate ion  |
| $\text{NO}_3^-$ nitrate ion                                       | $\text{CrO}_4^{2-}$ chromate ion  |
| $\text{SO}_3^{2-}$ sulfite ion                                    | $\text{Cr}_2\text{O}_7^{2-}$ dichromate ion   |
| $\text{SO}_4^{2-}$ sulfate ion                                    | $\text{C}_2\text{H}_3\text{O}_2^-$ acetate ion (sometimes written as $\text{CH}_3\text{CO}_2^-$ ) |
|   | $\text{CN}^-$ cyanide ion   |

# Chemical Compounds

Compounds: \_\_\_\_\_

---

Each compound has the same \_\_\_\_\_ of the same elements.

- Example: Water = 2 hydrogen atoms and 1 oxygen atom (Ratio H:O = 2:1)

## Chemical bonds

Atoms can **bond** with other atoms, and ions can **bond** with other ions to form **compounds** such as water (H<sub>2</sub>O), carbon dioxide (CO<sub>2</sub>), and table salt (sodium chloride).

**Chemical bonds** are the **electrical attractive** \_\_\_\_\_ that hold atoms or ions together in a compound.

There are *three types* of **chemical bonding**:

- 1) **Covalent Bonding**
- 2) **Ionic Bonding**
- 3) **Metallic Bonding**

In this chapter, you will learn about the first two types, *covalent bonding* and *ionic bonding*. You will learn about *metallic bonding* in chapter 5.

## Some Terminology

All matter can be classified as either **mixtures** or **pure substances**. You will learn about mixtures in chapter 6. **Pure substances** are described in the chart *on the right*:

**Chemistry** is the study of matter and the **changes** it undergoes.

\_\_\_\_\_ **changes**, such as *melting* or *boiling*, result in changes in *physical properties* and **do not** involve the formation of new *pure substances*.

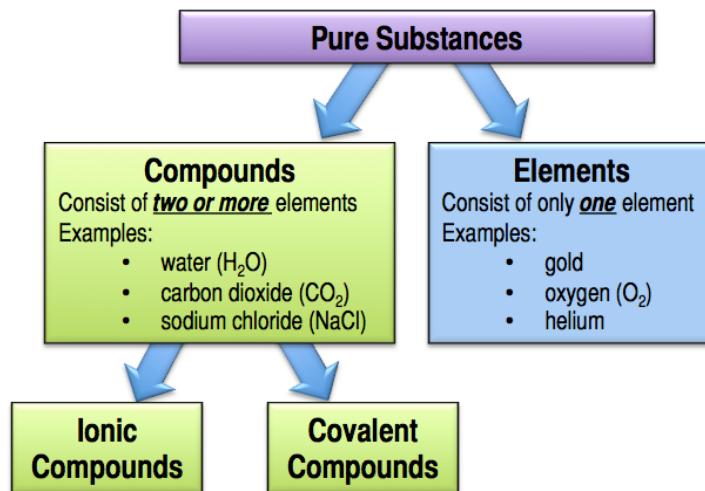
- For example, the melting of ice is simply H<sub>2</sub>O being changed from the *solid* phase to the *liquid* phase. The chemical bonds between oxygen and hydrogen atoms do not change in that process.

\_\_\_\_\_ **changes**, on the other hand, result in the formation of **new pure substances**.

- To make a **new** pure substance, **chemical bonds must be** \_\_\_\_\_ **and/or new chemical bonds are** \_\_\_\_\_.
- This happens in a process called a **chemical reaction**, which we will study in chapter 6.

A **major principle** of chemistry is that the observed (macroscopic) properties of a substance are related to its “microscopic” structure.

- The *microscopic* structure entails details such as the kind of atoms/ions and the pattern in which they are **bonded** to each other.



## Covalent Chemical Bonding

**Covalent bonding** is defined as the chemical bonding *force* that results from the \_\_\_\_\_ between two atoms.

The resulting collection of atoms results in the formation of either \_\_\_\_\_ or *polyatomic ions*.

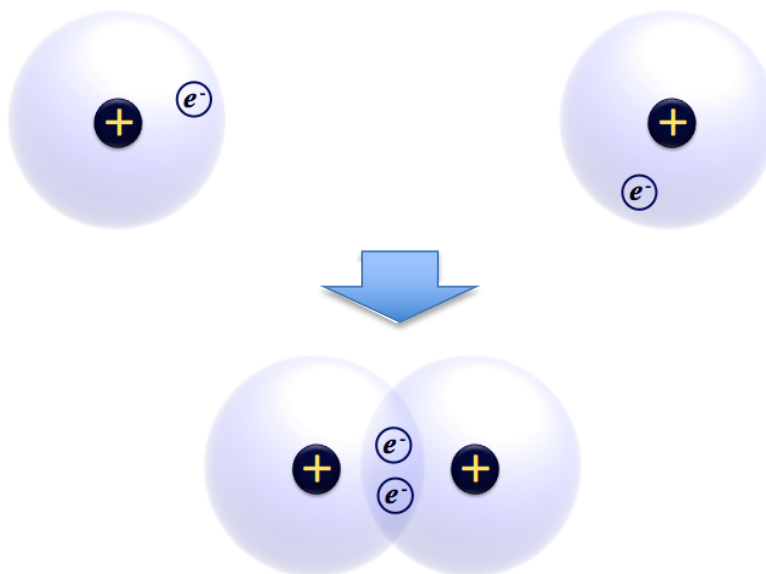
A **molecule** is an electrically \_\_\_\_\_ group of atoms *held together by covalent bonds*.

Covalent bonding occurs between \_\_\_\_\_ atoms.

Covalent bonding occurs because the bound atoms are at a \_\_\_\_\_ energy than the unbound atoms.

**Why does sharing of electron pairs result in an attractive electrostatic force capable of holding atoms together?**

Consider the two *hydrogen atoms* coming together to form a covalent bond.



In covalent bonding, the atoms \_\_\_\_\_ *electron pairs*.

The shared electron pair spends significantly *more* \_\_\_\_\_ in the area between the positive nuclei of the hydrogen atoms than in other regions.

The electron pair between the nuclei create a **positive-negative-positive** electrostatic attractive “sandwich” and this \_\_\_\_\_ holds the atoms together.

### The Octet Rule in the Formation of Molecules

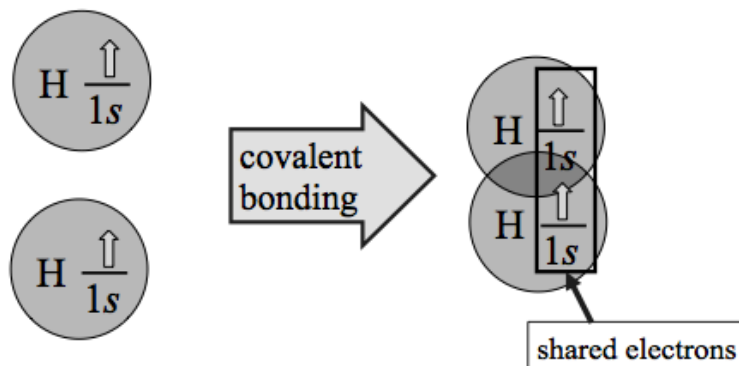
The positive-negative-positive model **cannot** explain why a covalent bond *does not* form between two *helium* atoms.

The octet rule in the formation of molecules is: *molecules tend to form such that the atoms are surrounded by an octet (eight) of valence electrons (except for hydrogen and helium that have two electrons)*.

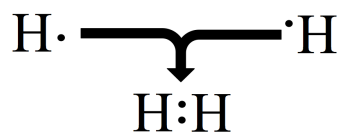
**Example: H<sub>2</sub>**

(recall that H and He are stable with two valence electrons)

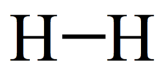
When a covalent bond forms, each hydrogen atom “feels” **two** electrons in its outermost shell.



The H<sub>2</sub> covalent bond can also be illustrated with electron dot structures.



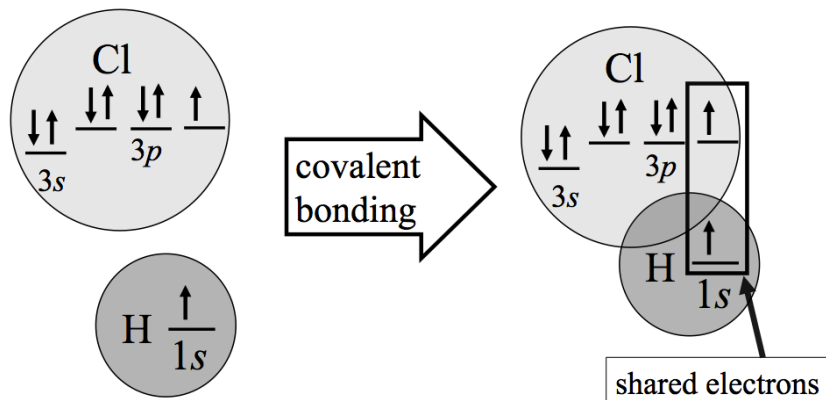
The two electrons \_\_\_\_\_ the atoms are shared in a covalent bond.  
Chemist use a line to represent \_\_\_\_\_ electrons in a covalent bond.



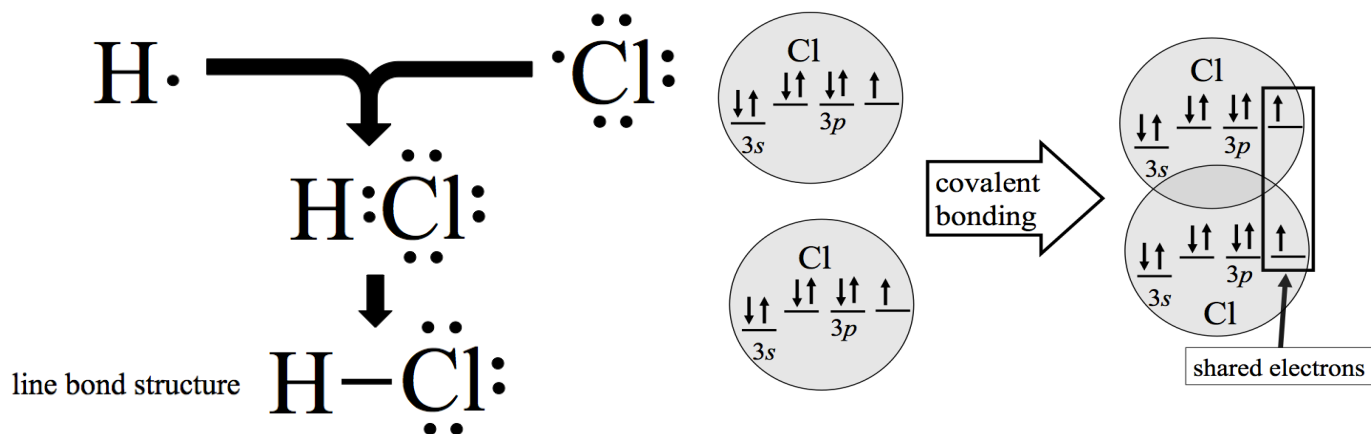
These drawings are called \_\_\_\_\_.

Let's do another example: **Hydrogen Chloride (HCl)**

When a covalent bond forms, the hydrogen atom “feels” **two** electrons in its outermost shell, and the chlorine atom “feels” **eight** electrons in its outermost shell.

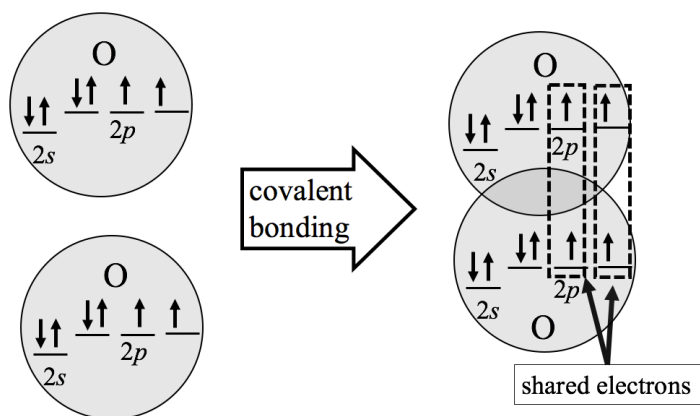


The HCl covalent bond can also be illustrated using electron dot structures.



Let's do another example:  $\text{Cl}_2$  (chlorine gas).

When a covalent bond forms, each chlorine atom "feels" **eight** electrons in its outermost shell.



**You try it:**

Draw the *line bond structure* for  $\text{Cl}_2$ .

- Start with the electron dot structure for two Cl atoms.

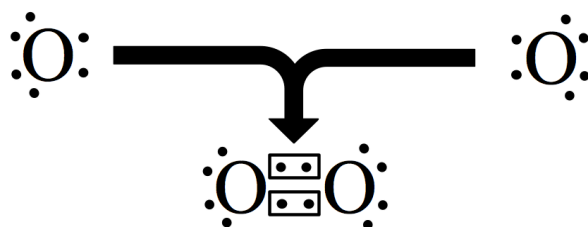
Let's do **oxygen gas ( $\text{O}_2$ )**.

In  $\text{O}_2$ , *two pairs* of electrons are shared.

When a covalent bond forms, each oxygen atom "feels" **eight** electrons in its outermost shell.

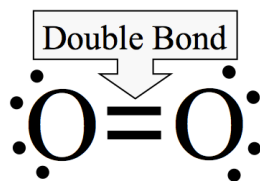
Let's draw the line bond structure for **oxygen gas ( $\text{O}_2$ )**.

- Oxygen atoms have 6 valence electrons.
- We will rotate the electrons so they can form bonding pairs.



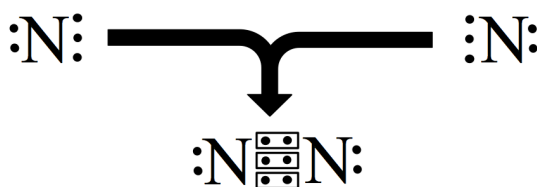
We use lines to represent \_\_\_\_\_ electron pairs.

When atoms are bonded with 2 pairs of electrons it is called a \_\_\_\_\_ .



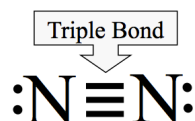
Let's draw the line bond structure for nitrogen gas (N<sub>2</sub>)

- Nitrogen atoms have 5 valence electrons.
- We will rotate the electrons so they can form bonding pairs.



We use lines to represent electron pairs.

When atoms are bonded with **3 pairs** of electrons it is called a \_\_\_\_\_.



## Naming Covalent Compounds

The covalent bonding that we will see in this course will *always* involve \_\_\_\_\_ **elements only**.

The nonmetal atoms can share electrons to form **molecules** (*molecular compounds*) or **polyatomic ions**.

A chemical substance whose simplest units are **molecules** is called a \_\_\_\_\_ **compound**.

When discussing molecules we use a \_\_\_\_\_ that shows the *types* (*elements*) and *numbers of atoms* that make up a single molecule.

The number of atoms of each element contained in the molecule is written as a **subscript** after the element's symbol.

- **Examples:**

| line bond structure                 | molecular formula    |
|-------------------------------------|----------------------|
| $\text{H}-\text{H}$                 | $\text{H}_2$         |
| $\text{H}-\ddot{\text{O}}-\text{H}$ | $\text{H}_2\text{O}$ |

When there is only **one atom** of a particular element present in a molecule the subscripted "1" is **omitted** for that element.

Some molecules **only** contain **one** element, for example **H<sub>2</sub>**, **Cl<sub>2</sub>**, and **O<sub>2</sub>**.

- These molecules often take the name of the elements they contain.
- **Examples:**

| molecular formula | name     |
|-------------------|----------|
| $\text{H}_2$      | hydrogen |
| $\text{O}_2$      | oxygen   |

## Naming Binary Covalent (Molecular) Compounds

\_\_\_\_\_ covalent compounds contain only *two* \_\_\_\_\_ (the “bi-” prefix indicates “two”).

- Examples of binary covalent compounds are  $\text{HCl}$ ,  $\text{H}_2\text{O}$ , and  $\text{CO}_2$ .

### Educational Goals:

Given the **name** of a *binary covalent molecule*, be able to write the **molecular formula**.

Given the **molecular formula** of a *binary covalent molecule*, be able to write the **name** of the molecule.

### Method for Naming Binary Covalent (Molecular) Compounds

1. List the name of the first element in the formula.
2. List the second element and add the **-ide** \_\_\_\_\_.
3. Use Greek \_\_\_\_\_ to indicate the number of each atom in the formula.

- **Exception:** If there is just one atom of the \_\_\_\_\_ element in the formula, do not use **mono-** for the *first element in the name*.

- **Example:**  $\text{CO}_2$   
~~mon~~carbon dioxide → carbon dioxide

- The "o" or "a" at the *end of the Greek prefix* is omitted when the element's name begins with a vowel.

- **Example:**  $\text{CO}$   
carbon monoxide → carbon monoxide

| Greek Prefix | Number |
|--------------|--------|
| mono         | 1      |
| di           | 2      |
| tri          | 3      |
| tetra        | 4      |
| penta        | 5      |
| hexa         | 6      |
| hepta        | 7      |
| octa         | 8      |
| nona         | 9      |
| deca         | 10     |

### Example Problem:

Name the following compound:  $\text{CCl}_4$

**Answer:** carbon tetrachloride

### Understanding Check

Write the **names** of the following molecules:

$\text{CF}_4$  \_\_\_\_\_

$\text{N}_2\text{O}$  \_\_\_\_\_

$\text{SF}_6$  \_\_\_\_\_

### Method for Writing the Molecular Formula of a Binary Covalent Compound

1. Write the symbol of the *first element* in the compound's name, then the symbol of the *second element* in the compound's name.
2. Indicate *how many atoms* of each element the molecule contains using **subscripts** after the atomic symbol.

- The *numbers of atoms* are given in the Greek prefixes in the molecule's name.
- NOTE: If there is no Greek prefix in front of the first element in the name, that means the number is **1**.

### Example:

Write the *molecular formula* for **dinitrogen tetrafluoride**.





## Understanding Check

Write the **molecular formula** for the covalent compounds:

- nitrogen trichloride \_\_\_\_\_
- dinitrogen pentoxide \_\_\_\_\_
- sulfur dioxide \_\_\_\_\_

For covalent compounds with *more than two types of atoms*, we use *common names* or IUPAC system names. You are not responsible for knowing *common names*. You will learn some IUPAC system names in later chapters.

Examples of *common names*:

- Glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>)
- Acetone (C<sub>3</sub>H<sub>6</sub>O)

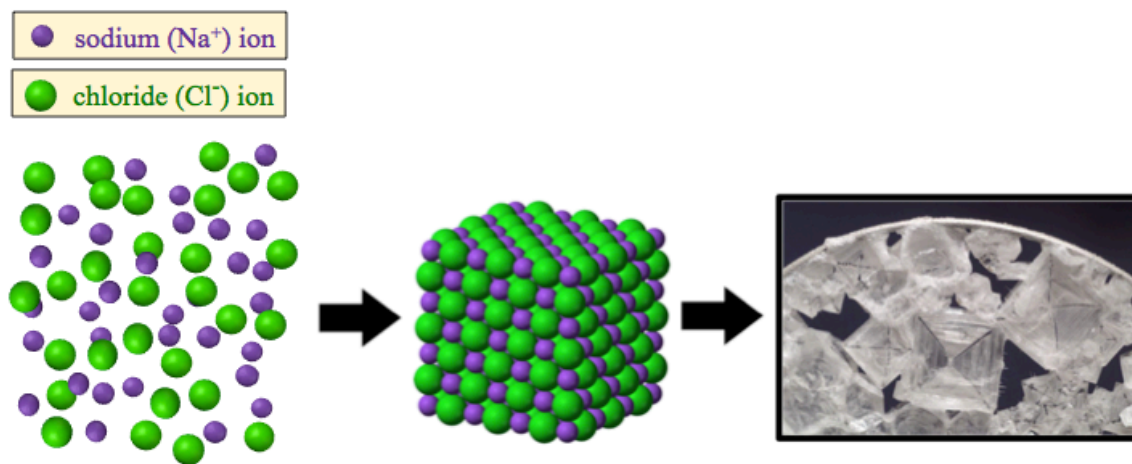
## Ionic Compounds

Definition of **ionic bonding**: Chemical bonding that results from the electrostatic attraction between \_\_\_\_\_ **numbers** of *cations* and *anions*.

- Compounds composed of ions are called **ionic compounds**.

**Example of an ionic compound: sodium chloride (NaCl)**

**Many** sodium ions combine with **many** chloride ions in a *three-dimensional pattern* that minimizes the distance between the oppositely charged cations and anions and maximizes the distance between the like-charged particles.



We call this structure a \_\_\_\_\_ or *crystal* \_\_\_\_\_.

It is this regular, repeating structure on the scale of the individual ions that give crystals the interesting geometrical shapes that we see on the macro-scale when we look at them with our eyes or with a microscope.

**Ionic bonding (ionic compounds)** results from:

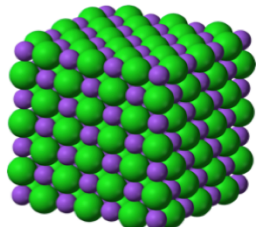
- Combining **metal ions** with **nonmetal** ions.
- Combining **polyatomic ions** with *other* ions.

The cations and anions will combine in a ratio such that the **total** of the *positive* (+) and *negative* (−) charges equals \_\_\_\_\_!

- **Example:** Sodium Chloride (NaCl)

● sodium (Na<sup>+</sup>) ion

● chloride (Cl<sup>−</sup>) ion



Sodium ions have a charge of 1+

Chloride ions have a charge of 1-

They combine in a 1-to-1 \_\_\_\_\_ in the crystal

For every sodium ion, there is one chloride ion!

The charges \_\_\_\_\_ up to ZERO!

### Formula Units

The use of *molecular formulas* would not make sense for ionic compounds; they do not form molecules, instead they form crystals.

We write \_\_\_\_\_ (as opposed to *molecular formulas*) for **ionic compounds**.

The *formula unit* looks like the molecular formula used for covalent compounds, however it means something *entirely* different.

The *formula unit* uses **subscripted numbers** after the ion's symbol that indicate the **ratio** that the cations and anions combine in the ionic crystal.

- As in the case of molecular formula, when a subscript would have a value of “1,” the subscript is omitted.
- We write the cation symbol first followed by a numerical subscript (if needed), then we write the anion symbol followed by a numerical subscript (if needed).

**Example:** For sodium chloride, since sodium ions and chloride ions combine in a **one-to-one ratio**, we write the formula unit of sodium chloride as:



**Example:** *Calcium* ions combine with *fluoride* ions to form an **ionic compound**.



Calcium ions have a charge of 2+

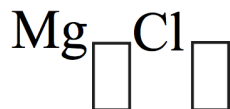
Fluoride ions have a charge of 1-

They combine in a \_\_\_\_\_ ratio in the crystal

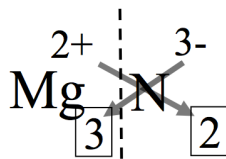
For *every* calcium ion, there are \_\_\_\_\_ fluoride ions.

We write the **formula unit** for calcium fluoride as:



**Understanding Check:**Write the *formula unit* for the compound formed by combining magnesium and chloride ions.**Understanding Check:**Write the *formula unit* for the compound formed by combining potassium and oxide ions.**Understanding Check:**Write the *formula unit* for the compound formed by combining magnesium and nitride ions.

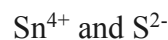
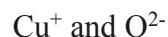
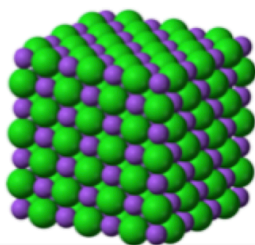
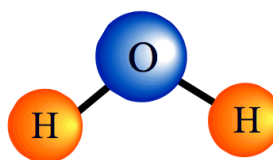
Dr. Zoval's Caveman Style, Works Every Time Method:



The Criss-Cross Method

**Formula Units**

Write the formula for the ionic compound formed between each of the following pairs of ions:

**Formula Unit vs. Molecular Formula****Formula Unit** =  
Lowest **RATIO** of ions**Example: NaCl**Ratio of Na<sup>+</sup> to Cl<sup>-</sup> = 1 to 1● sodium (Na<sup>+</sup>) ion● chloride (Cl<sup>-</sup>) ion**Molecular Formula** =  
Actual **number** of atoms**Example: H<sub>2</sub>O***two* hydrogen atoms and  
*one* oxygen atom

## Naming Ionic Compounds

### Educational Goals:

Given the **name** of an *ionic compound*, be able to write the **formula unit**.

Given the **formula unit** of an *ionic compound*, be able to write the **name**.

### Method for Writing Formula Units for Ionic Compounds

1) Write the symbol of the first ion (the cation) in the compound's name, then the symbol of the second ion (the anion) in the compound's name.

2) Indicate the **ratio** of the ions in the compound using *subscripts* after each ion.

The ratio of the ions is deduced by *balancing the charges* of the ions so that the total charge in the crystal is equal to **zero**.

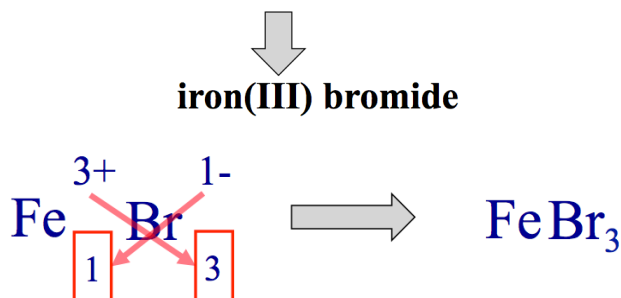
- We find the ion's charge from its position on the periodic table or, for polyatomic ions, we look it up in a table.
- You will know the charge for the metals *that occur with various charges* because the charge will be written in the compound's name in Roman numerals.

For \_\_\_\_\_ *ions*:

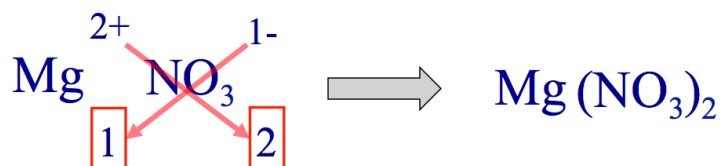
When the *subscript* for a **polyatomic ion** is *greater than 1*, the polyatomic ion formula is written in parenthesis and the subscript is written after/outside of the parenthesis.

**Example:** Write the *formula unit* for **iron(III) bromide**.

- You will know the **charge** for the metals *that occur with various charges* because the charge will be written in the compound's name in **Roman numerals**.



**Example:** Write the *formula unit* for **magnesium nitrate**.



- **For polyatomic ions:**  
When the *subscript* for a **polyatomic ion** is *greater than 1*, the polyatomic ion formula is written in parenthesis and the subscript is written after/outside of the parenthesis.

**Understanding Check:** Write the *formula unit* for each of the following compounds:

- sodium bicarbonate
- sodium fluoride
- iron(III) chloride
- sodium carbonate
- copper(II) sulfate
- magnesium hydroxide

### Method for Writing the Names of Ionic Compounds

- Write the \_\_\_\_\_ name first, then the \_\_\_\_\_ name.
  - Monoatomic **anions** (anions composed of one element) use the “ide” suffix.
  - We get the names of *polyatomic ions* from the polyatomic ion table.
- If the cation is one of the *metals with various charges*, write the charge using parenthesis and Roman numerals after the metal’s name.

**Example:** Name the following compound: **MgCl<sub>2</sub>**

Name: **magnesium chloride**

**Example:** Name the following compound: **CuBr<sub>2</sub>**

- What *must* the charge of the copper ion be? **2+**

Name: **copper(II) bromide**

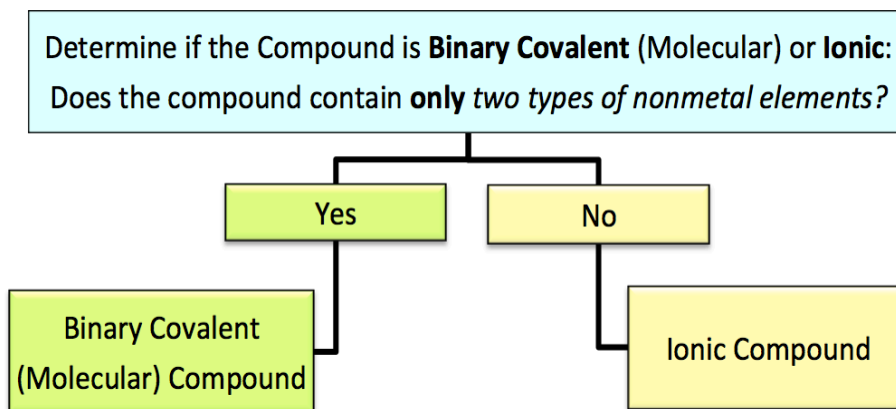
Complete the names of the following ionic compounds with variable charge metal ions:

|                                |                       |
|--------------------------------|-----------------------|
| FeBr <sub>2</sub>              | iron( __ ) bromide    |
| CuCl                           | copper( __ ) chloride |
| SnO <sub>2</sub>               | _____ ( __ ) _____    |
| Fe <sub>2</sub> O <sub>3</sub> | _____                 |

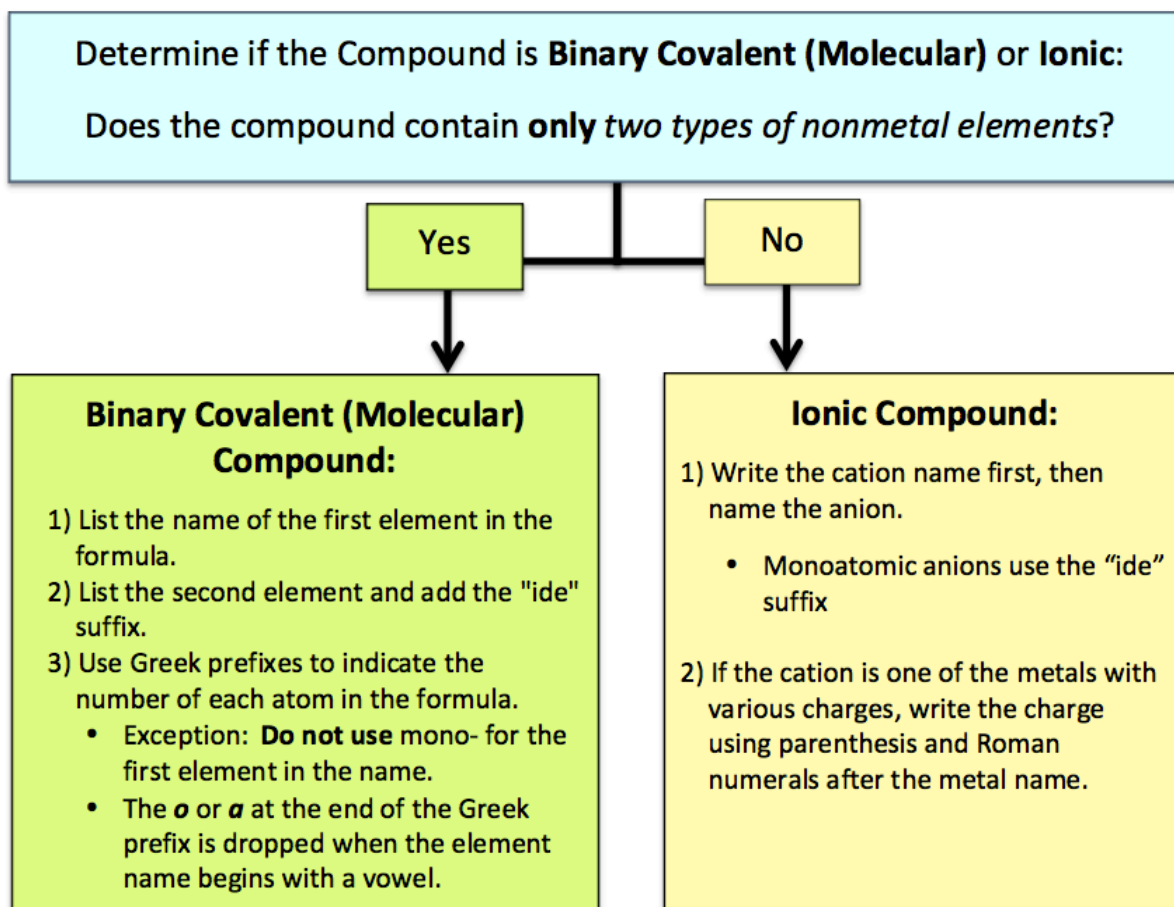
Name the following ionic compounds:

|                                |       |
|--------------------------------|-------|
| NaCl                           | _____ |
| ZnI <sub>2</sub>               | _____ |
| Al <sub>2</sub> O <sub>3</sub> | _____ |

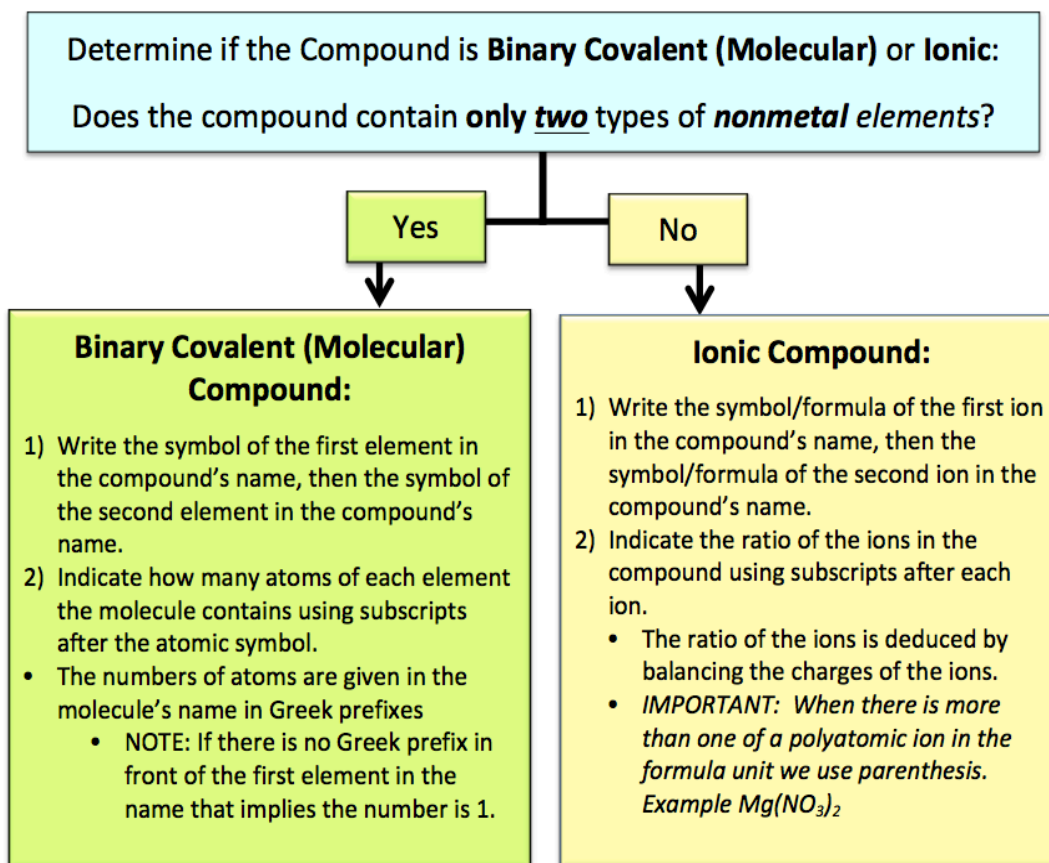
## Naming Compound Summary



### Given the Molecular Formula, Write the Name



## Given the Name, Write the Molecular Formula



## Molar Mass of Compounds

In this video, you will learn how to calculate the **molar mass** of a compound and how to use the molar mass of a compound to do *mole-mass conversions*.

- 1) Molar Mass of **Covalent Compounds (Molecules)**
- 2) Molar Mass of **Ionic Compounds**

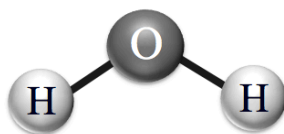
### Molar Mass of Covalent Compounds (Molecules)

The **molar mass** of a \_\_\_\_\_ tells us the mass (grams) of **1 mole** of the *molecules*.

- The *molar mass* of a molecule is also called the **molecular mass**.

To calculate the *molar mass* of a **molecule** we **add up** the *atomic molar masses* of all \_\_\_\_\_ in the molecule.

**Example:** Let's calculate the molar mass of H<sub>2</sub>O.



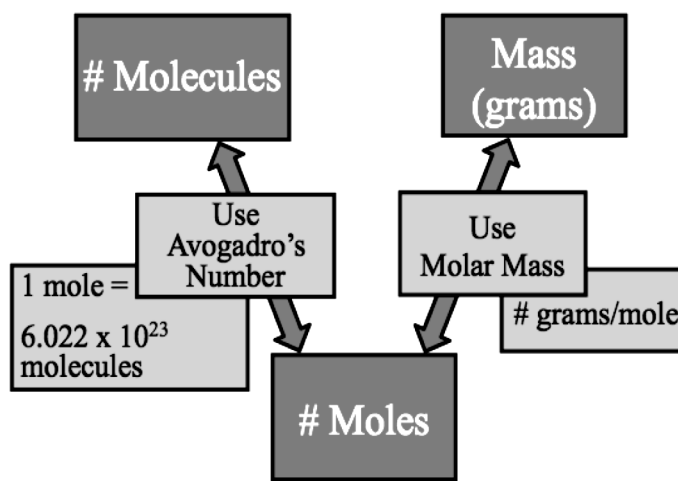
| Atom                                  | # of Atoms | Atomic Molar Mass | Total               |
|---------------------------------------|------------|-------------------|---------------------|
| oxygen                                | 1          | X 16.00 g/mole    | 16.00 g/mole        |
| hydrogen                              | 2          | X 1.01 g/mole     | 2.02 g/mole         |
| <b>Molar Mass of H<sub>2</sub>O =</b> |            |                   | <b>18.02 g/mole</b> |

One mole of H<sub>2</sub>O  
(6.022 x 10<sup>23</sup> molecules) } has a mass of 18.02 grams

**Understanding Check:** Calculate the molar mass of CH<sub>4</sub> (methane).

### Mass-Mole-Molecules Conversions

Note that, as in the case of atoms, the molar mass of a compound is the *relationship* between *moles* and *mass (grams)*, therefore we can **convert** between moles and grams of compounds.





**Example:** How many *grams* of CH<sub>4</sub> is contained in 3.65 *moles*?

Use the molar mass to write an *equivalence statement*:

- **1 mole CH<sub>4</sub> = 16.05 grams**

The equivalence statements can be written as conversion factors:

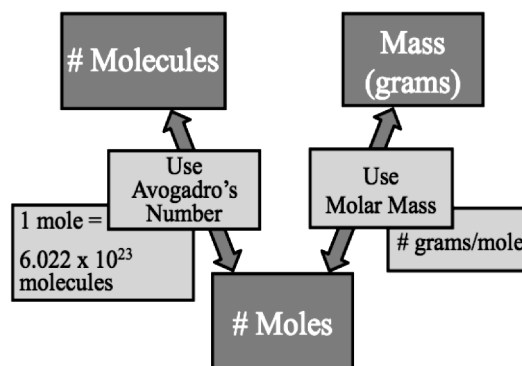


$$\frac{3.65 \text{ moles CH}_4}{1} \times \frac{16.05 \text{ grams CH}_4}{1 \text{ mole CH}_4} = 58.6 \text{ grams CH}_4$$

You have just learned how to convert between moles and mass of a compound and vice versa.

We do a **two-step calculation** to convert between **mass** and number of **molecules**.

We can *convert between molecules and moles* since **Avogadro's Number** applies to **molecules**; one mole of a molecular compound contains  $6.022 \times 10^{23}$  molecules.



**You try one:** How many H<sub>2</sub>O *molecules* are contained in 237 grams?

## Molar Mass of Ionic Compounds

When using the **molar mass** of *ionic compounds*, we calculate the mass of a compound based on the number of each ion as it appears in the formula unit.

- For this reason, the *molar mass of an ionic compound* is also called \_\_\_\_\_ **mass**.

**Example:** The molar mass of sodium chloride (NaCl)

The **formula unit** for *sodium chloride* is **NaCl** because there is a 1:1 ratio of sodium ions to chloride ions in the crystal.

One **mole** of sodium chloride contains **one mole** of *sodium ions* and **one mole** of *chloride ions*.

Although **ions** have *extra* or *missing* electrons, their molar masses are calculated by adding the *atomic molar masses* of the elements they contain.

- The reason we can do this is because the mass of electrons is negligible compared to the mass of protons and neutrons.

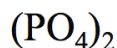
| Ion                                      | # of ions in the Formula Unit | Molar Mass of ion | Total                 |
|--|-------------------------------|-------------------|-----------------------|
| Sodium                                   | 1                             | x 22.99 g/mole    | = 22.99 g/mole        |
| Chloride                                 | 1                             | x 35.45 g/mole    | = 35.45 g/mole        |
| <b>Molar Mass (Formula Mass) of NaCl</b> |                               |                   | <b>= 58.44 g/mole</b> |

**Example:** What is the molar mass of iron(II) phosphate,  $\text{Fe}_3(\text{PO}_4)_2$  ?

One mole of iron(II) phosphate contains **three** moles of *iron(II) ions* and **two** moles of *phosphate ions*.

**three** moles of *iron(II) ions*

**two** moles of *phosphate ions*



The molar mass of **each** *iron(II)* ion is: 55.85 g/mole.

**each** phosphate ion contains:

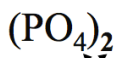
- one mole of phosphorus
- four moles of oxygen

The molar mass of **each** *phosphate ion* is: **94.97 g/mole**.

The molar mass (or formula mass) is calculated by adding the molar masses of the ions:

| Ion   | # of ions in the Formula Unit | Molar Mass of ion   | Total                  |
|---|-------------------------------|---|------------------------|
| Iron(II)  | 3                             | x 55.85 g/mole  | = 167.55 g/mole        |
| Phosphate   | 2                             | x 94.97 g/mole<br>based on:<br>one phosphorus <b>and</b><br>four oxygens <b>per ion</b> | = 189.94 g/mole        |
| <b>Molar Mass (Formula Mass) of <math>\text{Fe}_3(\text{PO}_4)_2</math></b> |                               |   | <b>= 357.49 g/mole</b> |

**An Alternative Method:**



**three** moles of *iron(II) ions*

**two** moles of *phosphate ions*  
contain:

- two** moles of phosphorous
- eight** ( $2 \times 4$ ) moles of oxygen

Three moles of **Fe**:  $3 \times 55.85 \text{ g/mole} = \mathbf{167.55 \text{ g/mole}}$

Two moles of **P**:  $2 \times 30.97 \text{ g/mole} = \mathbf{061.94 \text{ g/mole}}$

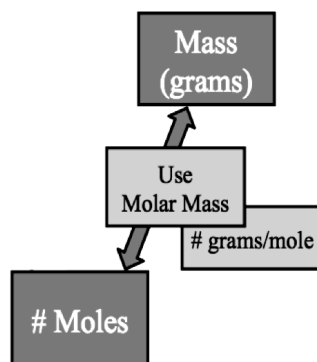
Eight moles of **O**:  $8 \times 16.00 \text{ g/mole} = \mathbf{128.00 \text{ g/mole}}$

The molar mass of  $\text{Fe}_3(\text{PO}_4)_2$  is  $\mathbf{357.49 \text{ g/mole}}$

**Understanding Check:** What is the molar mass of magnesium nitrate,  $\text{Mg}(\text{NO}_3)_2$  ?

### Mole-Mass Conversions for Ionic Compounds

Mole-Mass conversions for ionic compounds are done *exactly* as we did for covalent compounds; use the molar mass as a conversion factor.



**You Try One:** What is the **mass** (grams) of 4.95 **moles** of  $\text{Mg}(\text{NO}_3)_2$  ?