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.

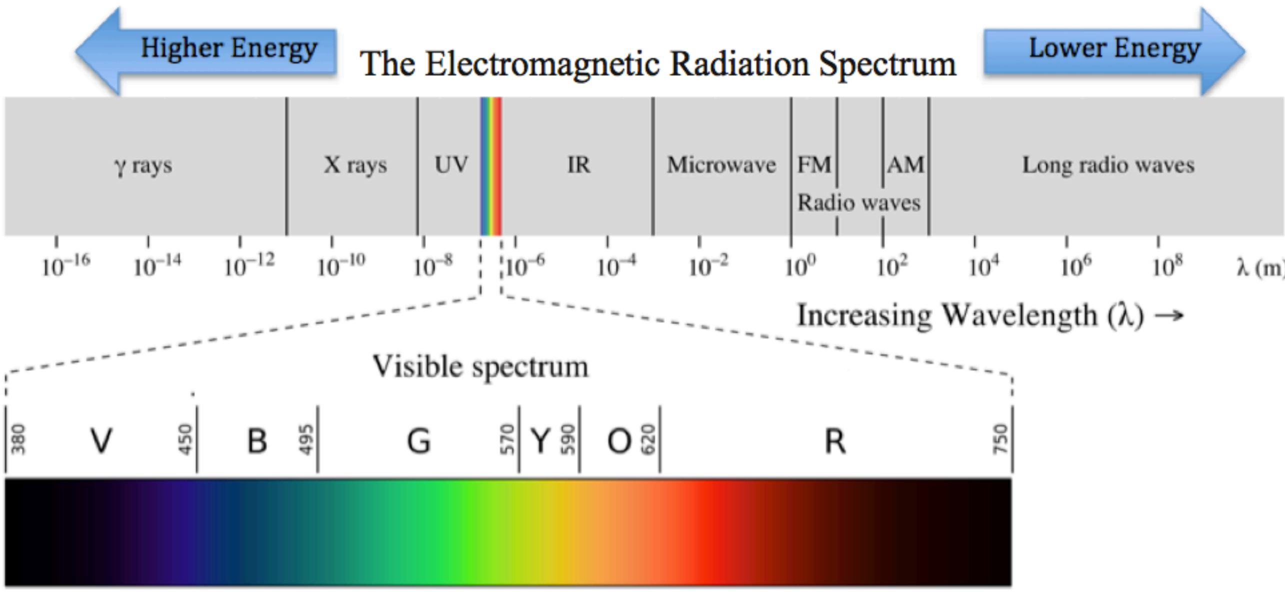
Atoms release energy when electrons move *back* to ________ ____________areas.

- 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**.

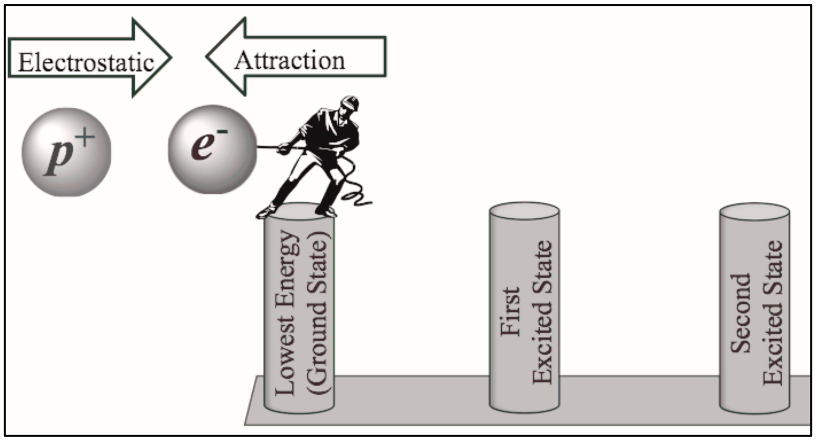
- Another word for "***discrete***" is "***distinct***."

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.

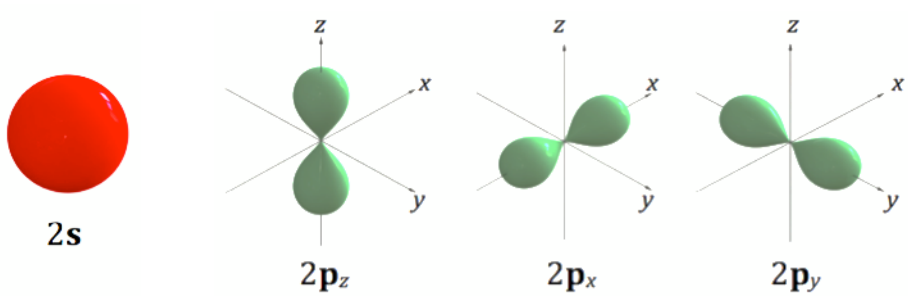
- It is called an _____ orbital.
- **"s"** represents the ***shape* of the orbital**, we use **1s** because **n**=**1**).
- **s** orbitals are ______________ in shape.

An Illustration of a 1s Orbital

- The ___________ is in the **center** of the orbital.

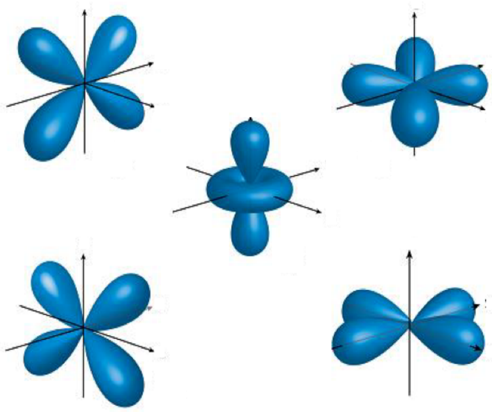
The **n**=**2** quantum level has _______ orbitals.

- 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** orbital are illustrated below:



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

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 then 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.

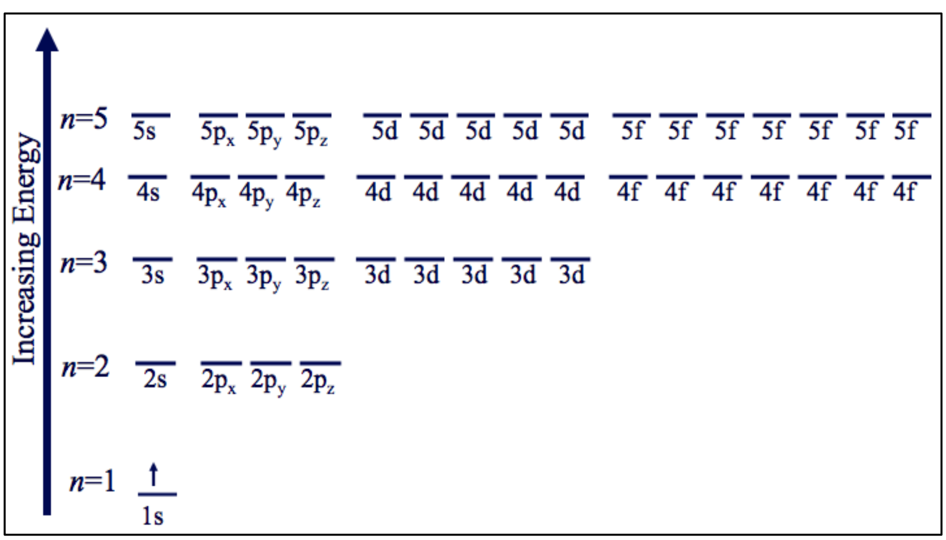
Although quantum levels with **n** > **4** contain orbitals other than **s**, **p**, **d**, and **f**, these other orbitals are n*ever occupied by electrons of any element in its ground state*.

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

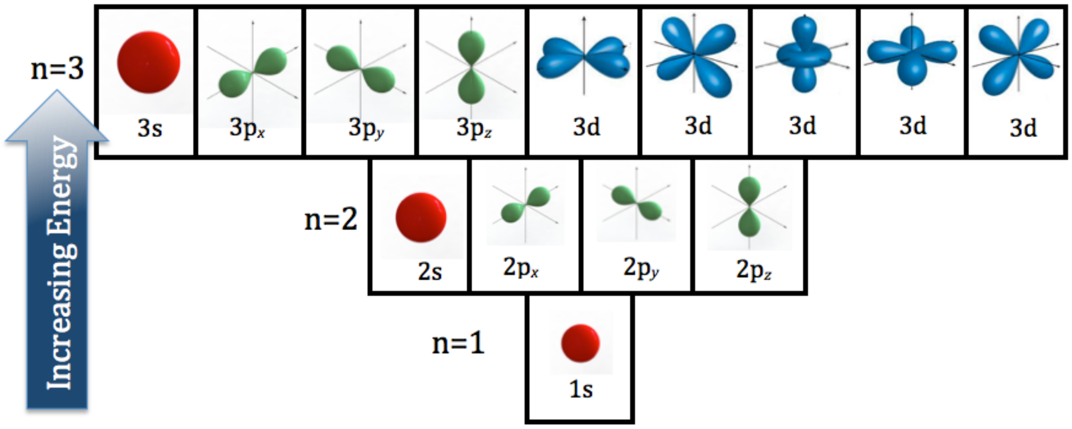
Energy Level Diagram for Hydrogen

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 diagra**



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

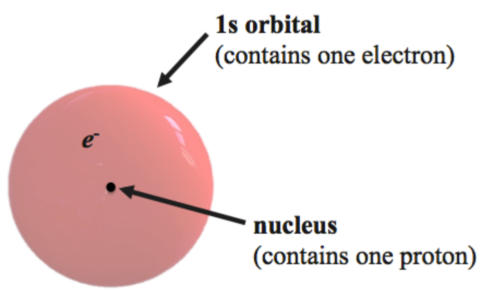


The different ____________ of the skyscraper represent the quantum levels (**n**).

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

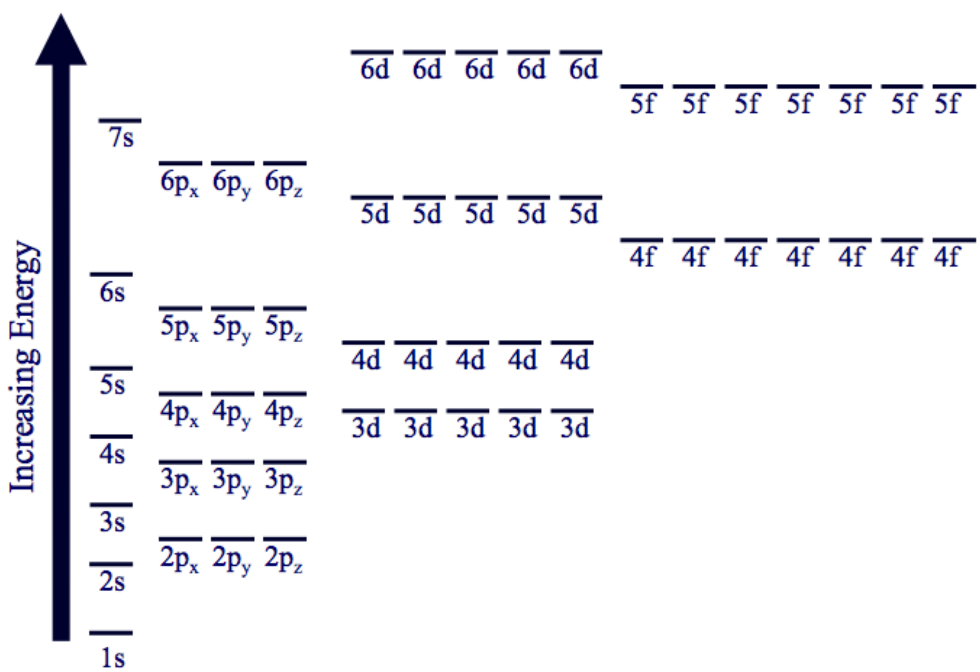
___________ on a particular floor are analogous to the various orbitals in a particular *quantum level*.

Model of a Hydrogen Atom

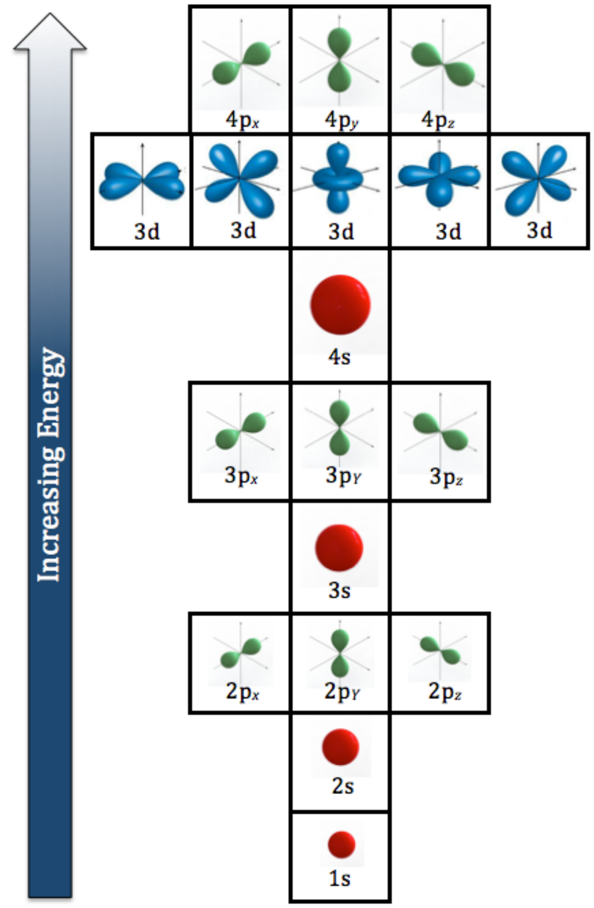


Atomic Model for Multi-Electron Atoms

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**.

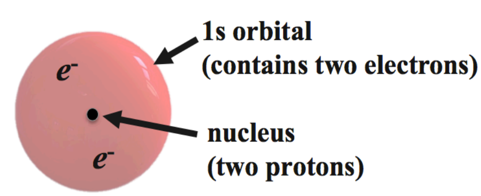
We call the spin states "**up**" or "**down**."

- *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!

Example:
Energy Level Diagram for a Helium Atom

Model of a Helium Atom



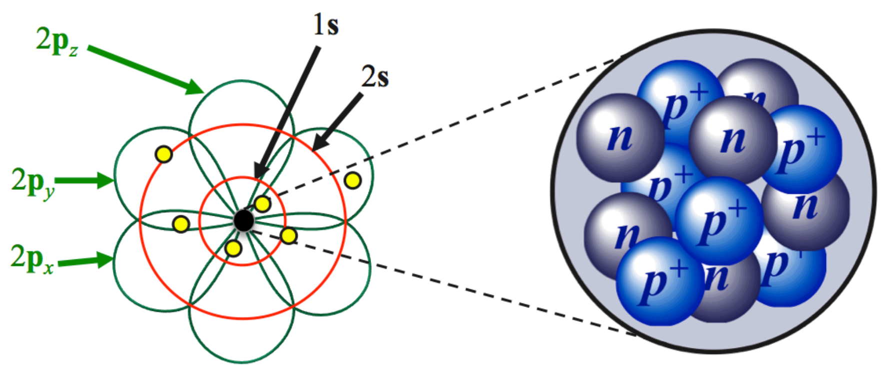
3) Hund's Rules

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

Drawing of a Carbon-12 Atom



Nucleus

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

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

Valence Electrons

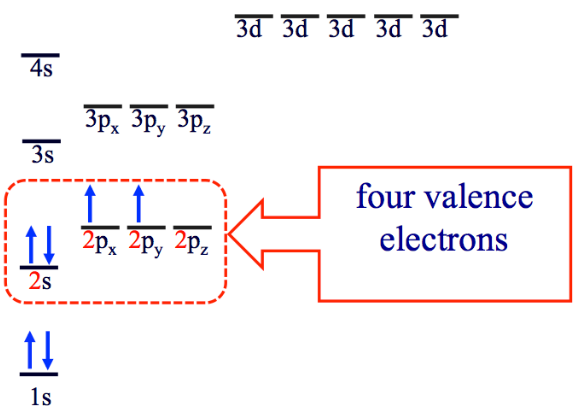
Valence electrons are the electrons held in the ________________________ ***shell*** (largest "**n**").

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.

Transition Metals

I	II											III	IV	V	VI	VII	VIII
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt									

(Inner) Transition Metals

Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

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_2O); sulfur atoms, also often "bond" with two hydrogen atoms to form hydrogen sulfide (H_2S).

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

a. hydrogen (H)

b. nitrogen (N)

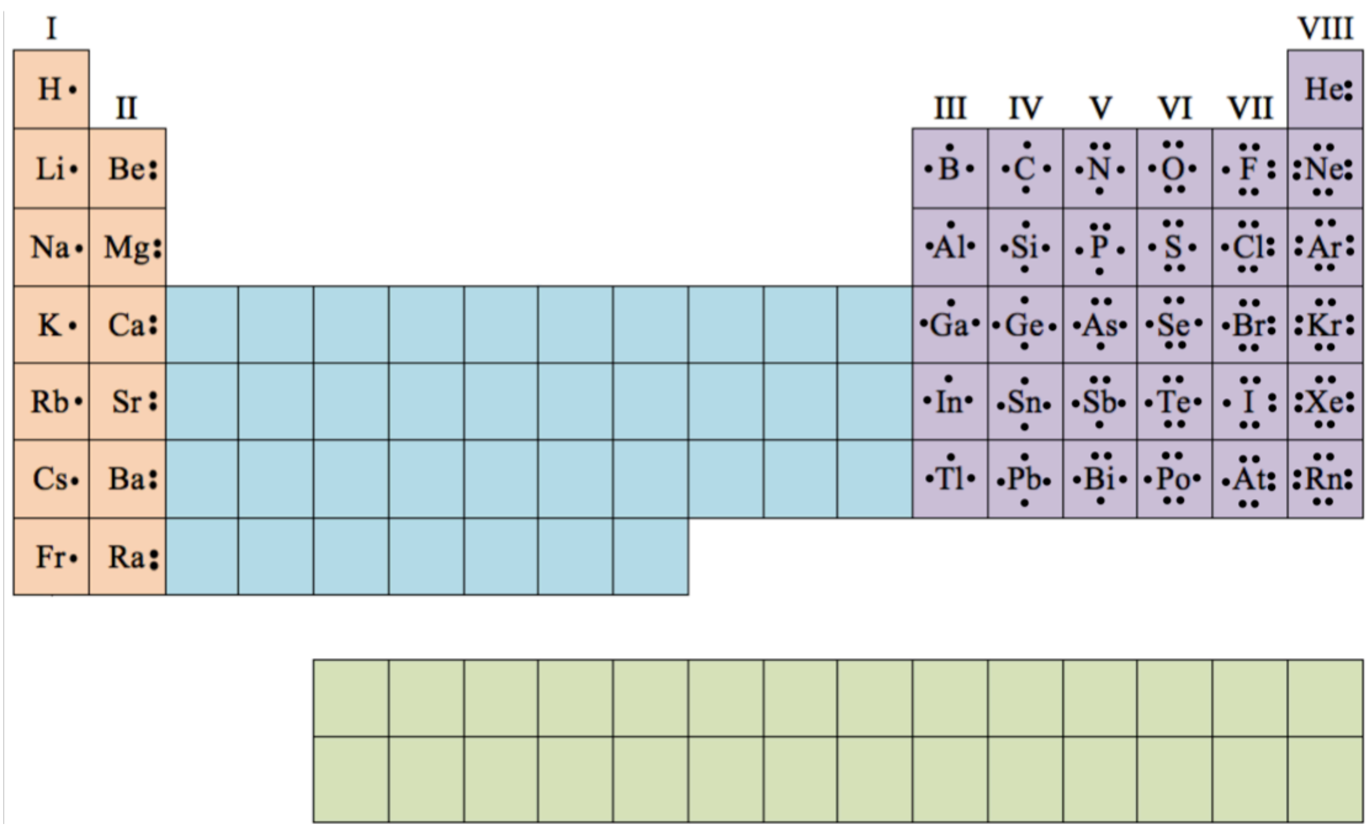
c. bromine (Br)

d. 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.

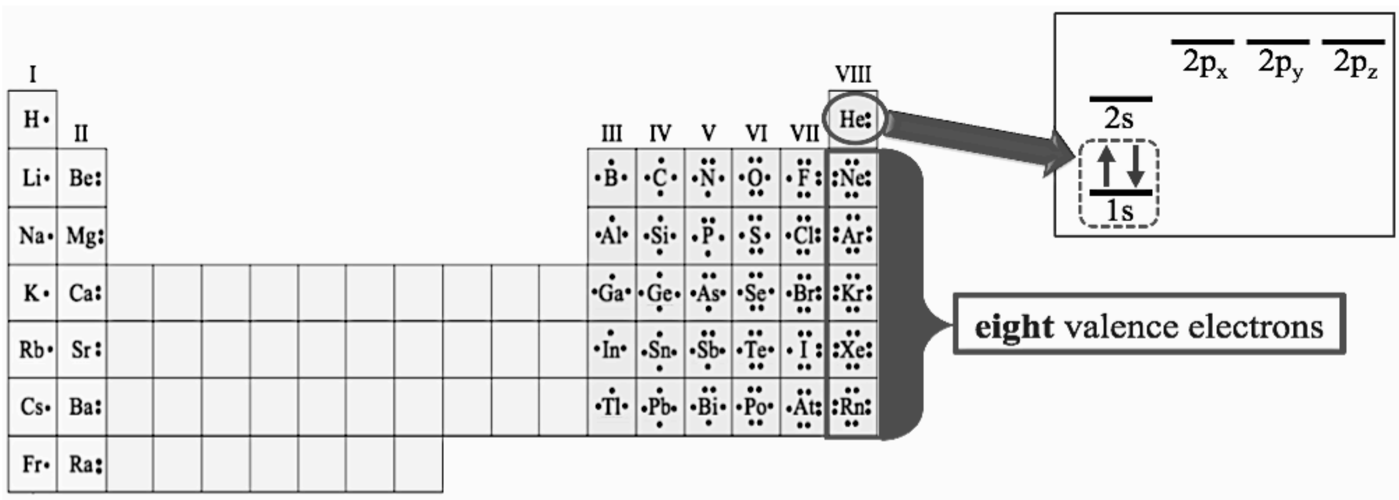


Noble Gases and the Octet Rule

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 it's 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_2O), carbon dioxide (CO_2), 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

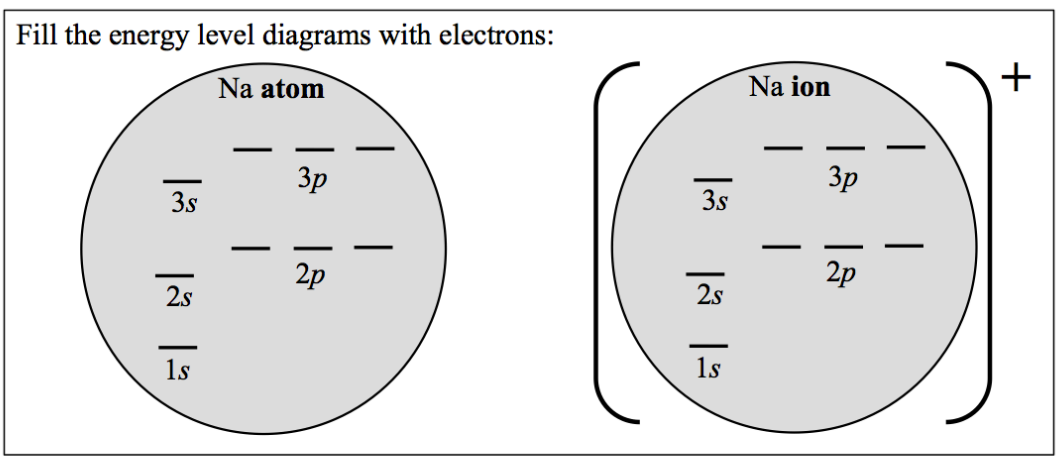
The ***Octet Rule*** can be used to predict the formation of ions.

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)

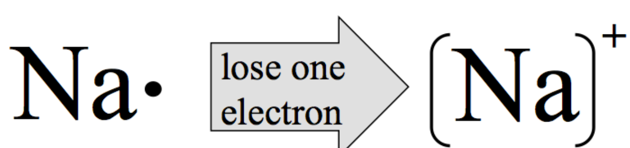
- A sodium atom has ______ protons and _____ electrons.
- How many valence electrons does the sodium atom have? _____
- How many valence electrons does sodium want? _____



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

Sodium will lose________ electron to become a sodium ion (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!



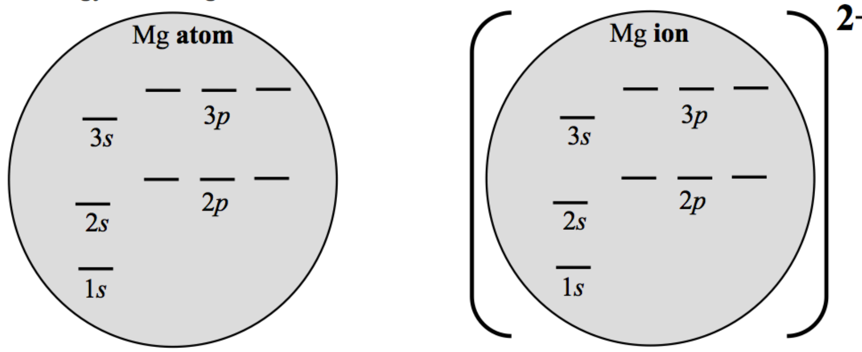
electron dot structure for a **Sodium Atom**

electron dot structure for a **Sodium Ion**

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.

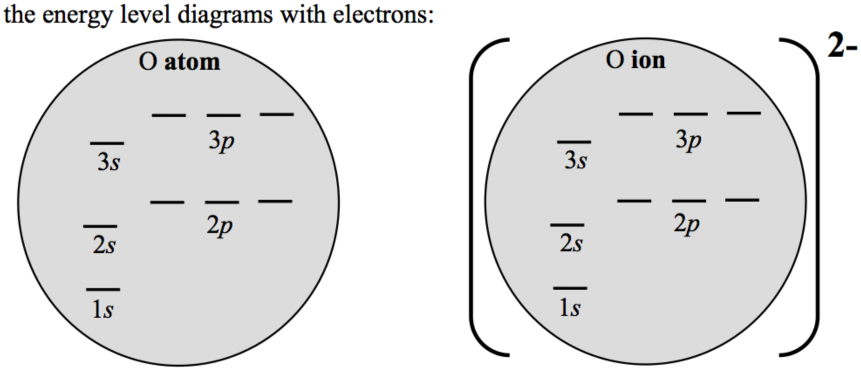
Magnesium will lose ______ electrons to become a magnesium ion (Mg^{2+}).

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 (0^{2-}).

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

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

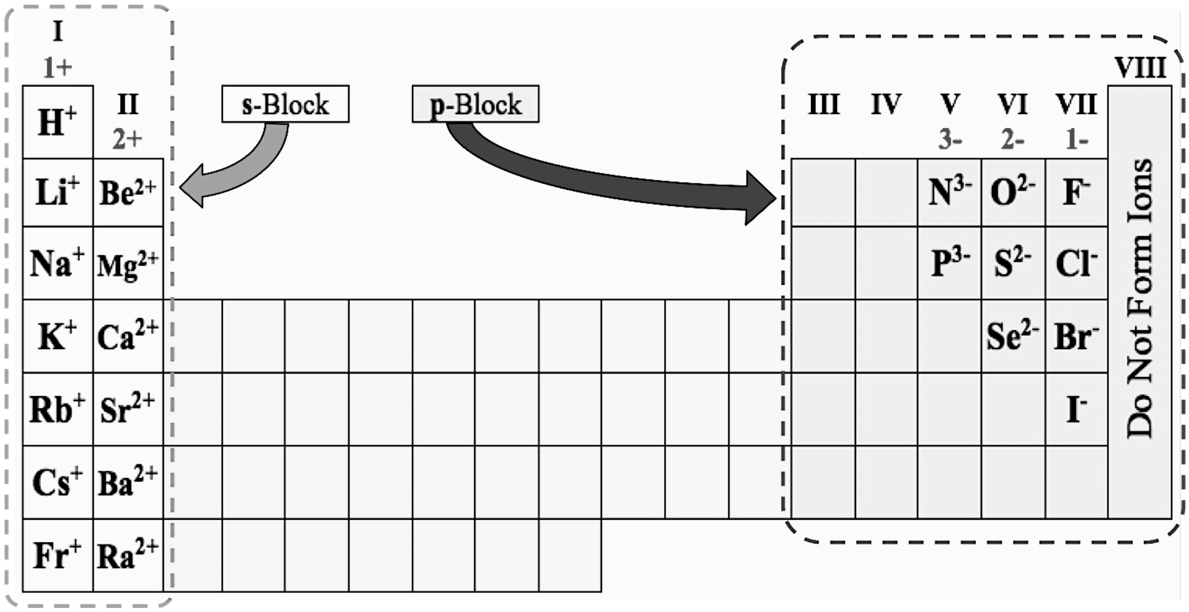
gain two electrons

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.

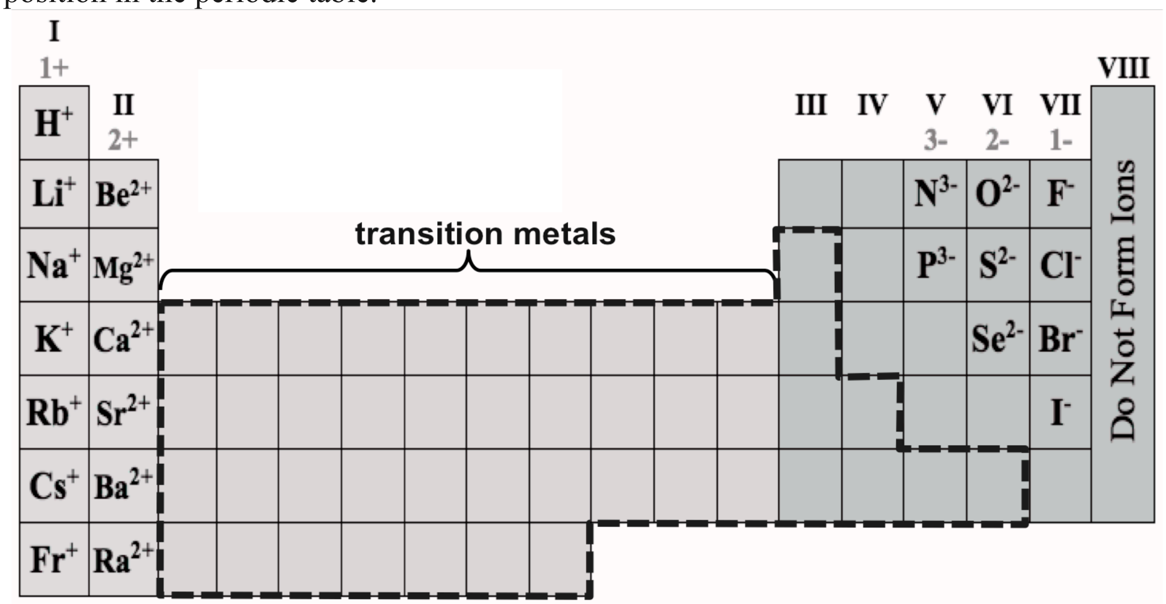
Understanding Check: What would be the charge of an **ion** formed from a **chlorine** atom?
HINT: Begin with the electron dot structure for a chlorine atom.

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
s-Block Elements			
Group I	1	Lose 1 electron	1+
Group II	2	Lose 2 electrons	2+
p-Block Nonmetal Elements			
Group III	There are no Group III nonmetals (only metals and metalloids)		
Group IV	4	Do not form ions, high energy to gain **or** 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.	

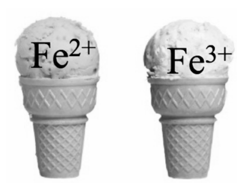
The charge of the ions formed from the **transition metals** and **p-block metals** ***cannot*** always be predicted by their position in the periodic table.



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

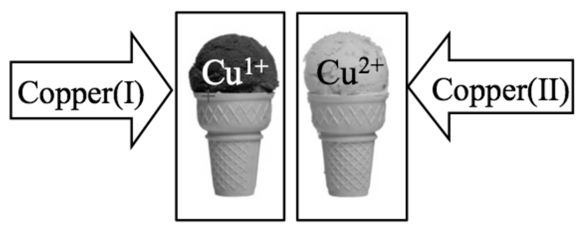
Example: Iron (Fe):

Iron (Fe) ions can come as Fe^{2+} or Fe^{3+}



Example: Copper (Cu):

Copper (Cu) ions can come as Cu^{1+}or Cu^{2+}



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^{1+} and "copper two" for Cu^{2+}.

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 $\mathbf{Cd^{2+}}$.

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 one charge			
Name	Charge	Name	Charge
aluminum ion	Al^{3+}	cadmium ion	Cd^{2+}
silver ion	Ag^{+}	zinc ion	Zn^{2+}
Ions that occur with multiple charges			
Name	Charge	Name	Charge
copper(I) ion	Cu^{+}	tin(II) ion	Sn^{2+}
copper(II) ion	Cu^{2+}	tin(IV) ion	Sn^{4+}
iron(II) ion	Fe^{2+}	lead(II) ion	Pb^{2+}
iron(III) ion	Fe^{3+}	lead(IV) ion	Pb^{4+}
cobalt(II) ion	Co^{2+}	mercury(I) ion	Hg^{+}
cobalt(III) ion	Co^{3+}	mercury(II) ion	Hg^{2+}

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^{+}** is referred to as a sodium ion.
  - **Mg^{2+}** 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^{2+}** is referred to as an iron(II) ion
  - **Fe^{3+}** 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^{-}** is referred to as a fluor**ide** ion.
  - **O^{2-}** is referred to as an ox**ide** ion.

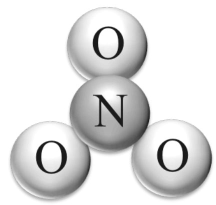
Polyatomic Ions

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

If the resulting particle has the s*ame 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)
Oxygen (8 electrons, 8 protons)
Oxygen (8 electrons, 8 protons)
Oxygen (8 electrons, 8 protons)
\+ **one extra electron**

Nitrate Ion
NO_3^-

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	
H_3O^+ hydronium ion	NH_4^+ ammonium ion
POLYATOMIC ANIONS	
OH^- hydroxide ion	HSO_4^- hydrogen sulfate (or bisulfate) ion
CO_3^{2-} carbonate ion	PO_4^{3-} phosphate ion
HCO_3^- bicarbonate (also called hydrogen carbonate) ion	HPO_4^{2-} hydrogen phosphate ion
NO_2^- nitrite ion	$H_2PO_4^-$ dihydrogen phosphate ion
NO_3^- nitrate ion	CrO_4^{2-} chromate ion
SO_3^{2-} sulfite ion	$Cr_2O_7^{2-}$ dichromate ion
SO_4^{2-} sulfate ion	$C_2H_3O_2^-$ acetate ion (sometimes written as $CH_3CO_2^-$)
	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_2O), carbon dioxide (CO_2), and table salt (sodium chloride).

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

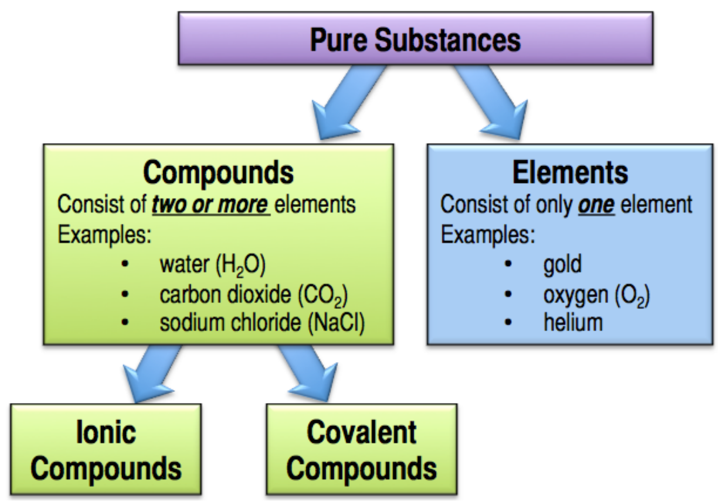
There are *three types* of **chemical bonding**:

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_2O 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.

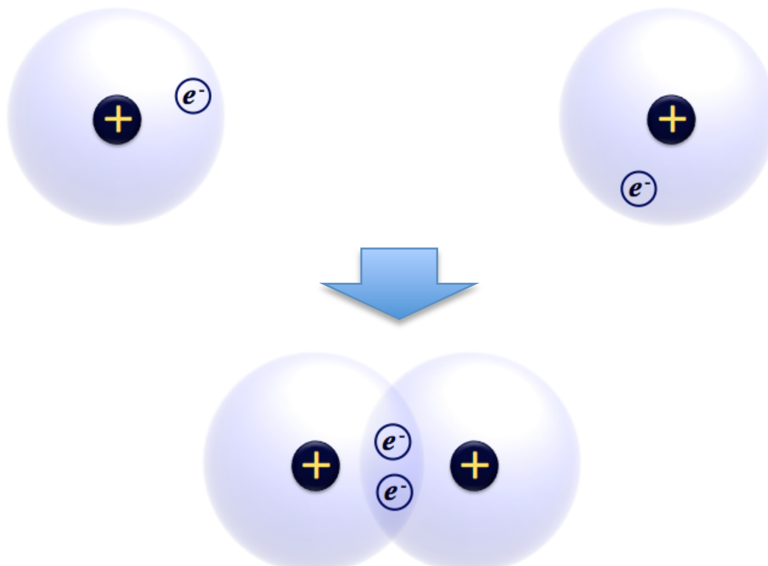
The resulting collection of atoms results in the formation of either _________________ or ***polyatomic ions***.

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

Covalent bonding occurs between __________________atoms.

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.

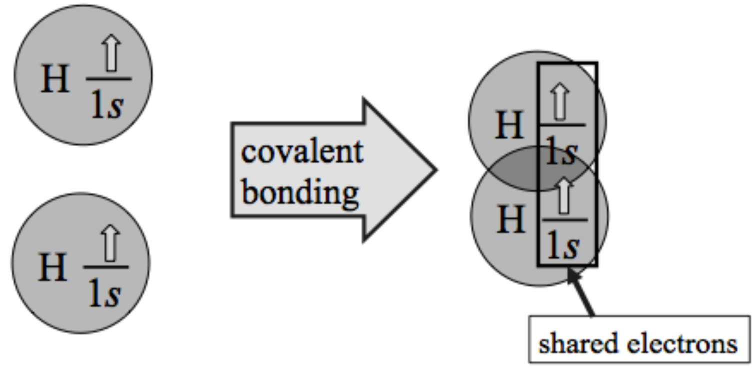
The Octet Rule in the Formation of Molecules

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

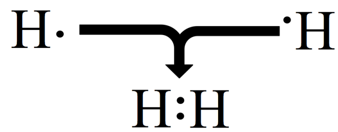
The octet rule in the formation of molecules is: ***molecules tend to form such that the atoms are surrounded by an octet (eight) of valence electrons (except for hydrogen and helium that have two electrons).***

Example: H_2
(recall that H and He are stable with two valence electrons)

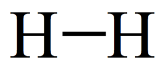
When a covalent bond forms, each hydrogen atom "feels" **two** electrons in its outermost shell.



The H_2 covalent bond can also be illustrated with electron dot structures.



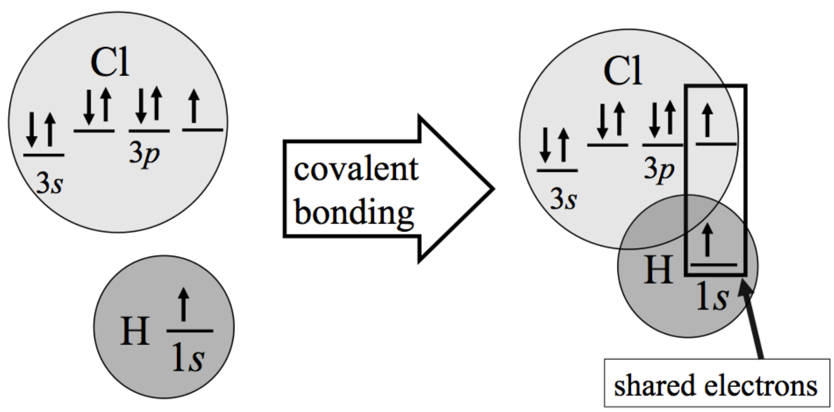
The two electrons _____________the atoms are shared in a covalent bond.
Chemist use a line to represent ____ electrons in a covalent bond.



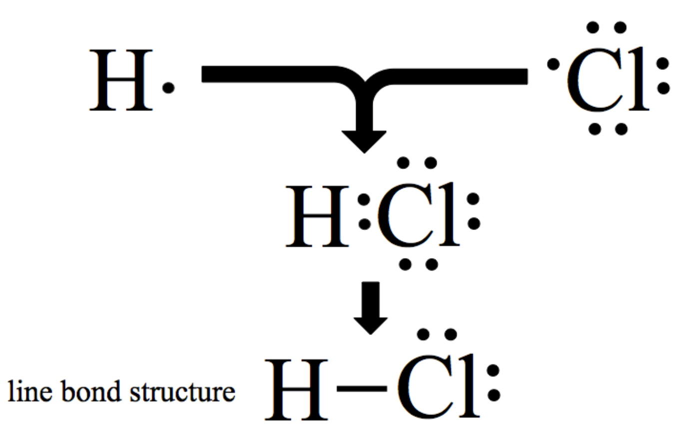
These drawings are called __________________________________.

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.

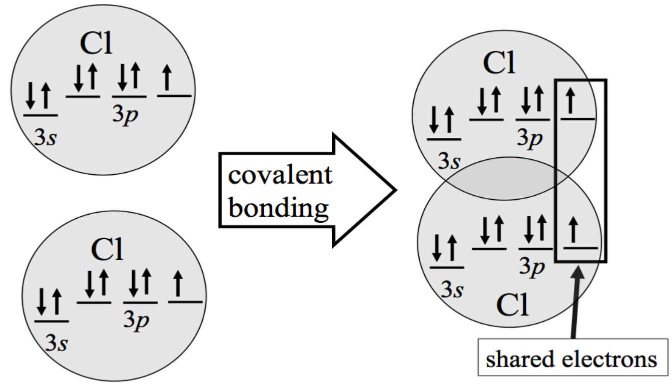


The HCl covalent bond can also be illustrated using electron dot structures.



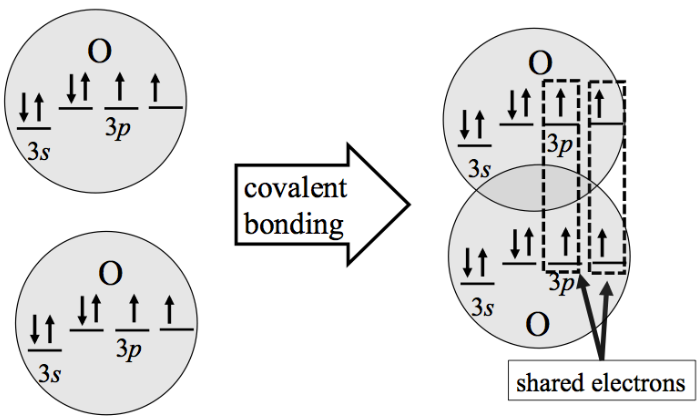
Let’s do another example: **Cl_2 (chlorine gas)**.

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



You try it:
Draw the ***line bond structure*** for Cl_2.
- Start with the electron dot structure for two Cl atoms.

Let’s do **oxygen gas (O_2)**.

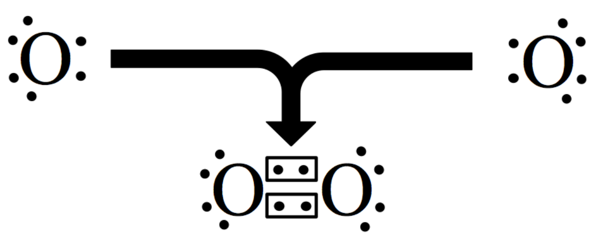


In O_2, ***two pairs*** of electrons are shared.

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

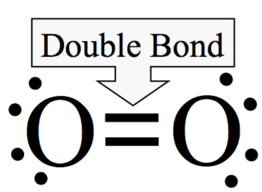
Let’s draw the line bond structure for **oxygen gas (O_2)**.

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



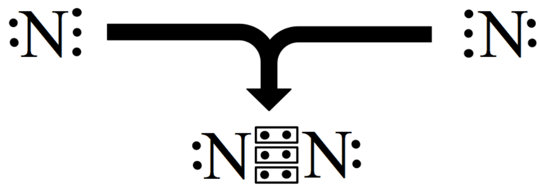
We use lines to represent ____________electron pairs.

When atoms are bonded with 2 pairs of electrons it is called a ___________ _________.



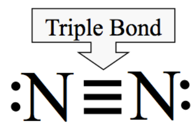
Let’s draw the line bond structure for nitrogen gas (**N_2**)

- 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
H–H	H_2
H–O–H	H_2O

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_2**, **Cl_2**, and **O_2**.

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

molecular formula	name
H_2	hydrogen
O_2	oxygen

Naming Binary Covalent (Molecular) Compounds

____________ **covalent compounds** contain only ***two*** ______________ (the "**bi-**" prefix indicates "***two***").

- Examples of binary covalent compounds are **HCl**, $\mathbf{H_2O}$, and $\mathbf{CO_2}$.

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:** $\mathbf{CO_2}$
       ~~mono~~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: CO**
       carbon mon~~o~~oxide → 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: $\mathbf{CCl_4}$

Answer: carbon tetrachloride

Understanding Check

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

$\mathbf{CF_4}$ ________________________________

$\mathbf{N_2O}$ ________________________________

$\mathbf{SF_6}$ ________________________________

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**.

N_2F_4

Understanding Check

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

- arsenic trichloride ______________
- dinitrogen pentoxide ______________
- tetraphosphorus decoxide ____________

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 ($C_6H_{12}O_6$)
- Acetone (C_3H_6O)

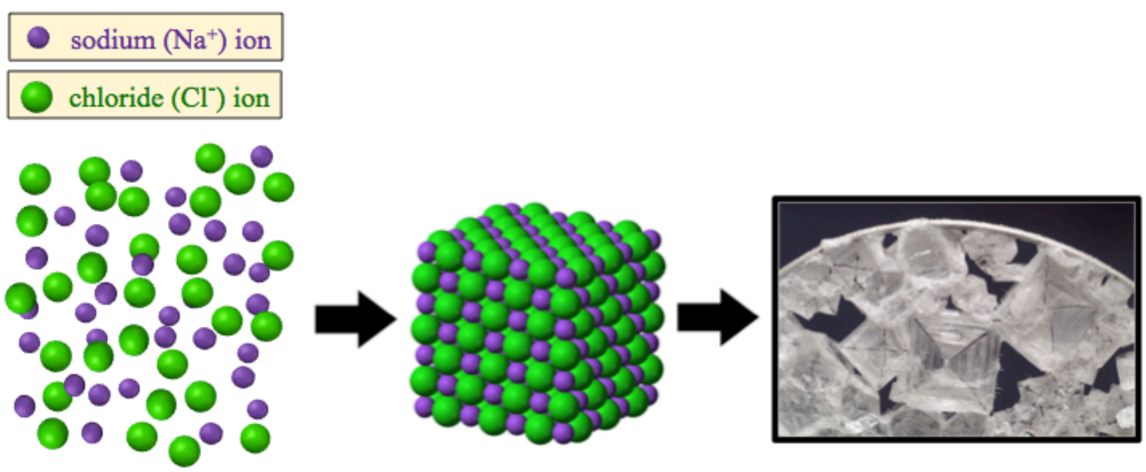
Ionic Compounds

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*** _____________.

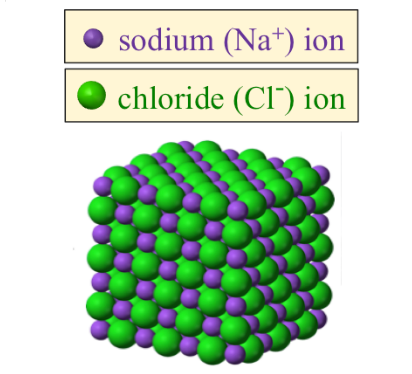
It is this regular, repeating structure on the scale of the individual ions that give crystals the interesting geometrical shapes that we see on the macro-scale when we look at them with our eyes or with a microscope.

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 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 apposed 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:

$$NaCl$$

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:

$$CaF_2$$

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

$$Mg_{\square}Cl_{\square}$$

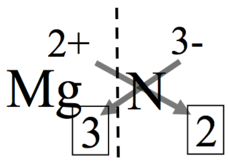
Understanding Check:
Write the ***formula unit*** for the compound formed by combining potassium and oxide ions.

$$K_{\square}O_{\square}$$

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

$$Mg_{\square}N_{\square}$$

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:

Cu^{+} and O^{2-}

Fe^{3+} and S^{2-}

Cu^{2+} and Cl^{-}

Mg^{2+} and O^{2-}

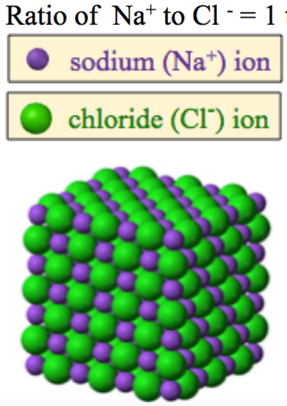
Sn^{4+} and S^{2-}

V^{3+} and Cl^{-}

Formula Unit vs. Molecular Formula

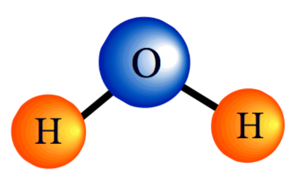
Formula Unit =
Lowest **RATIO** of ions

Example: NaCl
Ratio of Na^{+} to Cl^{-} = 1 to 1



Molecular Formula =
Actual **number** of atoms

Example: H_2O
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**.

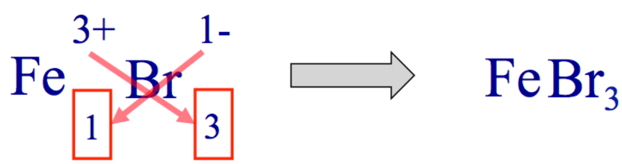
- We find the ion's charge from its position on the periodic table or, for polyatomic ions, we look it up in a table.
- You will know the charge for the metals t*hat occur with various charges* because the charge will be written in the compound's name in Roman numerals.

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.

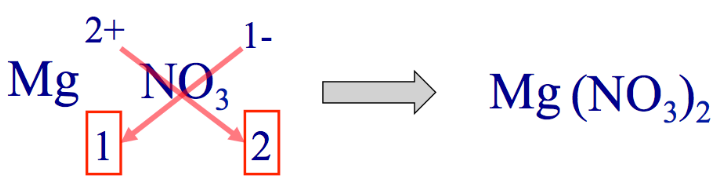
Example: Write the ***formula unit*** for **iron(III) bromide**.

- You will know the **charge** for the metals t*hat occur with various charges* because the charge will be written in the compound's name in **Roman numerals**.

iron(III) bromide



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:

a. sodium bicarbonate
b. sodium fluoride
c. iron(III) chloride
d. sodium carbonate
e. copper(II) sulfate
f. magnesium hydroxide

Method for Writing the Names of Ionic Compounds

1. 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.
2. 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: $\mathbf{MgCl_2}$

Name: **magnesium chloride**

Example: Name the following compound: $\mathbf{CuBr_2}$

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

Name: **copper(II) bromide**

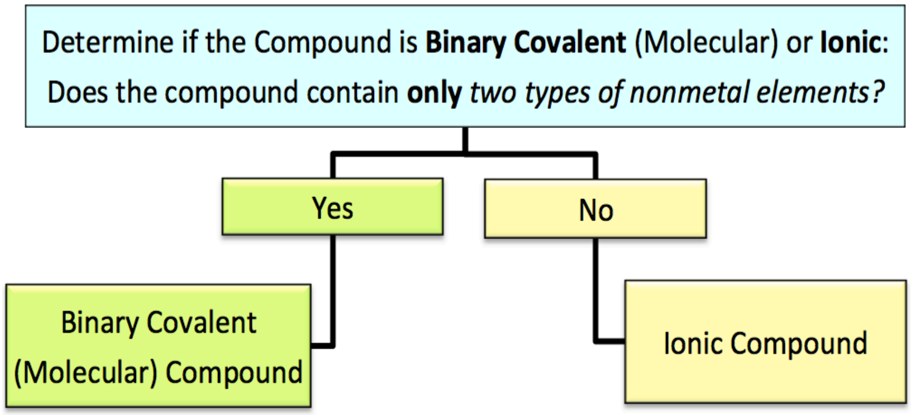
Complete the names of the following ionic compounds with variable charge metal ions:

$FeBr_2$	iron(__) bromide
$CuCl$	copper(__) chloride
SnO_2	______(__) _______
Fe_2O_3	___________________

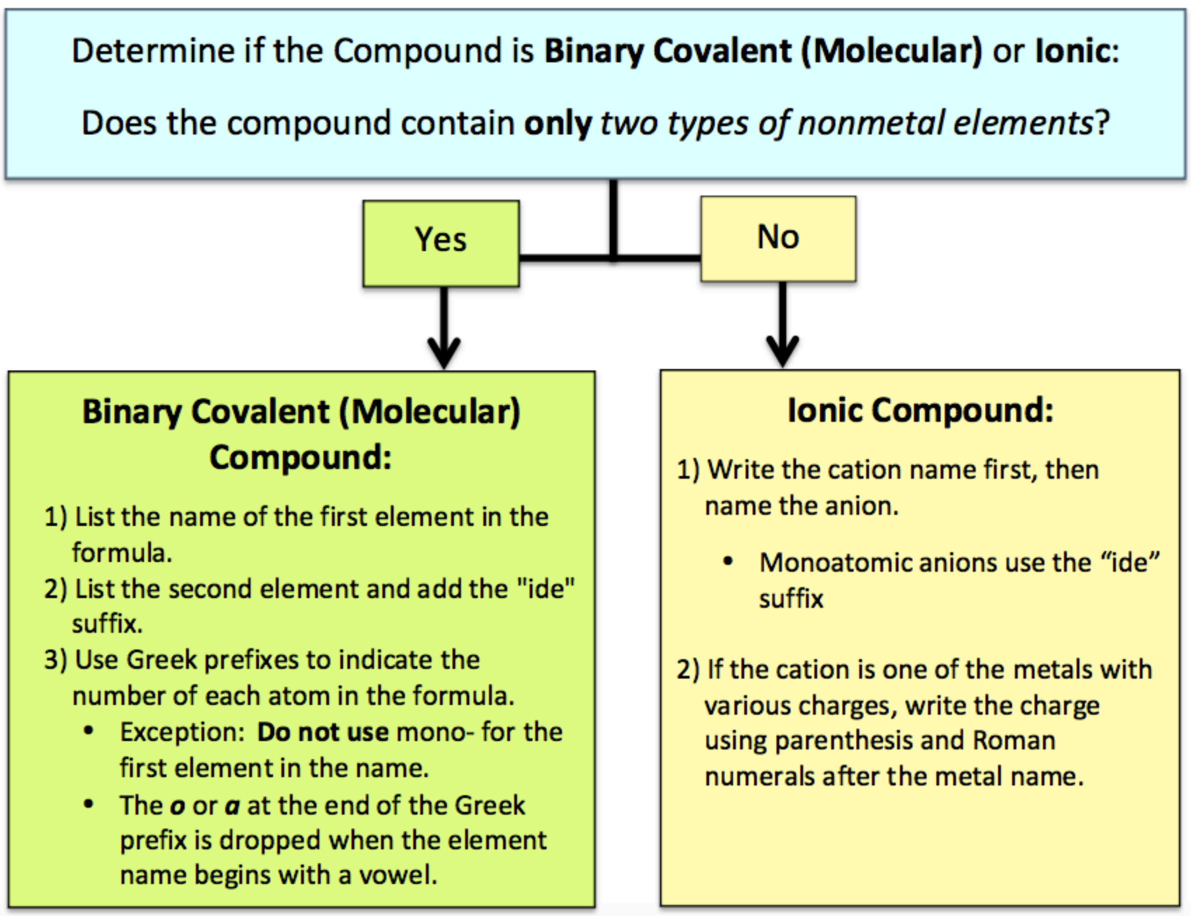
Name the following ionic compounds:

$NaCl$	_____________________
ZnI_2	_____________________
Al_2O_3	_____________________

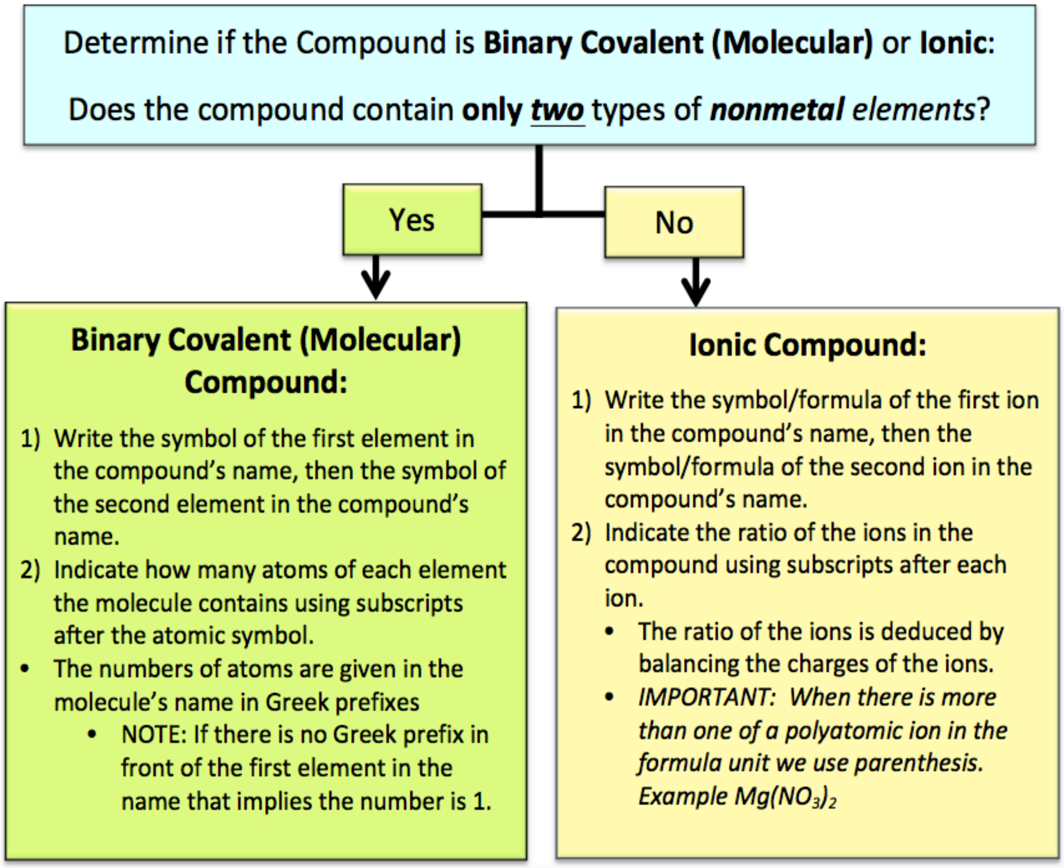
Naming Compound Summary



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**

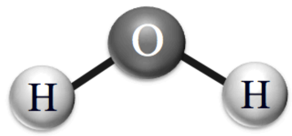
Molar Mass of Covalent Compounds (Molecules)

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_2O.



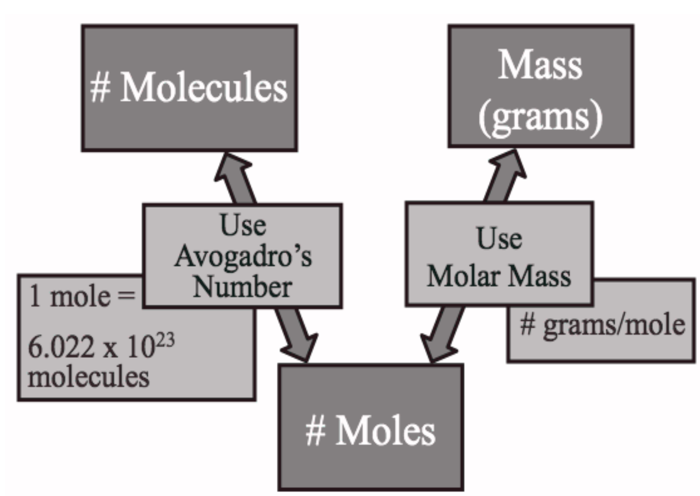
Atom	# of Atoms	Atomic Molar Mass	Total
oxygen	1	X 16.00 g/mole	16.00 g/mole
hydrogen	2	X 1.01 g/mole	2.02 g/mole
		Molar Mass of H_2O =	**18.02 g/mole**

One mole of H_2O (6.022 x 10^{23} molecules) has a mass of 18.02 grams

Understanding Check: Calculate the molar mass of CH_4 (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_4 is contained in 3.65 *moles*?

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

- **1 mole CH_4 = 16.05 grams**

The equivalence statements can be written as conversion factors:

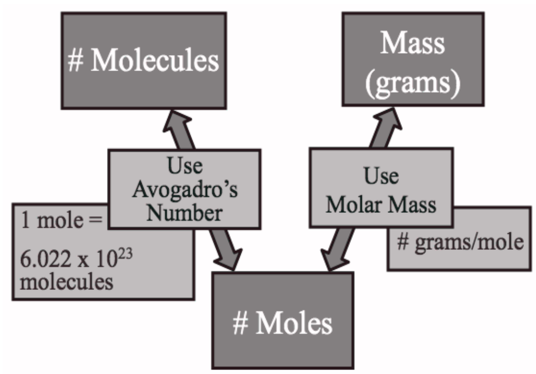
$$\left(\frac{1 \text{ mole } CH_4}{16.05 \text{ grams}}\right) \xleftrightarrow{\text{Conversion Factors}} \left(\frac{16.05 \text{ grams}}{1 \text{ mole } CH_4}\right)$$

$$\frac{3.65 \cancel{\text{ moles } CH_4}}{} \left| \frac{\textbf{16.05 grams } CH_4}{1 \cancel{\text{ mole } CH_4}} \right| = 58.6 \text{ grams } CH_4$$

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 6.022×10^{23} molecules.



You try one: How many H_2O ***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* elections, 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
Molar Mass (Formula Mass) of NaCl			**= 58.44 g/mole**

Example: What is the molar mass of iron(II) phosphate, $Fe_3(PO_4)_2$?

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*

Fe_3

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

two moles of *phosphate ions*

$(PO_4)_2$

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 based on: one phosphorus **and** four oxygens ***per ion***	= 189.94 g/mole
Molar Mass (Formula Mass) of $Fe_3(PO_4)_2$			**= 357.49 g/mole**

An Alternative Method:

Fe_3

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

$(PO_4)_2$

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

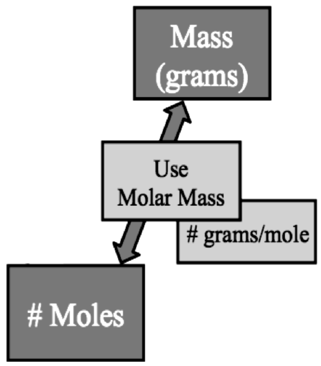
- ***two*** moles of phosphorous
- ***eight*** (**2** x 4) moles of oxygen

Three moles of **Fe**:	**3** x 55.85 g/mole =	**167.55** g/mole
Two moles of **P**:	**2** x 30.97 g/mole =	**061.94** g/mole
Eight moles of **O**:	**8** x 16.00 g/mole =	**128.00** g/mole
	The molar mass of $Fe_3(PO_4)_2$ is	**357.49** g/mole

Understanding Check: What is the molar mass of magnesium nitrate, $Mg(NO_3)_2$?

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 $Mg(NO_3)_2$?