Chapter 3: Compounds

Chapter 3 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.

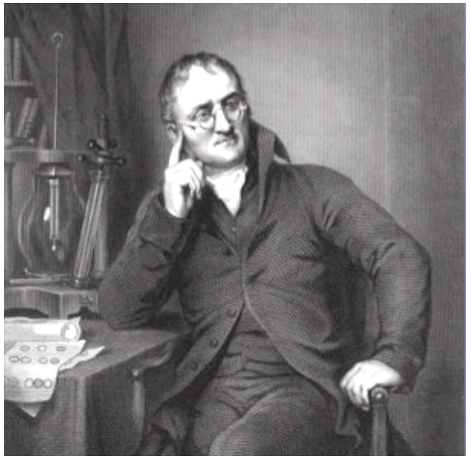
Scientists used ***light*** to study how electrons are arranged around the nucleus.

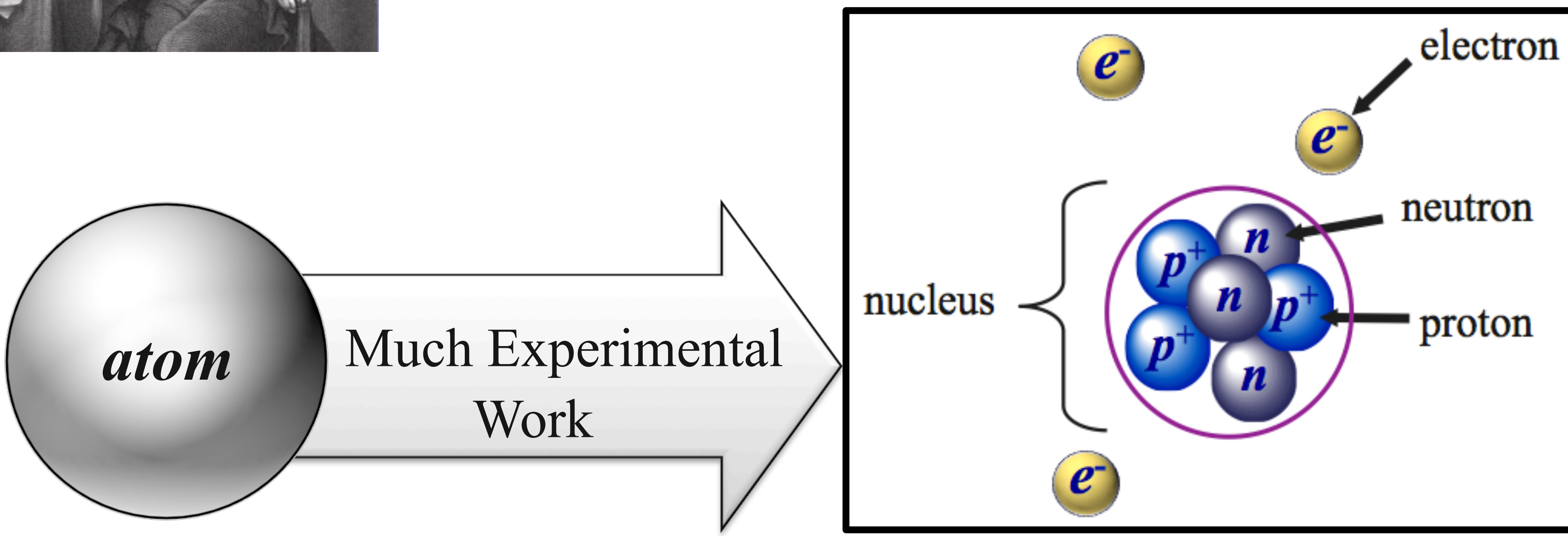
Energy, in the form of light or heat, can be ***absorbed*** by atoms.

Energy is absorbed by moving an electron to a new area.

Atoms **release** energy when electrons move *back* to ***low energy*** areas.

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





Dalton's Model of the Atom:

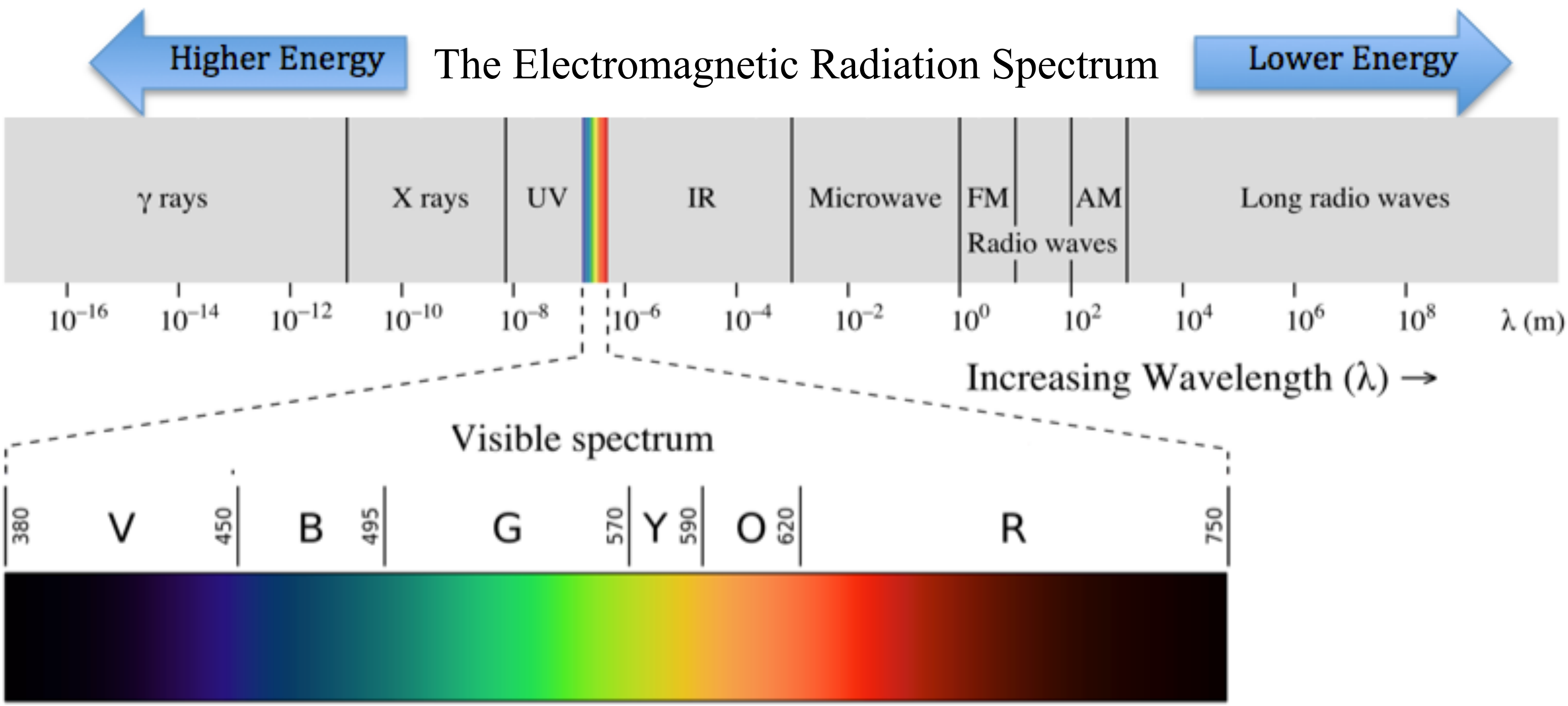
Atoms are the Smallest Particles

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.



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 (see video for color):

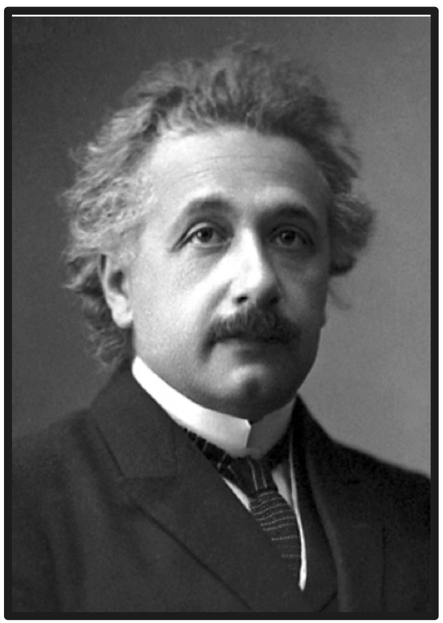
However, light with **discrete** (distinct) energies is emitted. For example, the pattern of light emitted from excited hydrogen atoms is (see video for color):

The Modern Model of the Atom

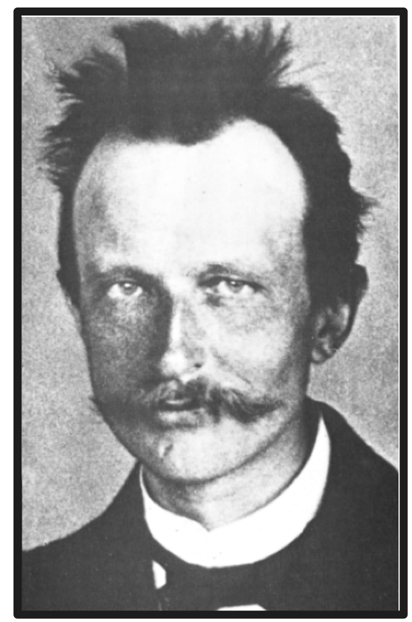
Our understanding of nature was dramatically changed when Max Planck and Albert Einstein introduced "***quantum mechanics***."

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

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



Albert Einstein (1879-1955)



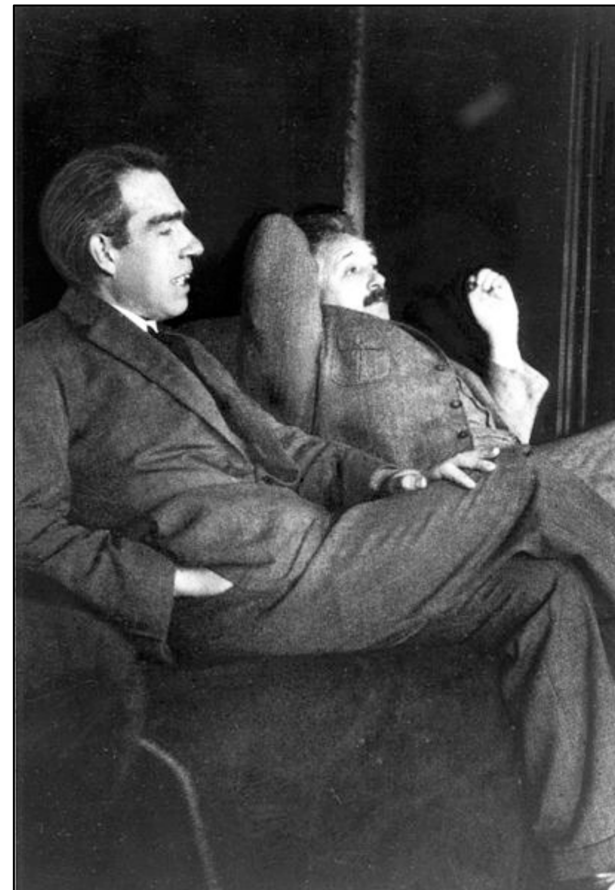
Max Planck (1858-1947)

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.

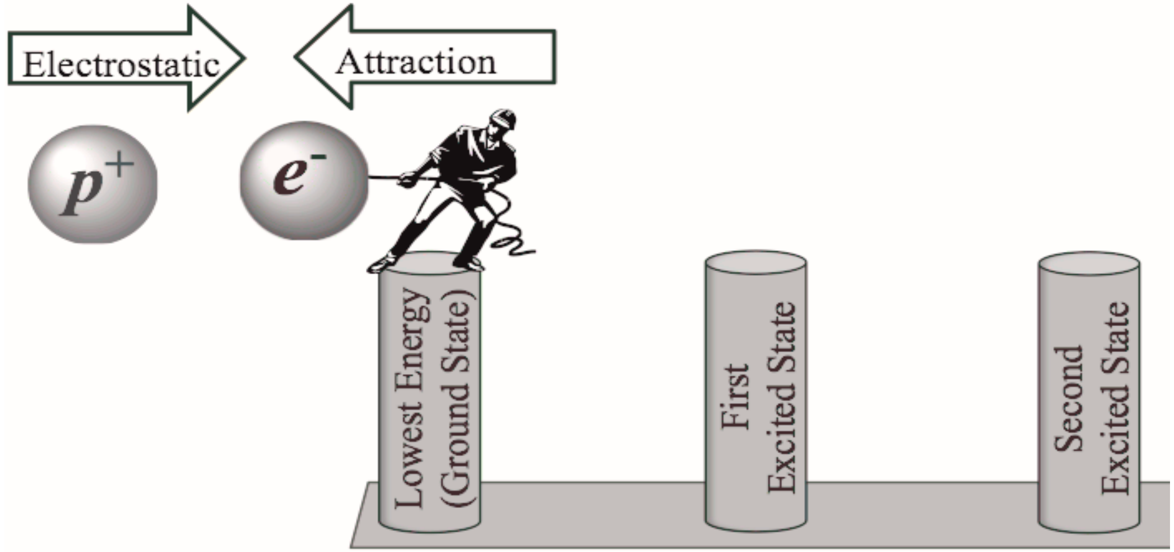
The observation that only discrete energies are emitted from excited atoms is explained using an atomic model that says that the ***electrons can only exist in certain areas and therefore atoms have discrete energies***.

- We say that the energy of atoms is "***quantized***."

- The first scientist to propose this model of the atom with discrete energy states was Niels Bohr.



Niels Bohr (left) with Albert Einstein



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

- 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).

When hydrogen's electron is in any other region than the ground state (lowest energy), we call that an ***excited state*** 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 *located* in atoms.**

2) **Understand how the location of electrons affect the *energy* of the atom.**

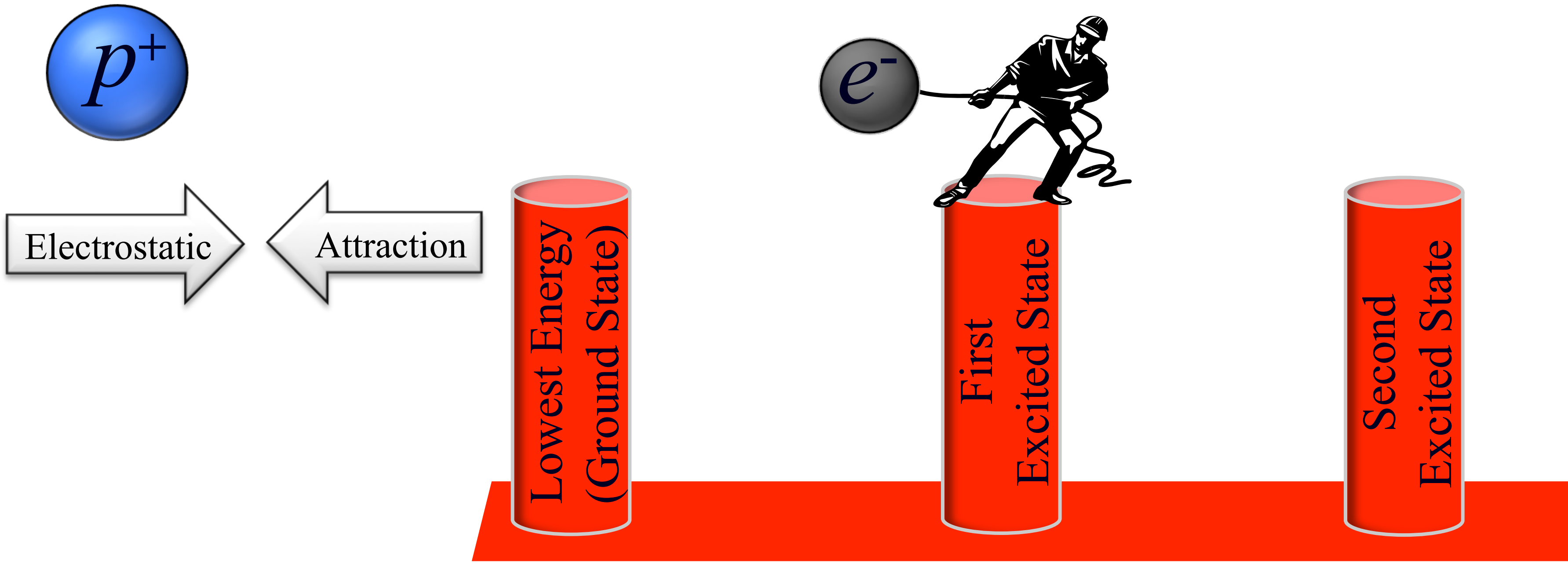
The Hydrogen Atom

Hydrogen is unique because it has only ***one*** electron.

Electrons exist in certain **three-dimensional regions** called ***orbitals***.

Orbitals can be described by these properties:

1. **The *average distance* an electron in a particular orbital is from the nucleus.**
    - As orbitals get larger, the average distance of an electron from the nucleus increases, therefore the ***larger the orbital*** occupied by an electron, the ***greater the energy***.



2. **The three-dimensional *shape* 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 the electron ***90%*** of the time.

The Language of Quantum Mechanics

The orbitals are *centered on the* ***nucleus***, and are labeled by a ***number***.

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 "***shell***" or "***quantum level***" and abbreviate it by using "***n***".

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 ***one*** orbital.

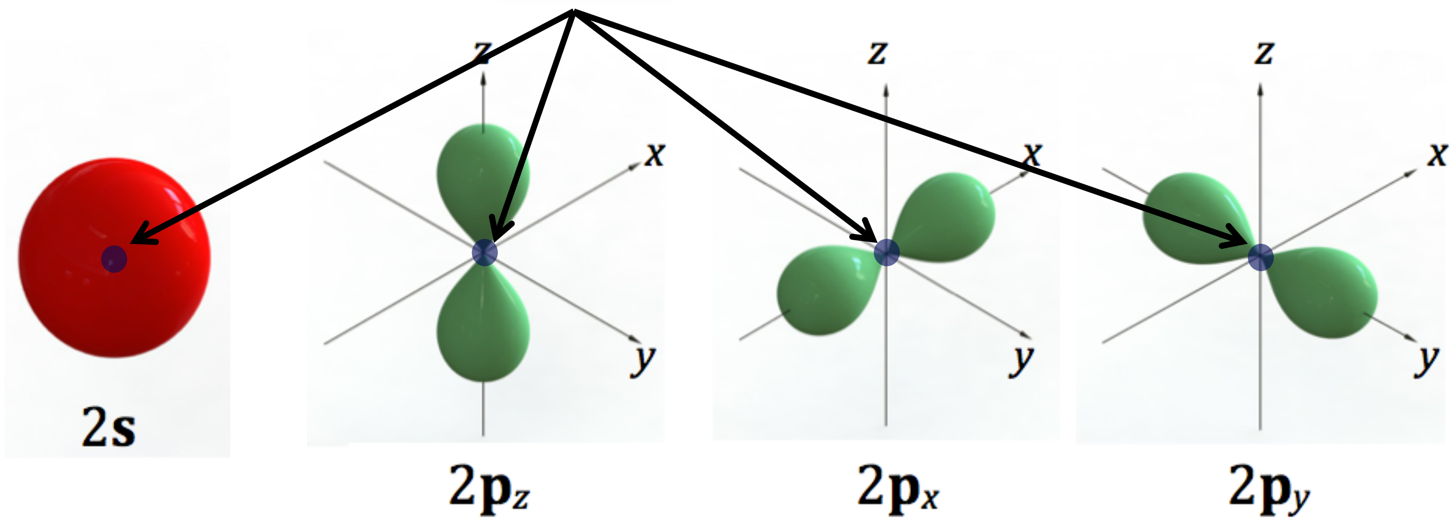
- It is called an ***s*** orbital.
- **"s"** represents the ***shape* of the orbital**, we use **1s** because **n=1**).
- **s** orbitals are ***spherical*** in shape.

Illustration of a **1s** Orbital

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

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

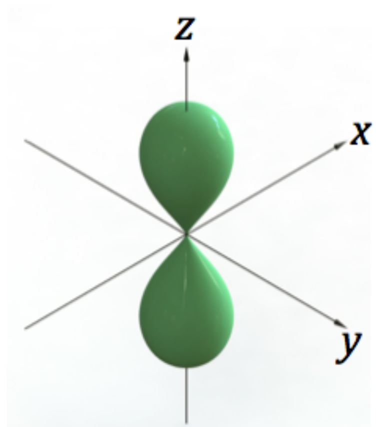
- There is **one 2s** orbital
  - ***All*** **s** orbitals are spherically shaped.
  - We use **2s** because **n**=**2**.
- There are **three 2p** orbitals.
  - **p** represents the shape; we use **2p** because **n**=**2**.
  - The **p** orbitals all have the **same shape** and only differ in *how they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



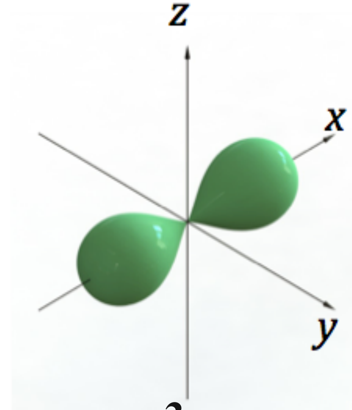
The **n**=**3** quantum level has ***nine*** orbitals.

- 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***.

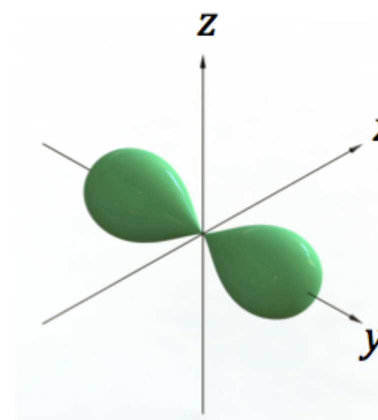

3s



$3p_z$

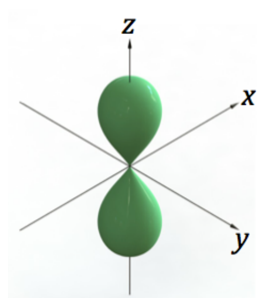


$3p_x$

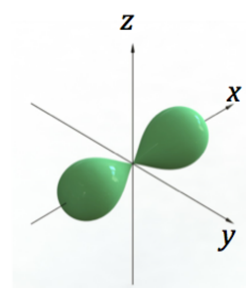


$3p_y$

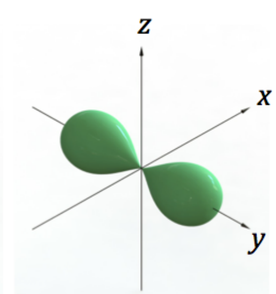

2s



$2p_z$

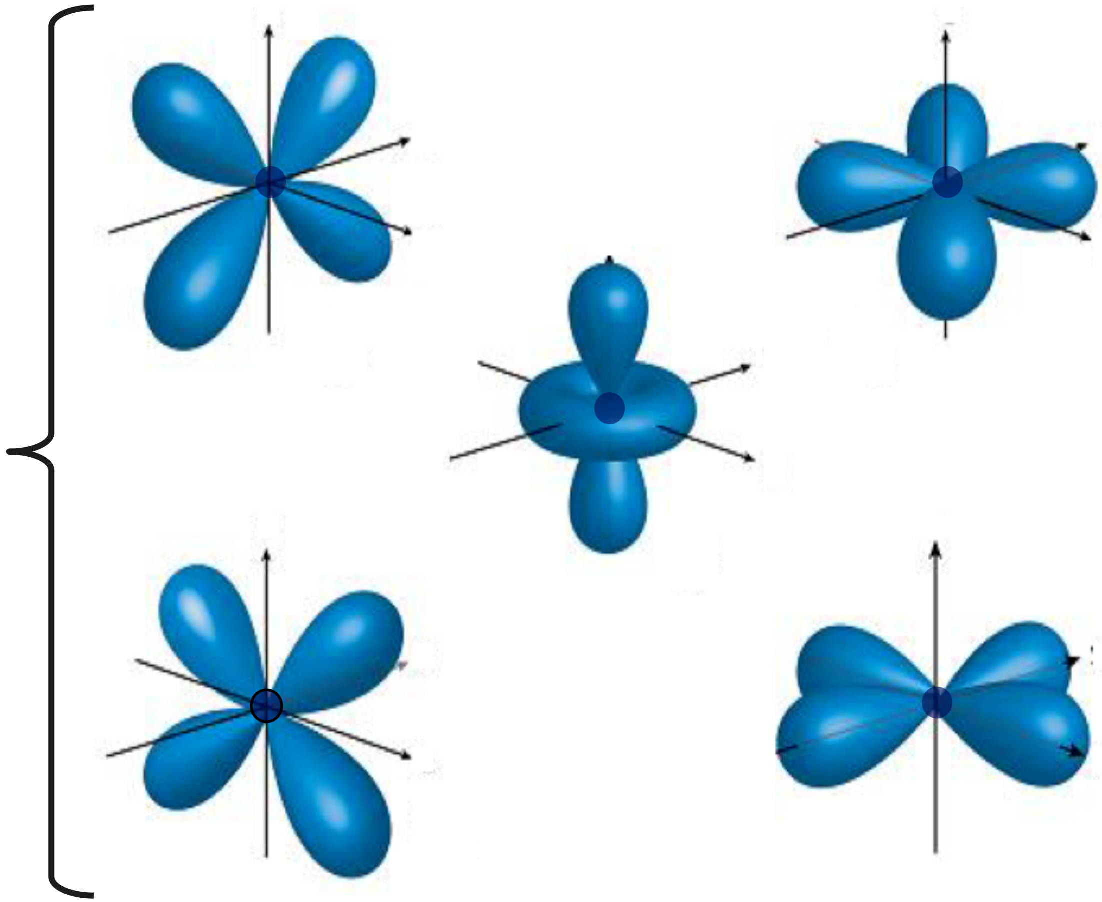


$2p_x$



$2p_y$

The **five 3d** orbitals



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

The **n**=**4** quantum level has ***sixteen*** 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*.

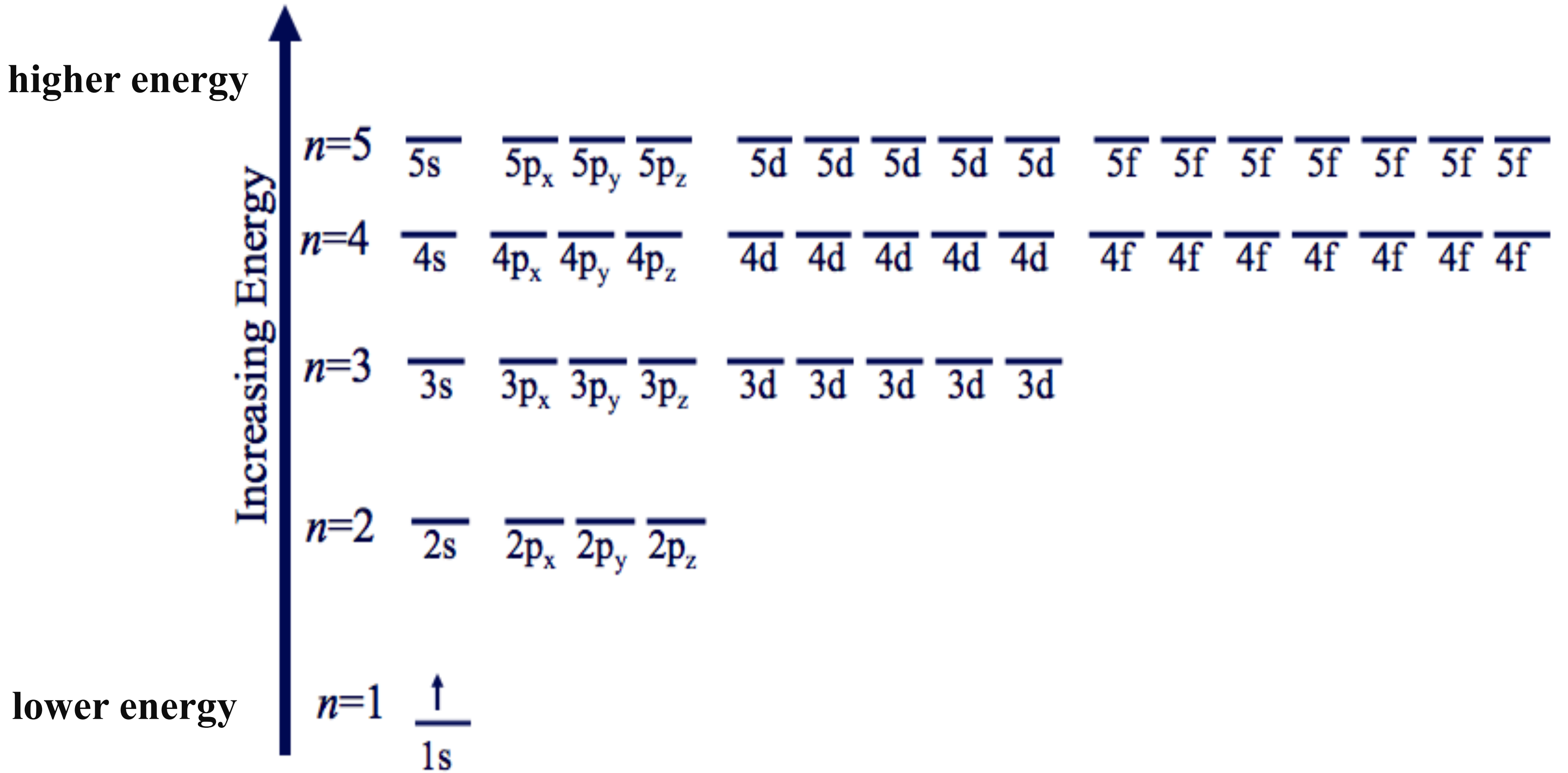
- The only time an electron can occupy any of those orbitals will be if the atom absorbs energy.

Energy Level Diagram for Hydrogen

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

The orbitals are arranged, from bottom to top, in order of increasing ***energy***.

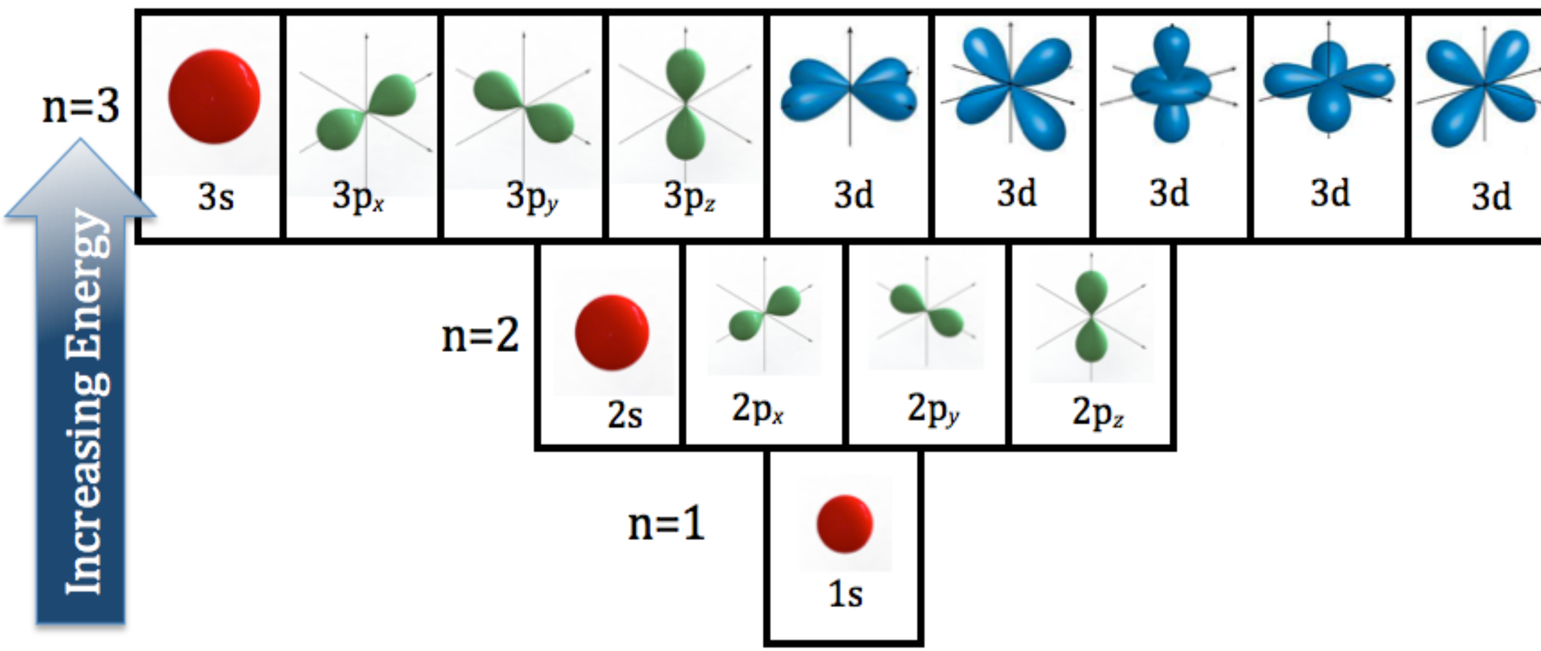
An electron is depicted as an ***arrow*** above the line that represents the orbital occupied by it.



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

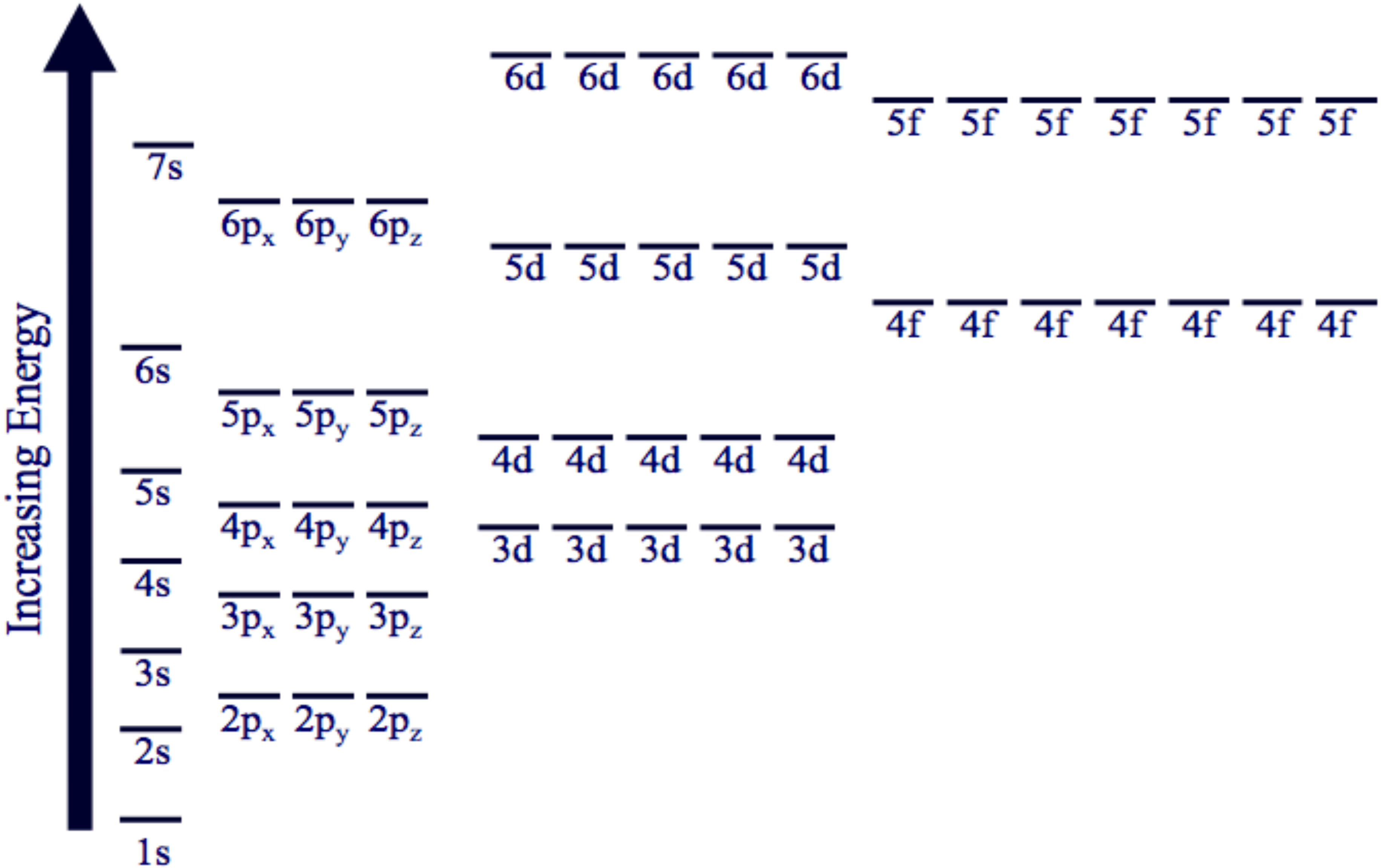
The different ***floors*** (levels) of the skyscraper represent the **quantum levels (n)**.

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

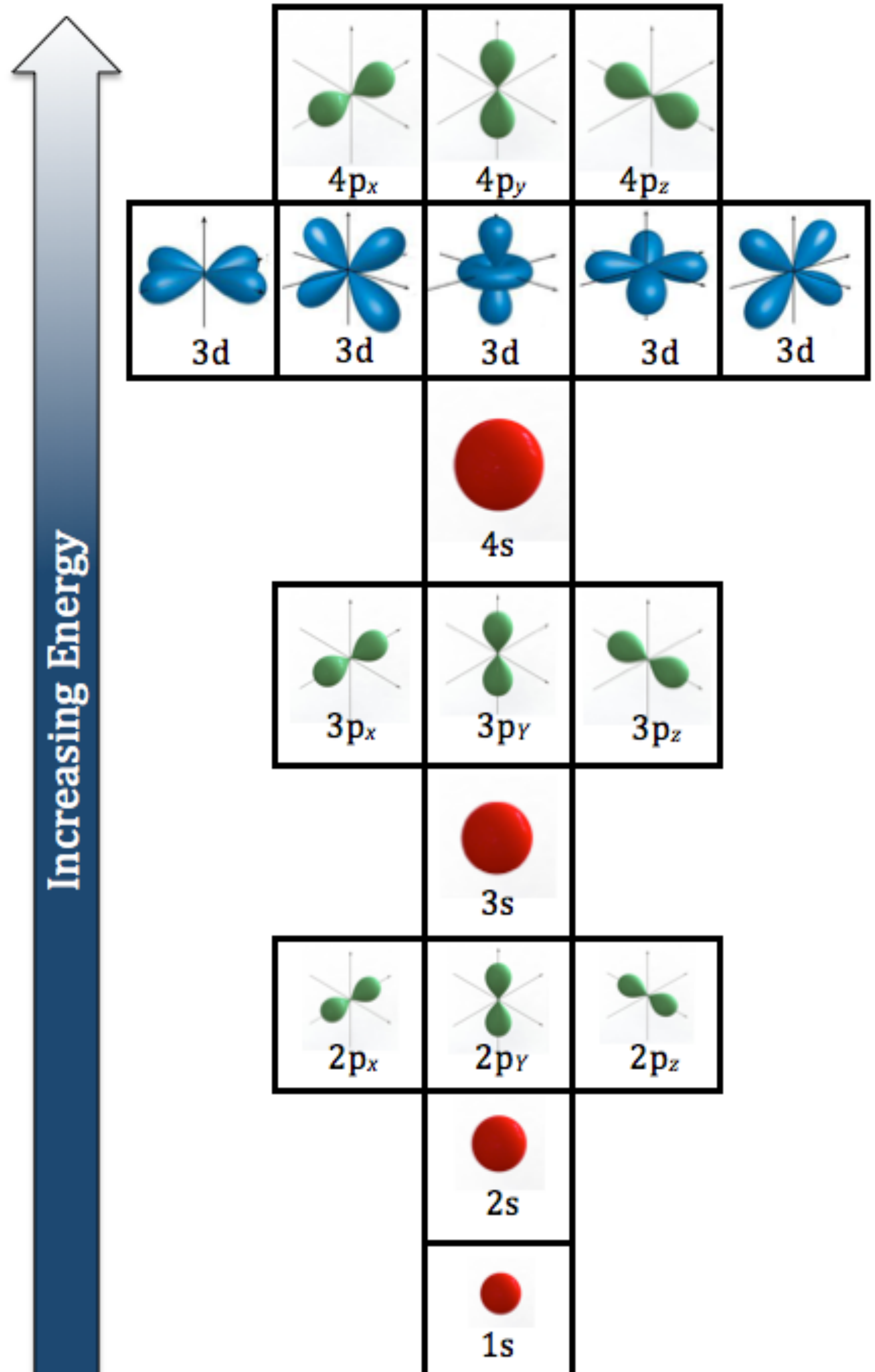


Atomic Model for Multi-Electron Atoms

Energy Level Diagram for Multi-Electron Atoms



Skyscraper Model for Multi-Electron Atoms



How are the electrons configured (arranged) into these orbitals?

Nature wants everything to be at the ***lowest*** possible energy.

Electron Configuration

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 three principles:

1) The Aufbau Principle

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

2) The Pauli Exclusion Principle

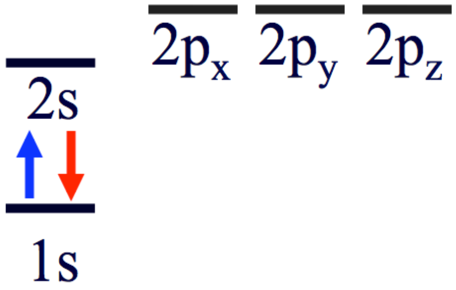
An orbital can hold a maximum of ***two*** electrons.

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

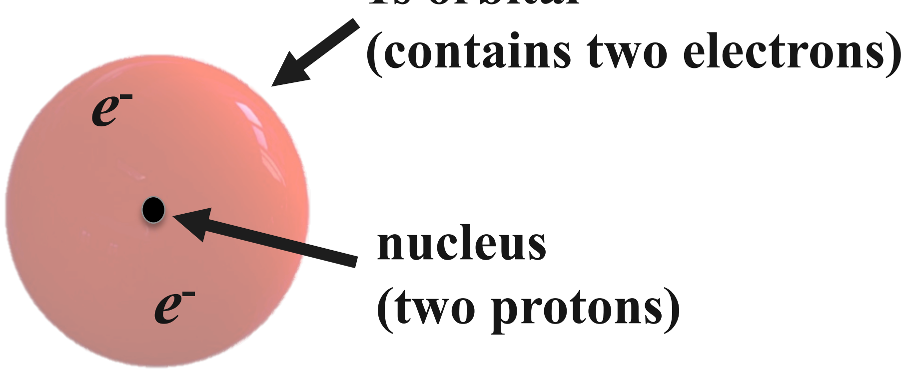
We call the spin states "**up**" or "**down**."

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

Example: The Electron Configuration of a Helium Atom (2 electrons)



Model of a Helium Atom



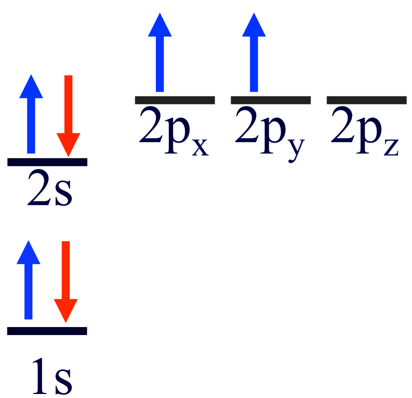
Having two electrons in the same orbital with **opposite** spin states is ***lower in energy than when both spins are the same.***

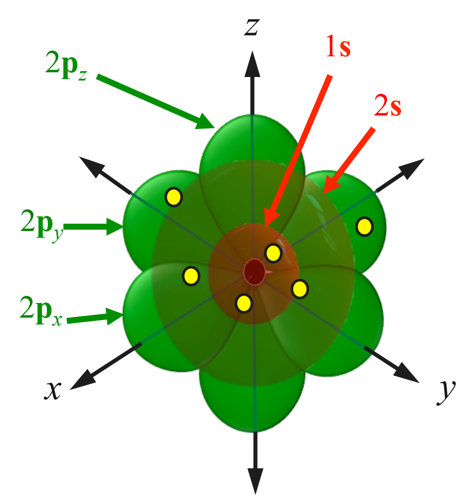
3) Hund's Rules

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.

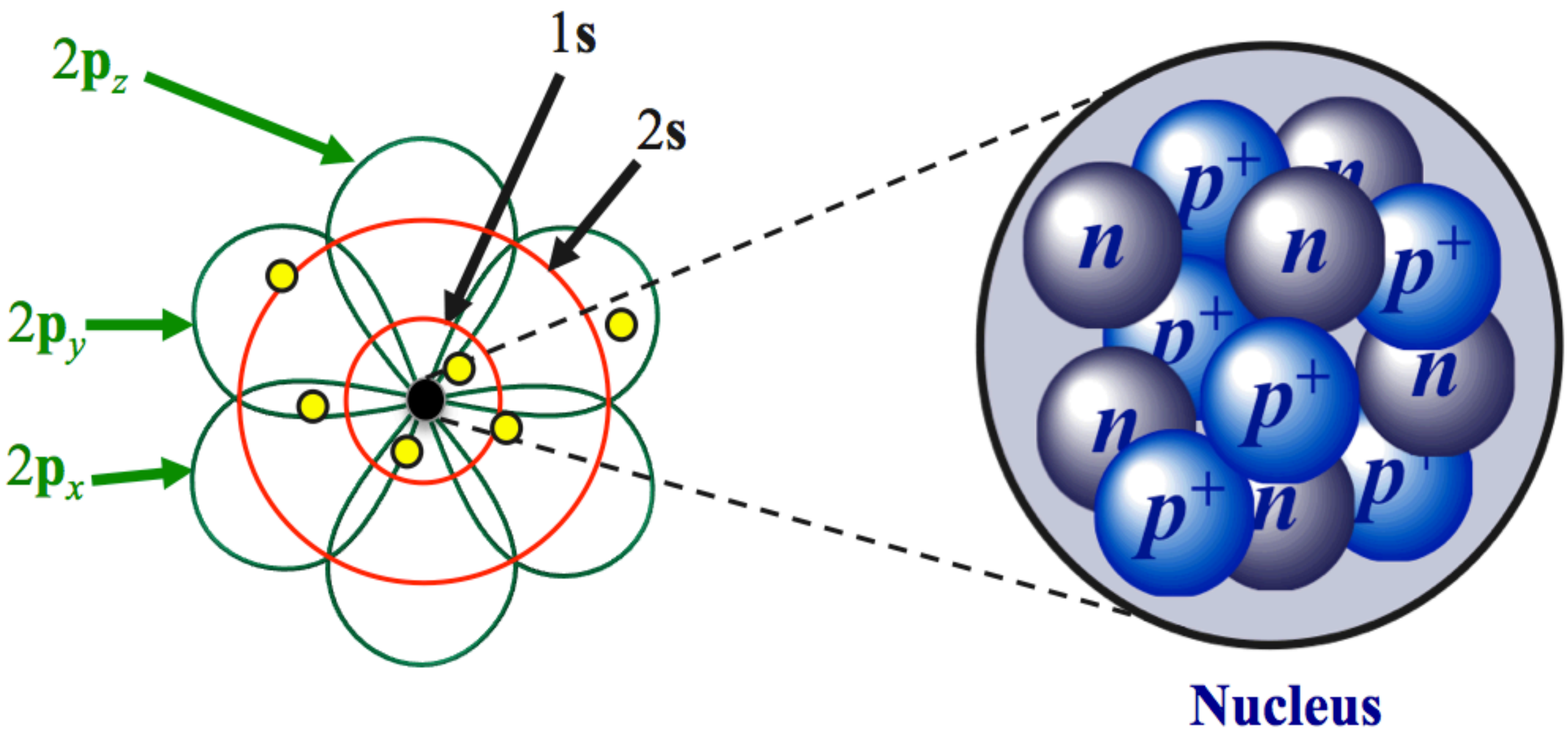
When electrons are configured into a *set of orbitals* that all have the same energy, the spins of the first electrons to be placed into each orbital are all in the same state (for example all "up").

Example: Electron Configuration of a Carbon Atom

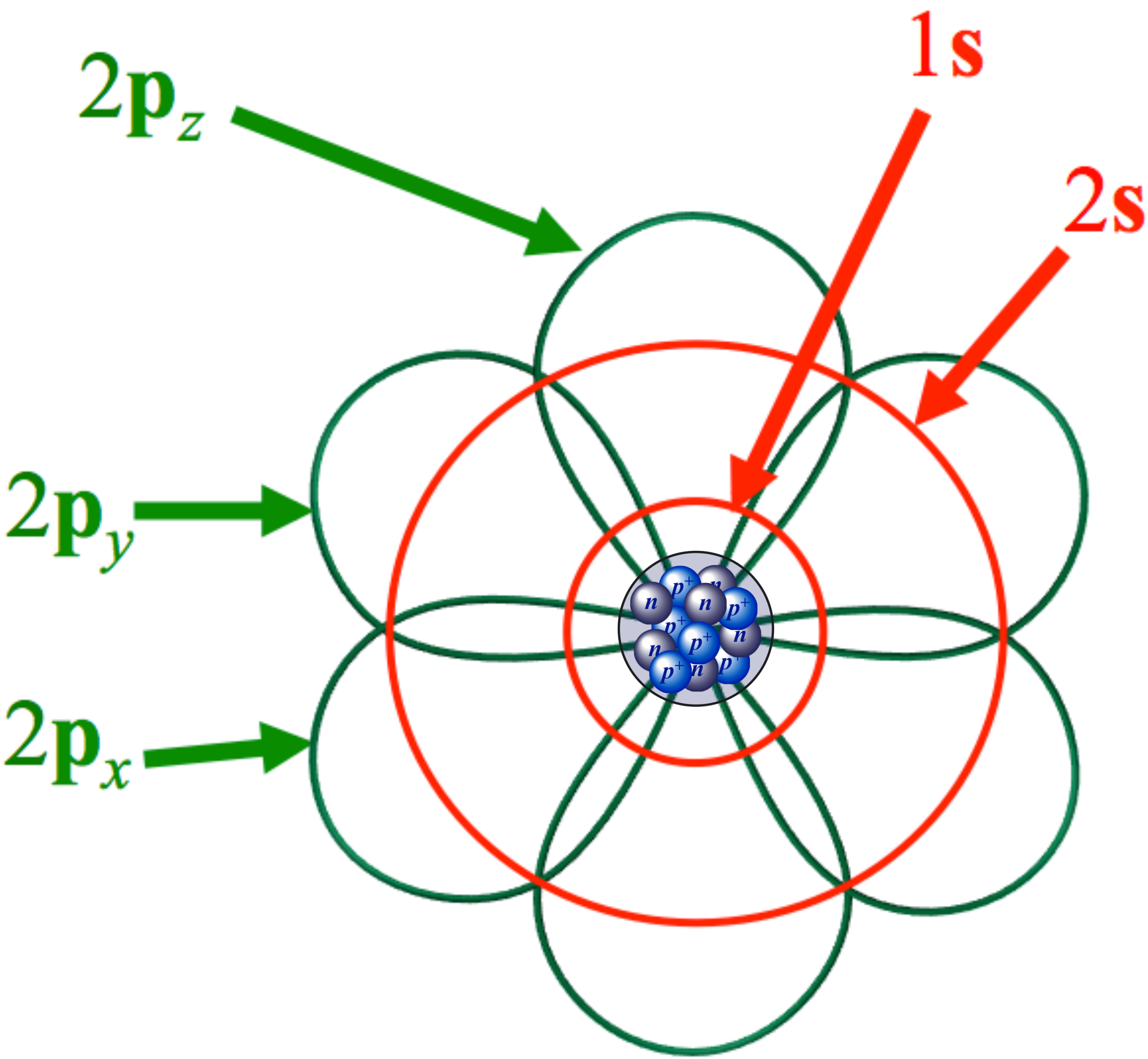




Drawing of a Carbon-12 Atom



Drawing of a Carbon-12 Atom



Understanding Check:
Energy Level Diagrams for Multi-Electron Atoms

Draw the energy level diagram for each of these atoms:

a) a neon (Ne) atom

b) an Iodine (I) atom

Valence Electrons

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

Language Reminder: "***shell***" = "***quantum level***" = "***energy level***"

Valence electrons are furthest away from the ***nucleus***.

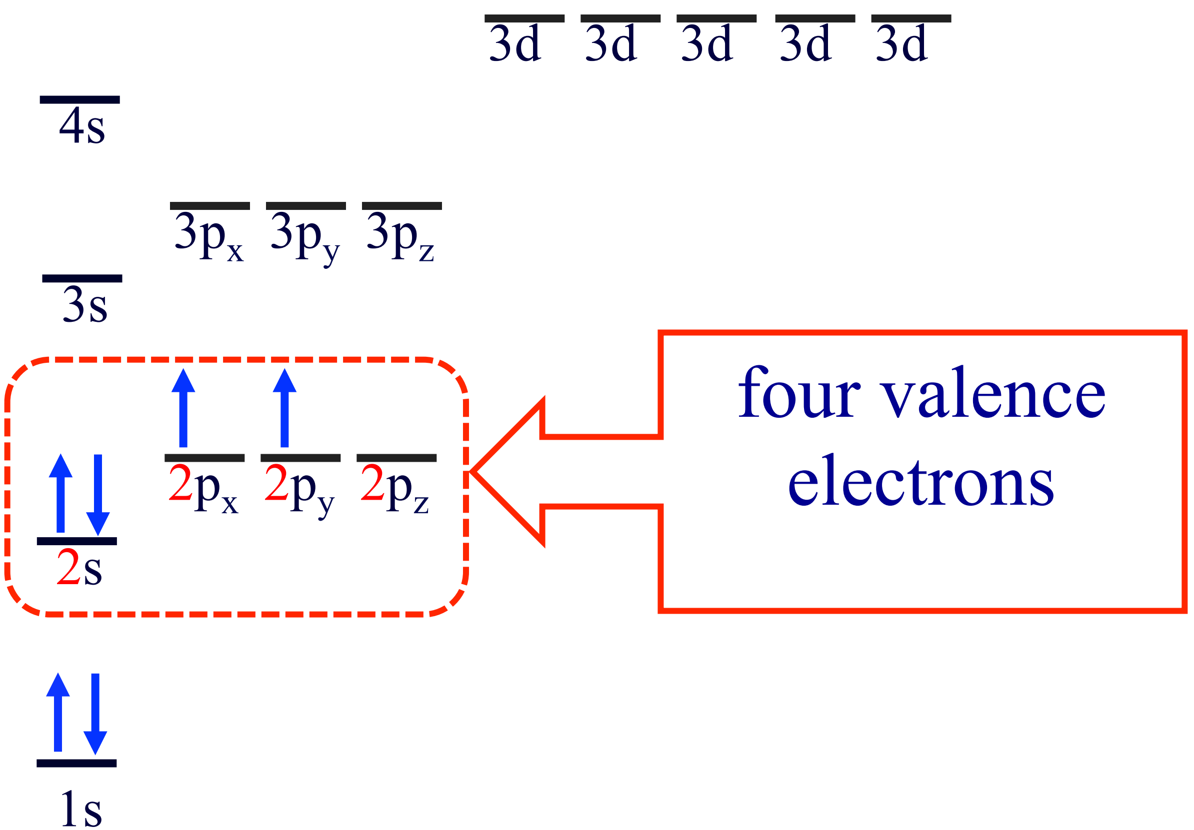
It is important to know how many valence electrons are in an atom because:

These are the electrons that are involved in ***chemical bonding*** to other elements to form ***compounds***.

These are the electrons that elements lose to become ***ions***.

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 ***column*** (group) have the *same number of valence electrons* as all others in that column.

s-Block

p-Block

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

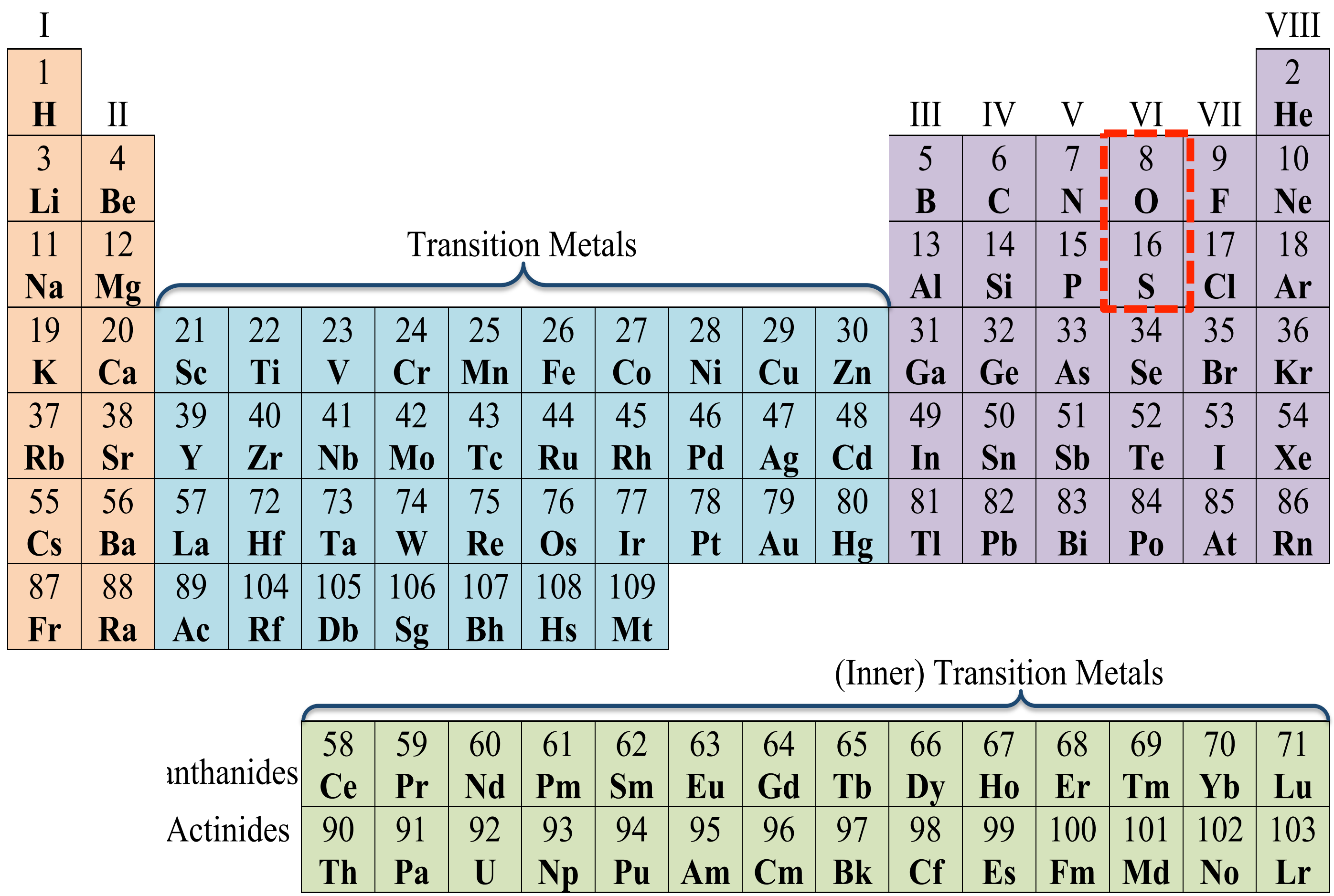
Lanthanides	58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
Actinides	90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

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

Different elements with the same number of valence electrons are said to be ***isoelectric***.

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

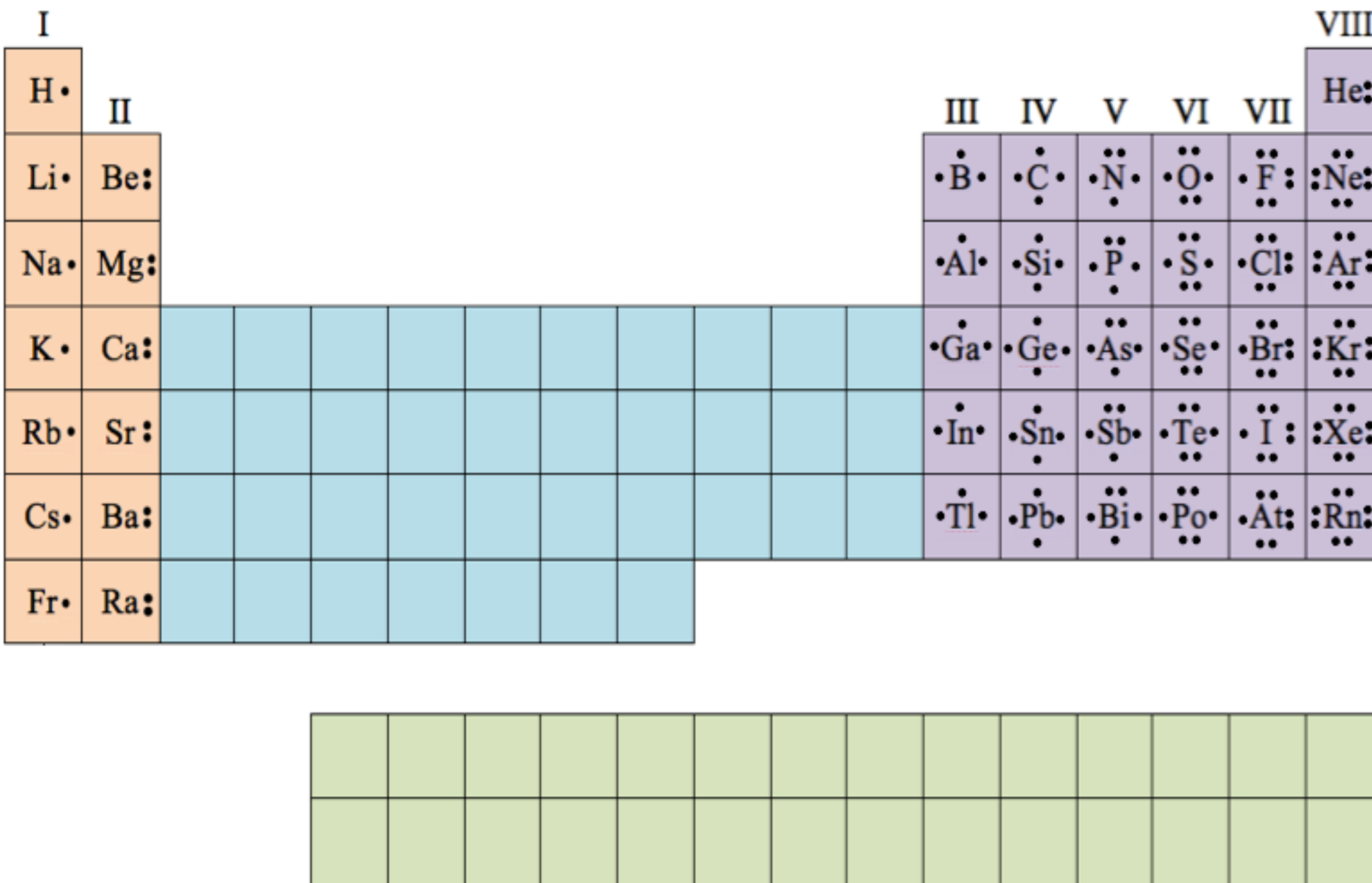
Use the periodic table to determine the number of valence electrons in each of these types of atoms:

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* **<u>dots</u>** drawn next to an element's symbol.

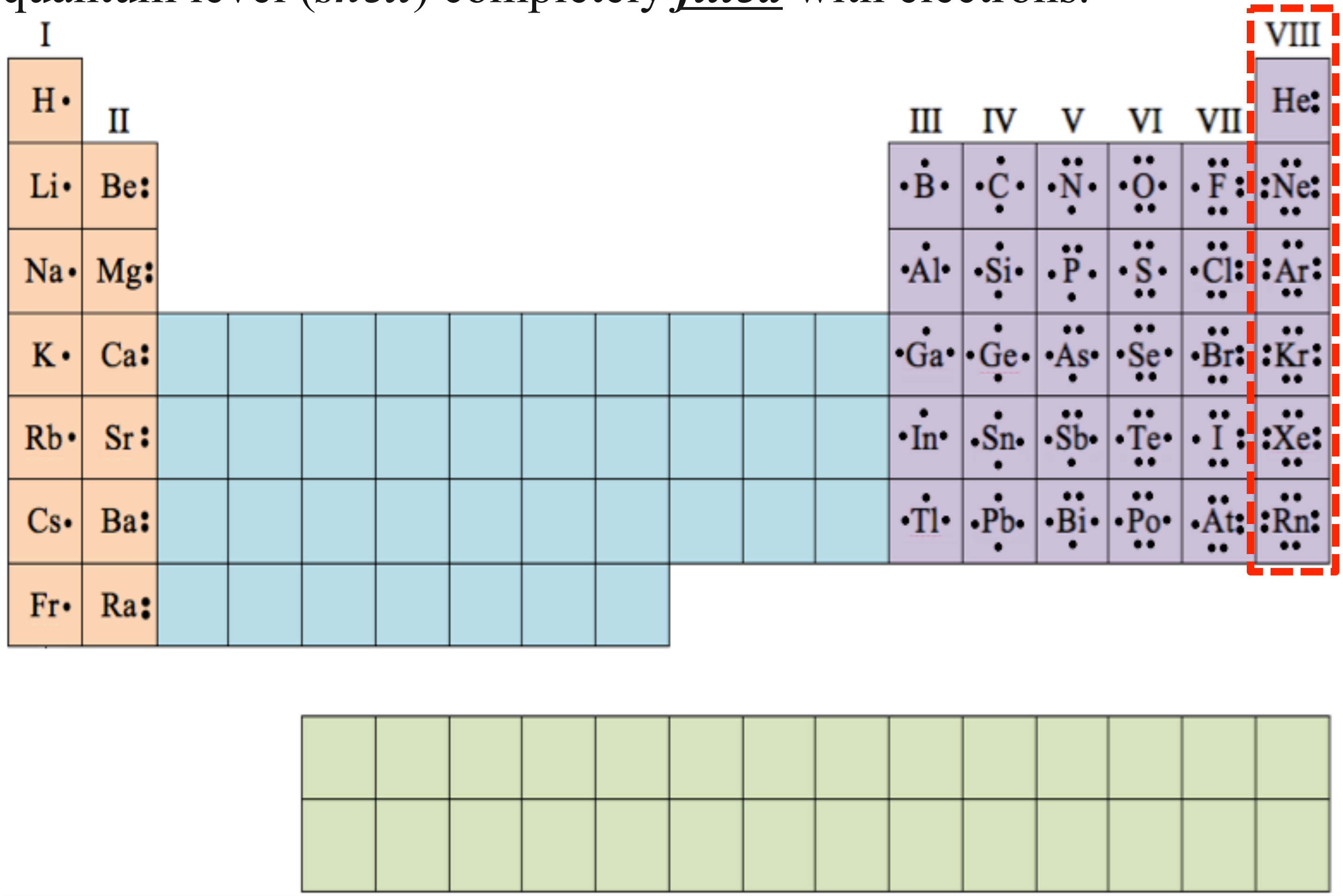


Noble Gases and the Octet Rule

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

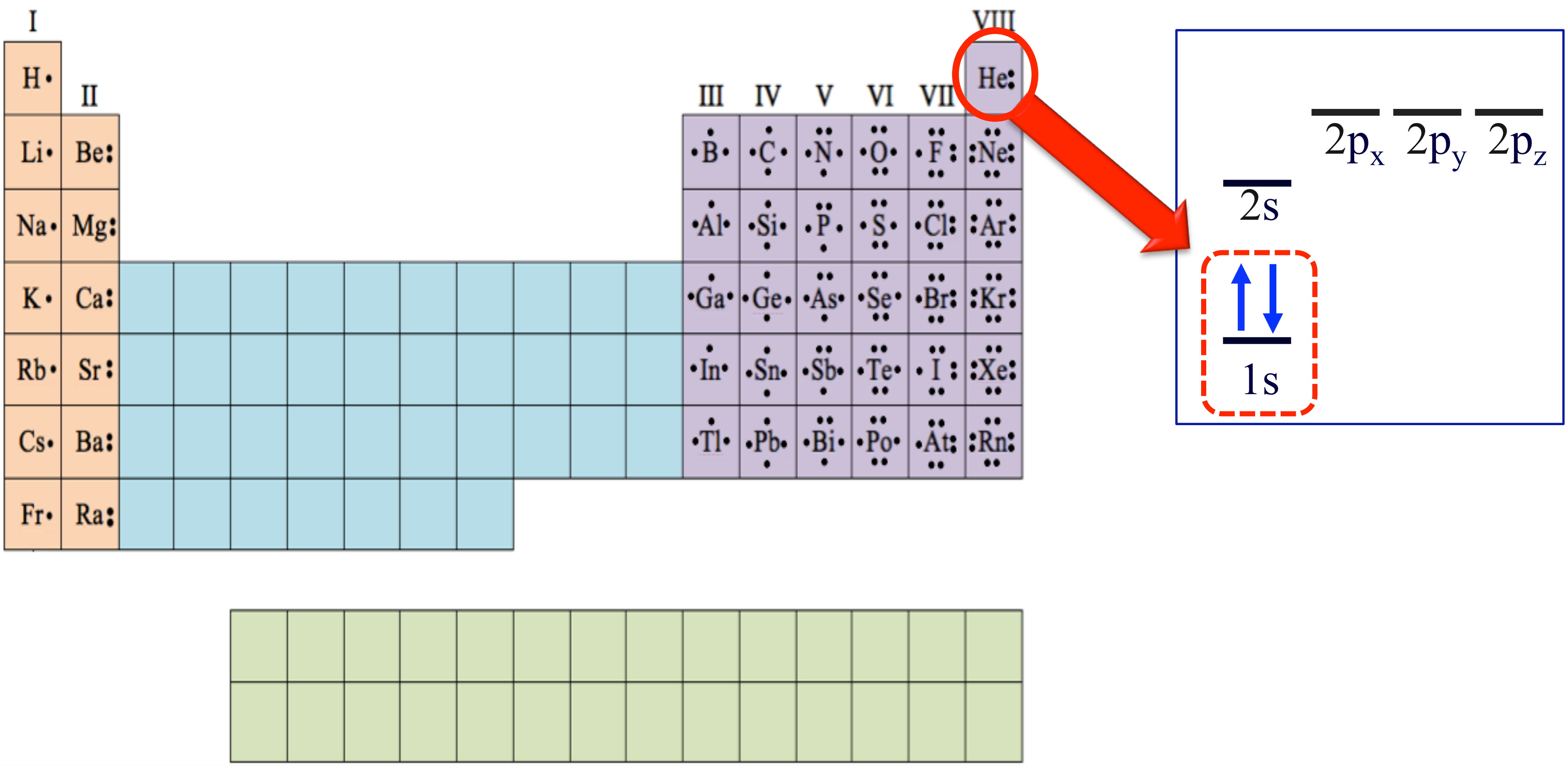
He, Ne, Ar, Kr, Xe, and Rn belong to the **noble gas family**, which gets it's name from the fact that these elements are resistant to change and, with few exceptions, do not lose or gain electrons.

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



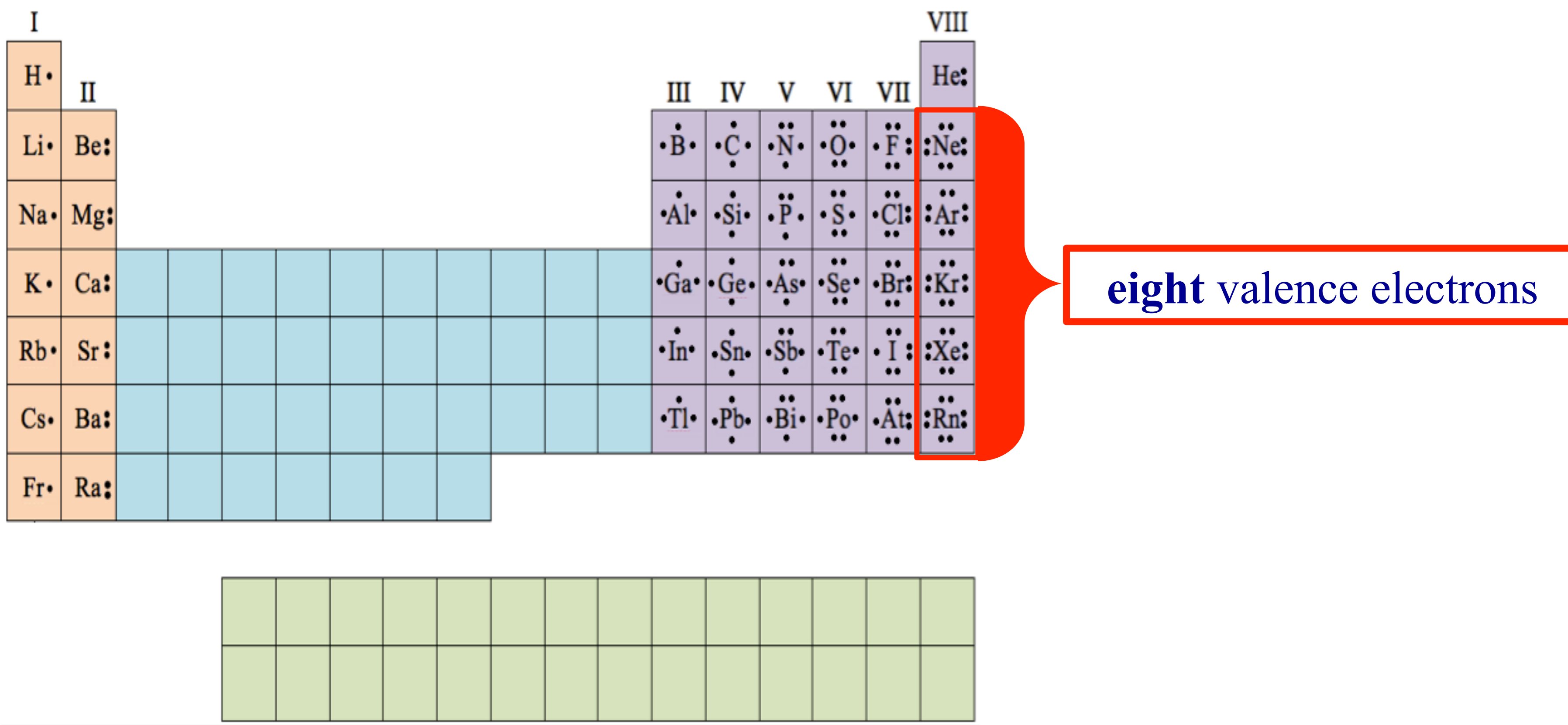
Noble Gases and the Octet Rule

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



Noble Gases and the Octet Rule

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



Noble Gases and the Octet Rule

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

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 ***ion*** is a small particle that has an ***electrical charge***.

Atoms can ***gain or lose electrons*** to become **ions**.

Metal atoms can ***lose*** 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 ***cations***.

Nonmetal atoms can ***gain*** electrons to form ***negative ions***.

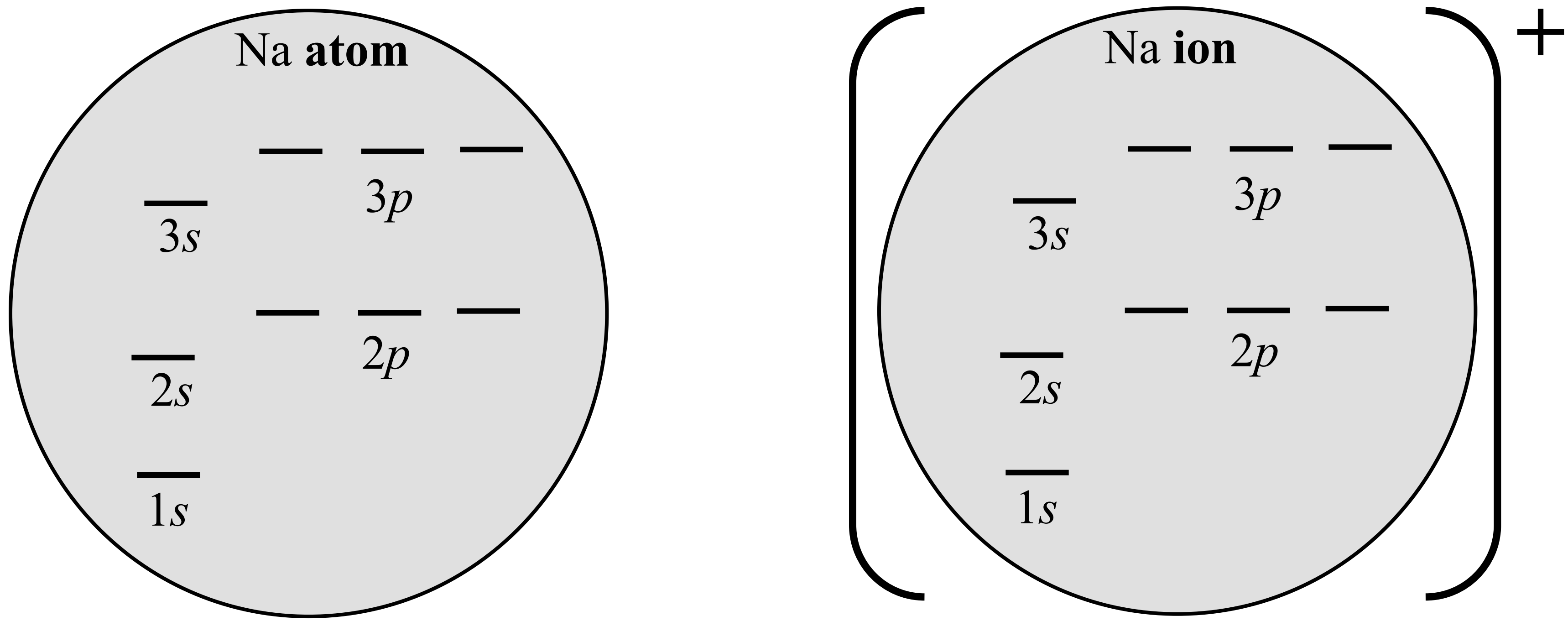
If an atom *gains* one or more electrons, it will then have more electrons than protons and have an overall *negative charge*.

- *Negative ions* are called ***anions***.

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

- A sodium atom has __11__ protons and __11__ electrons.
- How many valence electrons does the sodium atom have? __1__
- How many valence electrons does sodium "want?" __8__

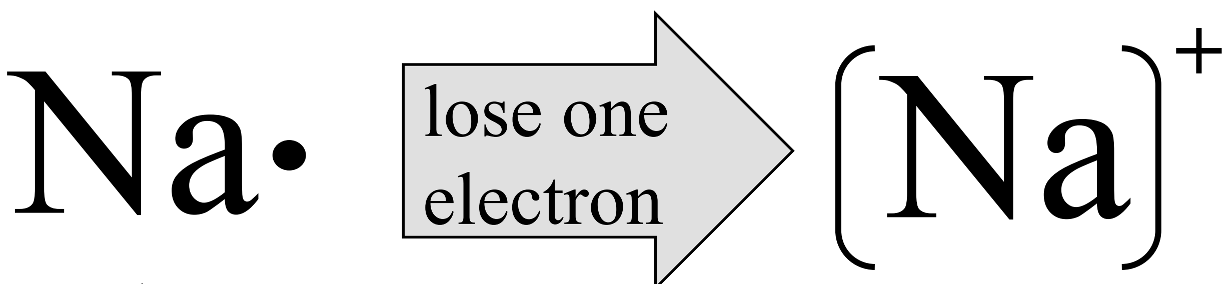
Fill the energy level diagrams with electrons:



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

Sodium will lose ***one*** electron to become a sodium ion (Na^+).

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



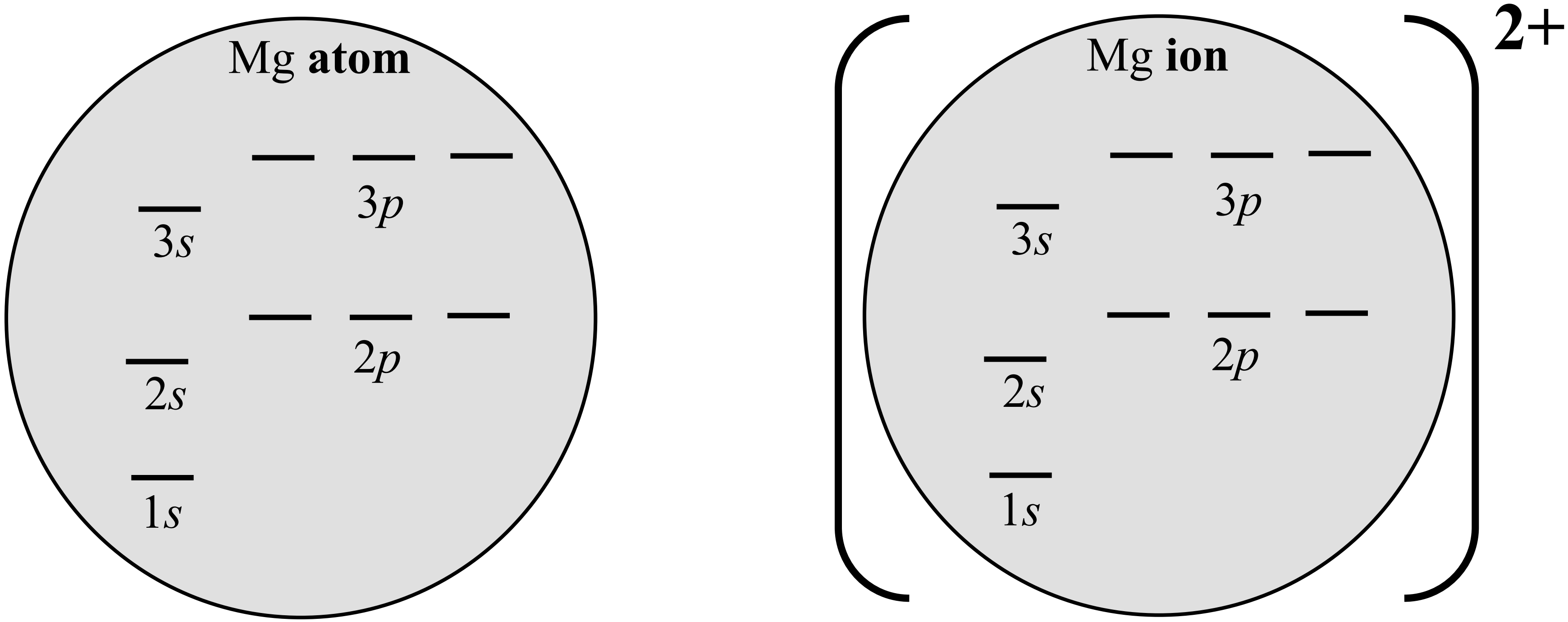
electron dot structure for a **Sodium Atom**

electron dot structure for a **Sodium Ion**

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

- A magnesium atom has __12__ protons and __12__ electrons.
- How many valence electrons does the magnesium atom have? __2__
- How many valence electrons does magnesium "want?" __8__

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 ***two*** electrons to become a magnesium ion (Mg^{2+}).

Understanding Check

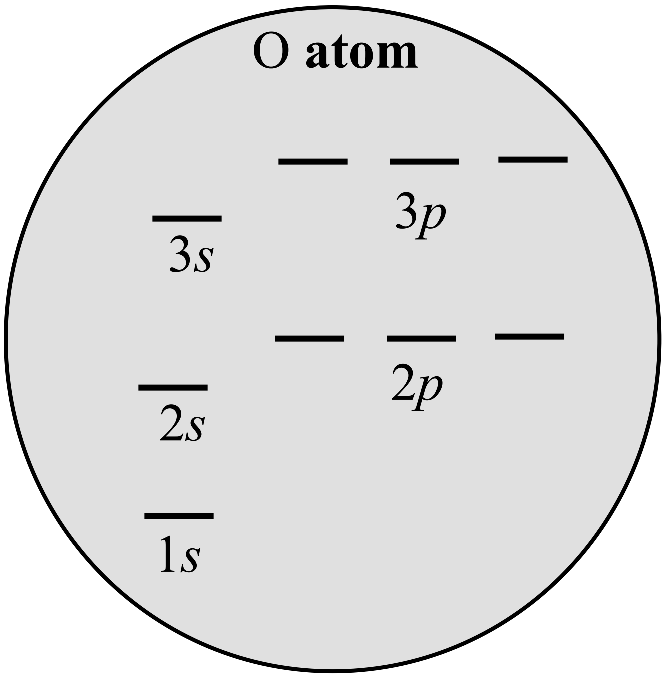
Based on the octet rule, what would be the charge of an aluminum ion?

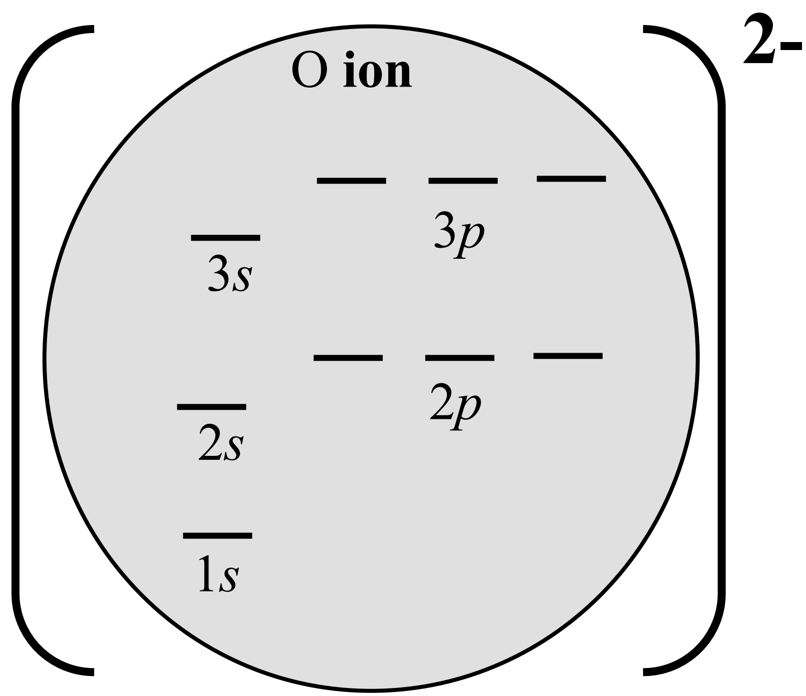
HINT: Begin with the energy level diagram (or the number of valence electrons) for an aluminum atom.

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

- A oxygen atom has __8___ protons and __8__ electrons.
- How many valence electrons does the oxygen atom have? __6__
- How many valence electrons does oxygen "want?" __8__

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 ***two*** 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**:

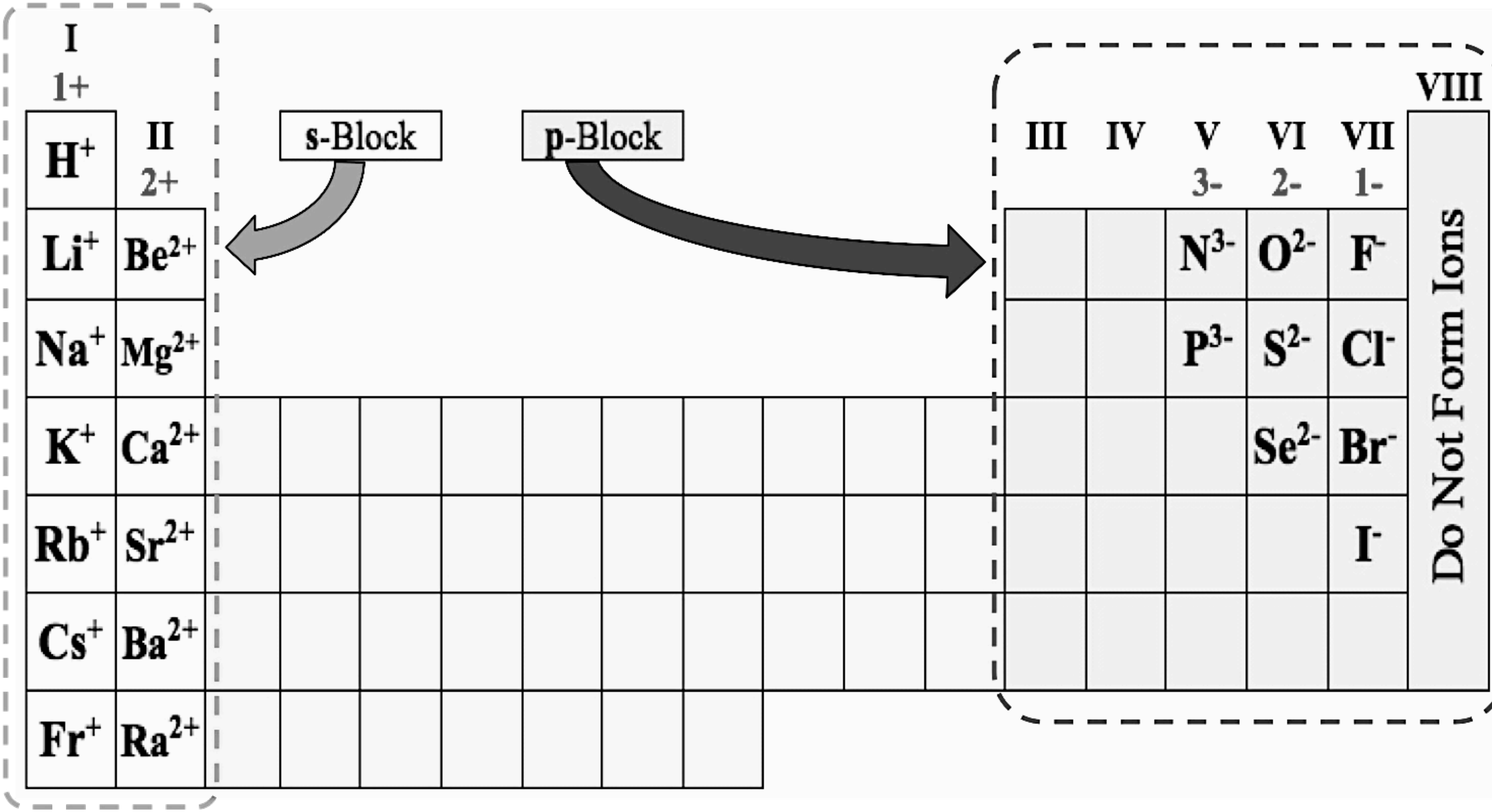
Oxygen has 6 valence electrons, if we add two electrons, its outer shell will have a full octet.

Understanding Check:

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

Begin with the electron dot structure for a chlorine atom.

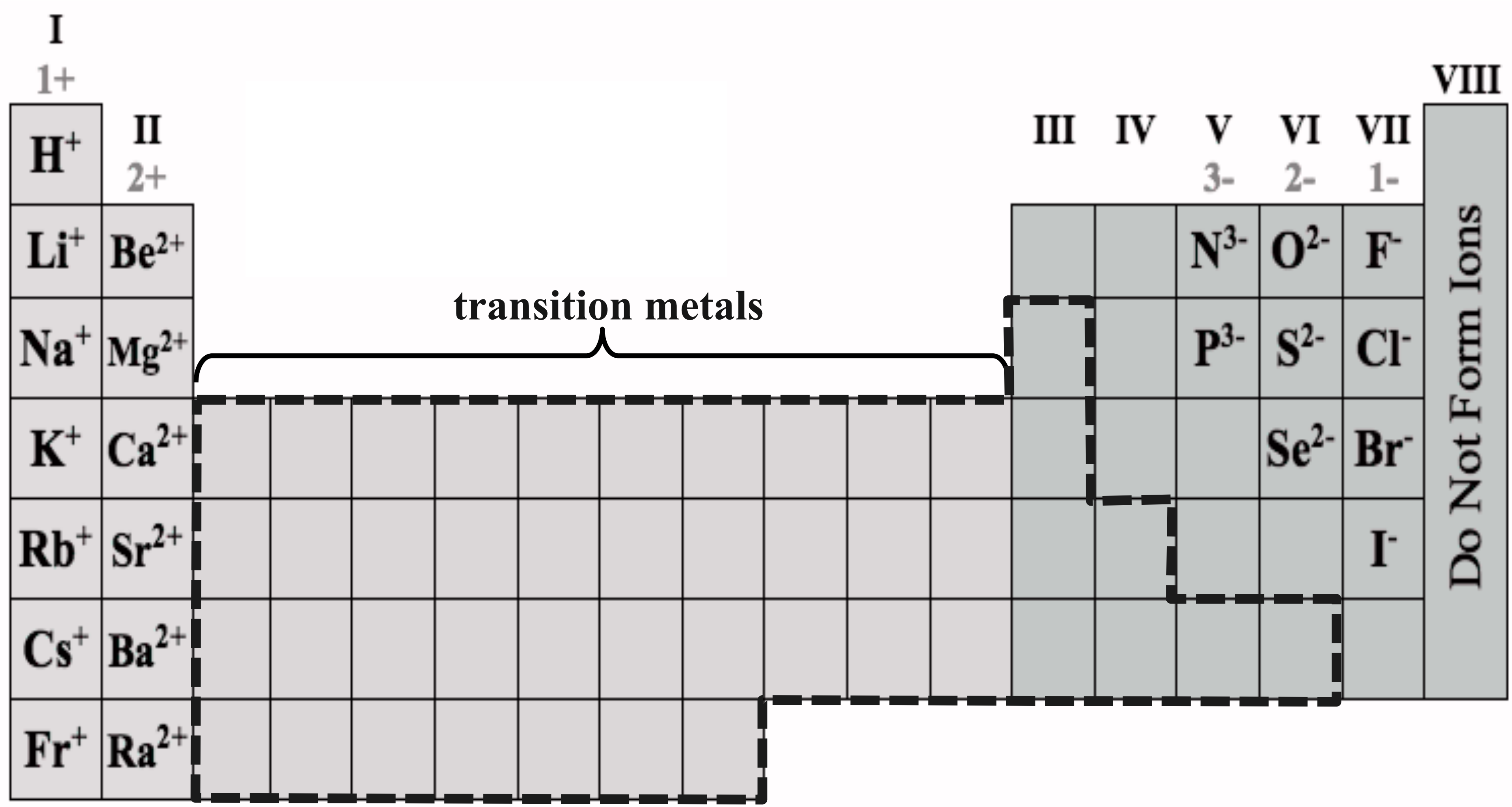
We can determine the charge of an ion formed from ***s-block elements*** and ***p-block nonmetals*** from the number of valence electrons in those elements, and therefore by their ***location*** on the periodic table.



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.	

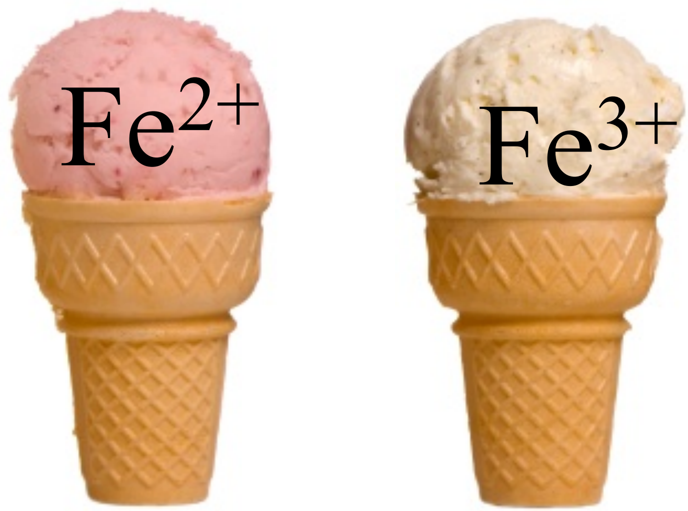
The charge of the ions formed from the **transition metals** and **p-block metals *cannot*** always be predicted by their position in the periodic table.

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 ***<u>Roman</u> 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 (as shown below).

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+}

You do not need to memorize the metal names and charges in this table; I will give you this table for with your exams.

Naming Monatomic Ions

A ***monatomic ion*** is an ion that is made when a ***single*** **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 "**-*ide***."

- **Examples:**

  F^- is referred to as a **fluoride ion**.

  O^{2-} is referred to as an **oxide ion**.

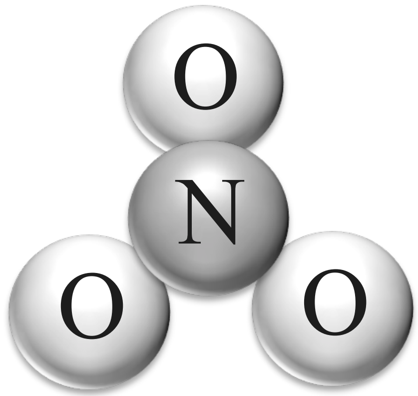
Polyatomic Ions

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

If the resulting particle has the s*ame number of protons as electrons*, then it will be ***electrically neutral***, and we call the particle a ***<u>molecule</u>***.

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 ***<u>polyatomic</u>*** **ion**.

Example of a Polyatomic Ion: Nitrate Ion



Nitrogen	(7 electrons, 7 protons)	
Oxygen	(8 electrons, 8 protons)	
Oxygen	(8 electrons, 8 protons)	Nitrate Ion NO_3^-
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	
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

An Introduction to Compounds

Compounds: matter that is constructed of two or more chemically bonded elements.

Each compound has the same proportion 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 <u>forces</u>*** that hold atoms or ions together in a compound.

Chemical Bonds

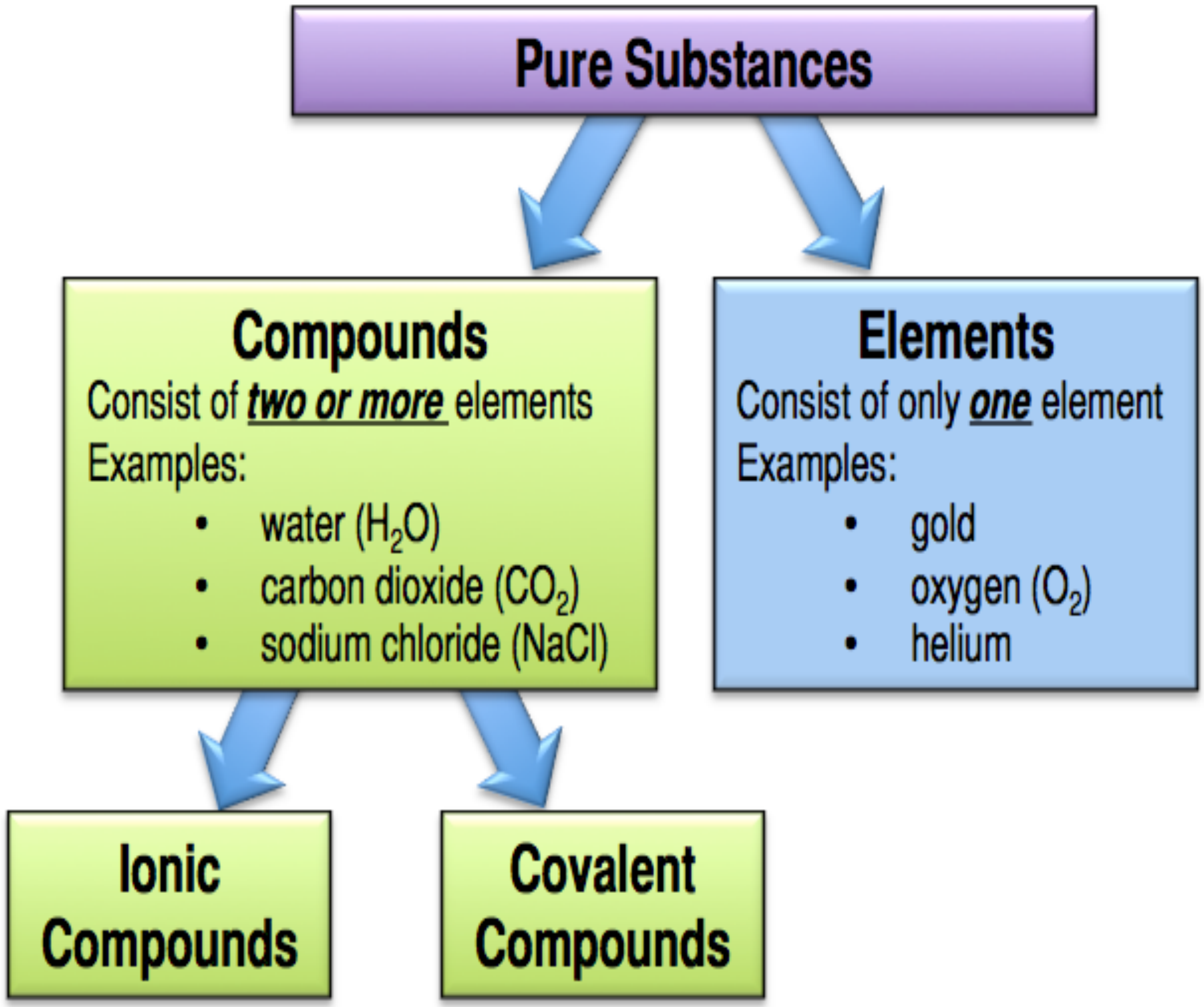
There are *three types* of **chemical bonding**:

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*.

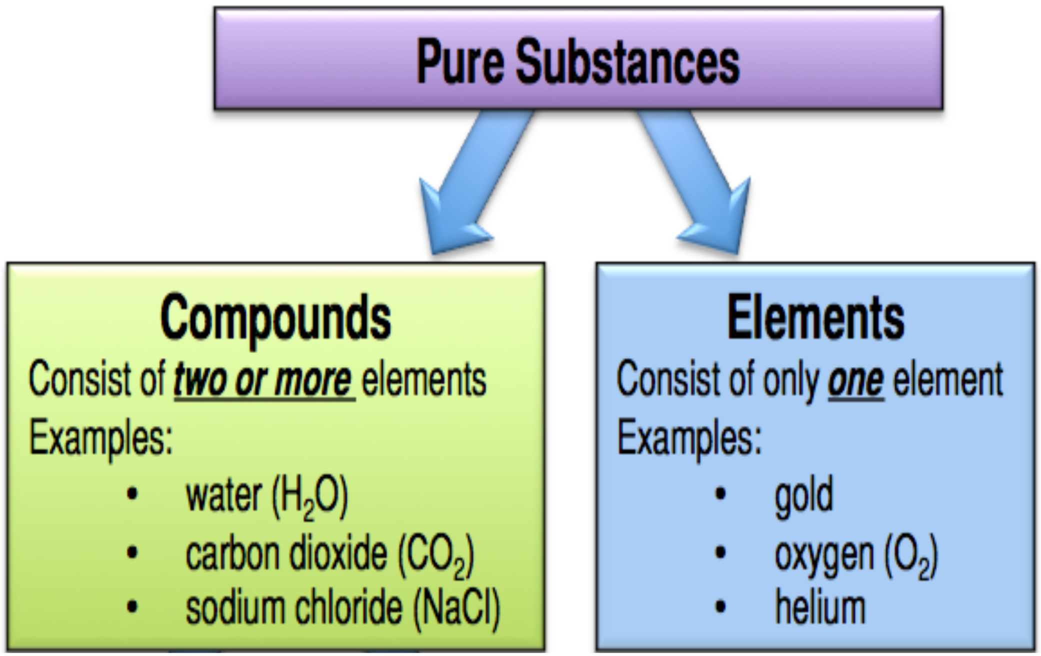
You will learn about *metallic bonding* in chapter 5.

Some Terminology



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

Physical* 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.

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

- To make a **new** pure substance, ***chemical bonds must be broken and/or new chemical bonds are made***.

- 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 ***sharing of electron pair(s)*** between two atoms.

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

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

- Contrast this with a polyatomic ion, which is an electrically ***charged*** group of atoms ***held together by covalent bonds***.

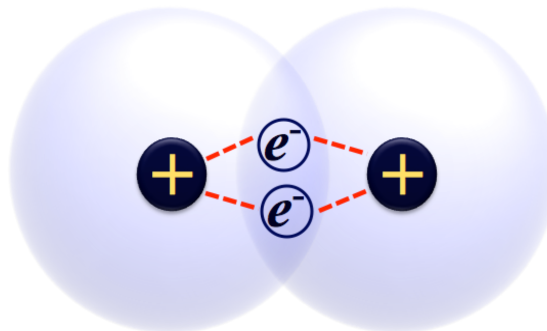
Covalent bonding occurs between ***nonmetal*** atoms.

Formation of a Covalent Bond

Covalent bonding occurs because the bound atoms are at a ***lower*** 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 ***share*** *electron pairs*.

Each hydrogen atom provides one electron in the shared pair.

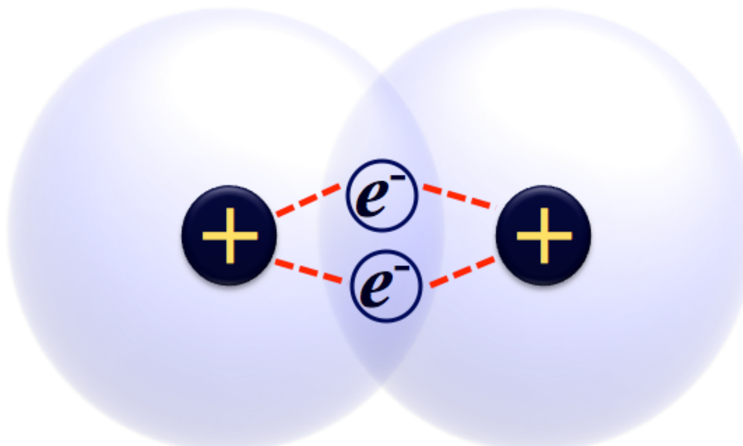
The shared electron pair spends significantly more time in the area between the positive nuclei of the hydrogen atoms than in other regions.

The electron pair between the nuclei create a positive-negative-positive electrostatic attractive "sandwich" and this ***force*** holds the atoms together.

- The dashed lines indicate the electrostatic attractive interactions.

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.

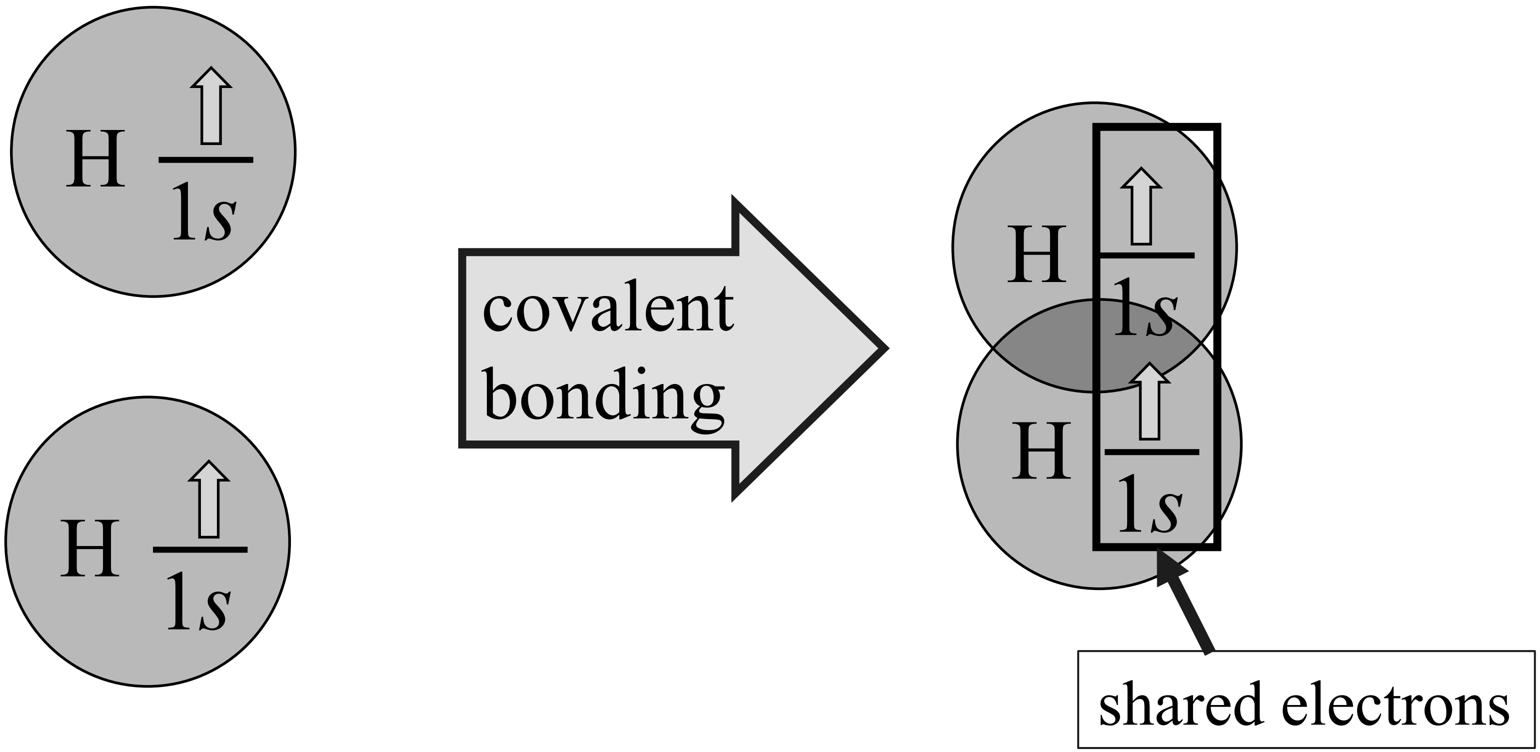


The octet rule in the formation of molecules is: ***molecules tend to form such that the atoms are surrounded by an octet (eight) of valence electrons (except for hydrogen and helium that have two electrons).***

The Octet Rule in the Formation of Molecules

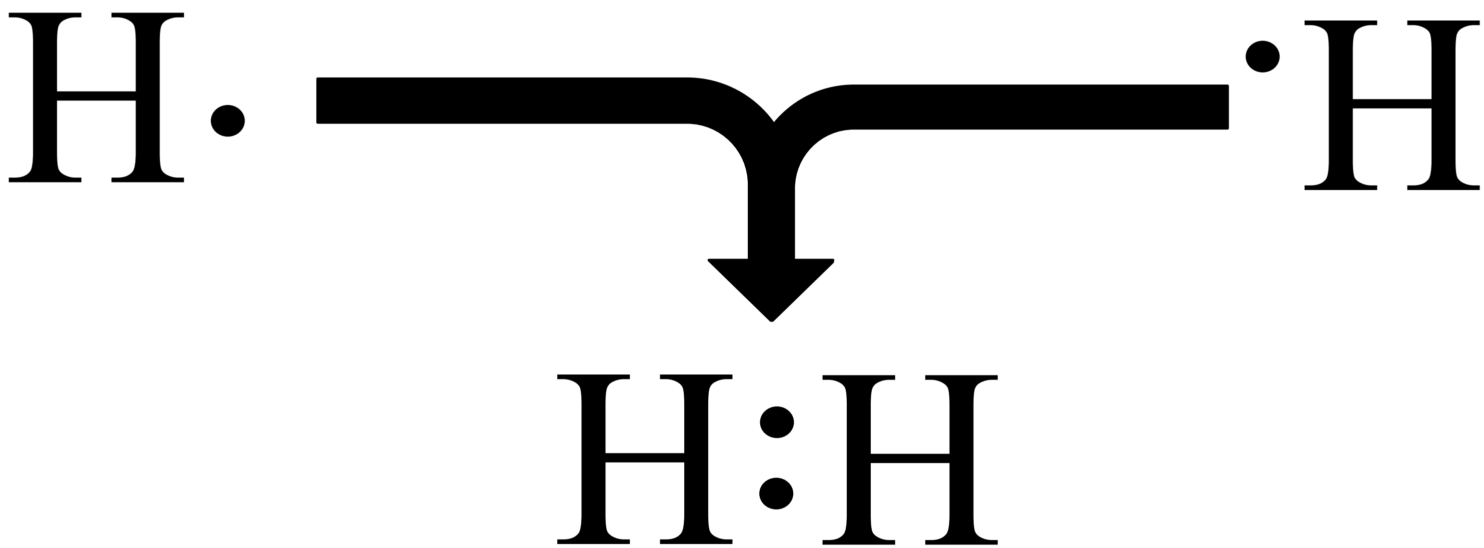
Example: $\mathbf{H_2}$

(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 ***__between__*** the atoms are shared in a covalent bond.

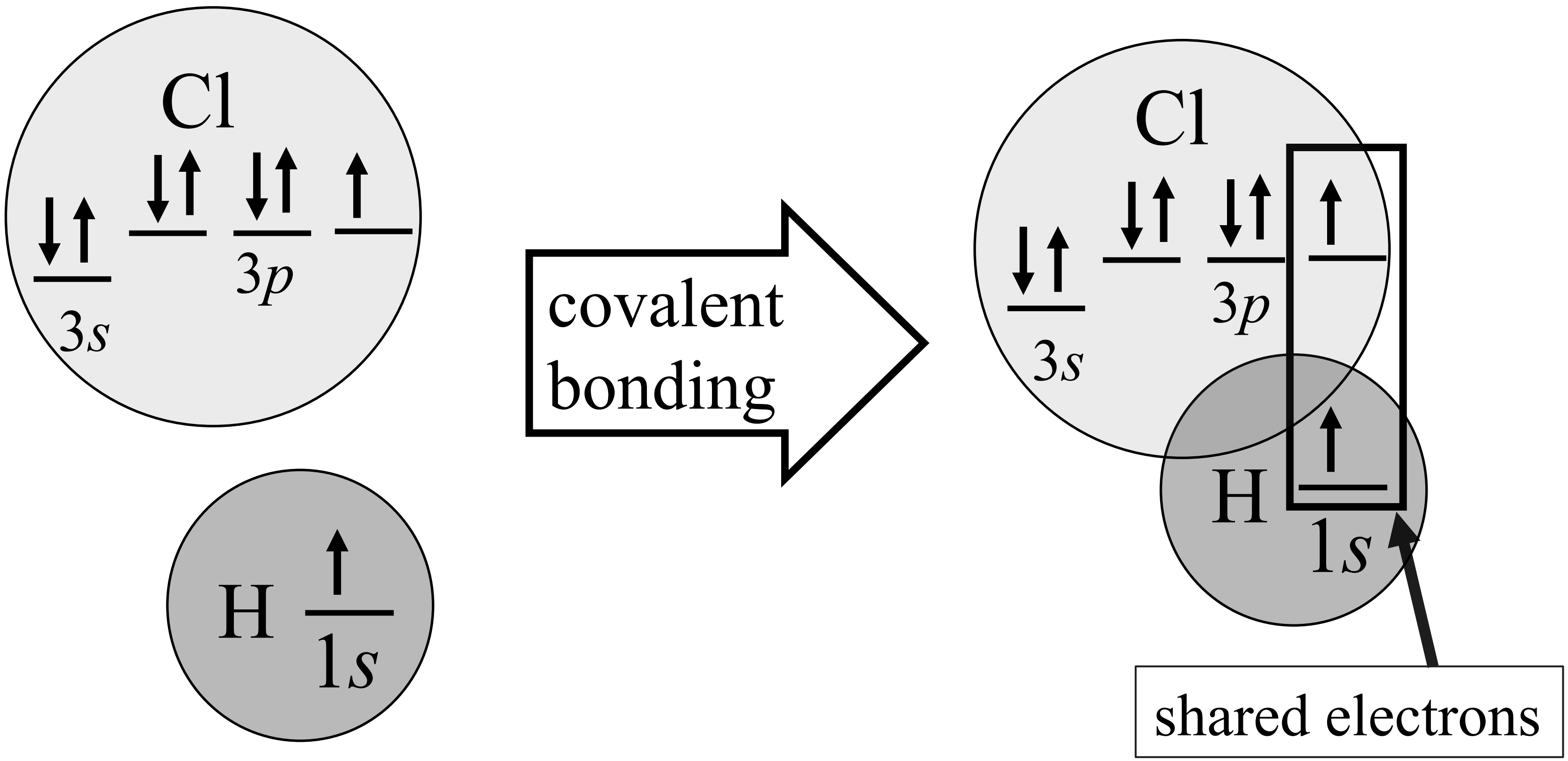
Chemist use a line to represent **__2__** electrons in a covalent bond.

These drawings are called ***__line bond structures__***.

H—H

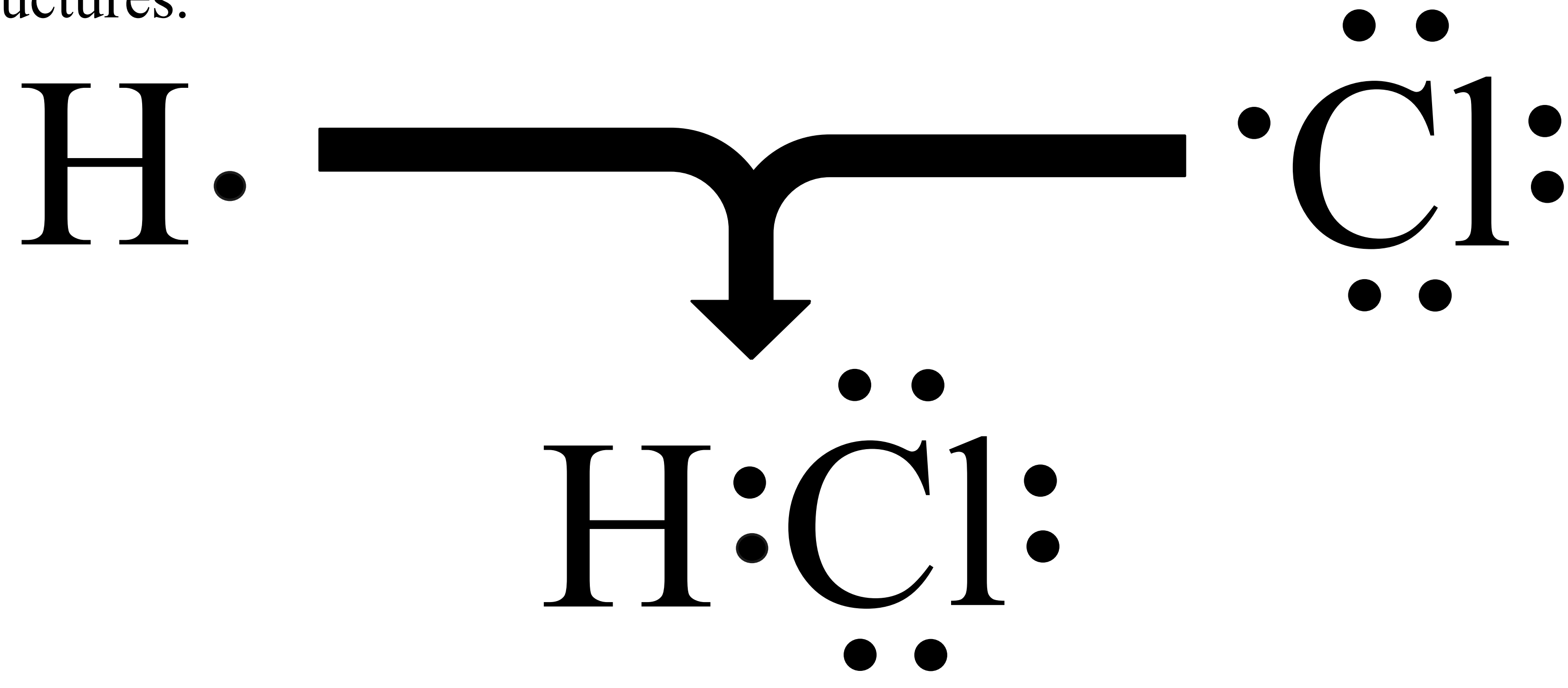
The Octet Rule in the Formation of Molecules

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

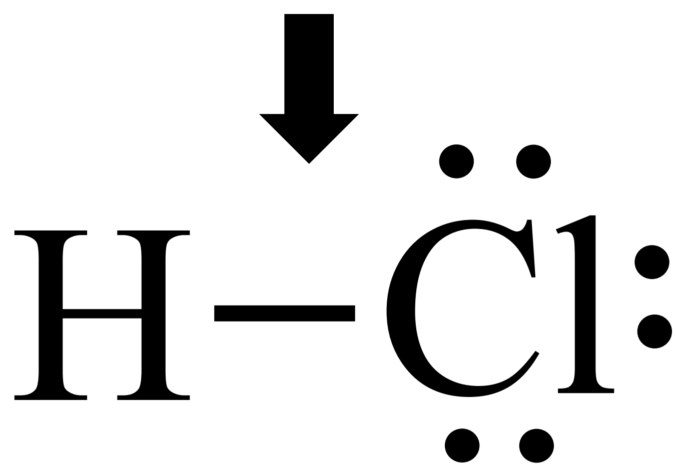


When a covalent bond forms, the hydrogen atom "feels" **two** electrons in its outermost shell, and the chlorine atom "feels" **eight** electrons in its outermost shell.

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

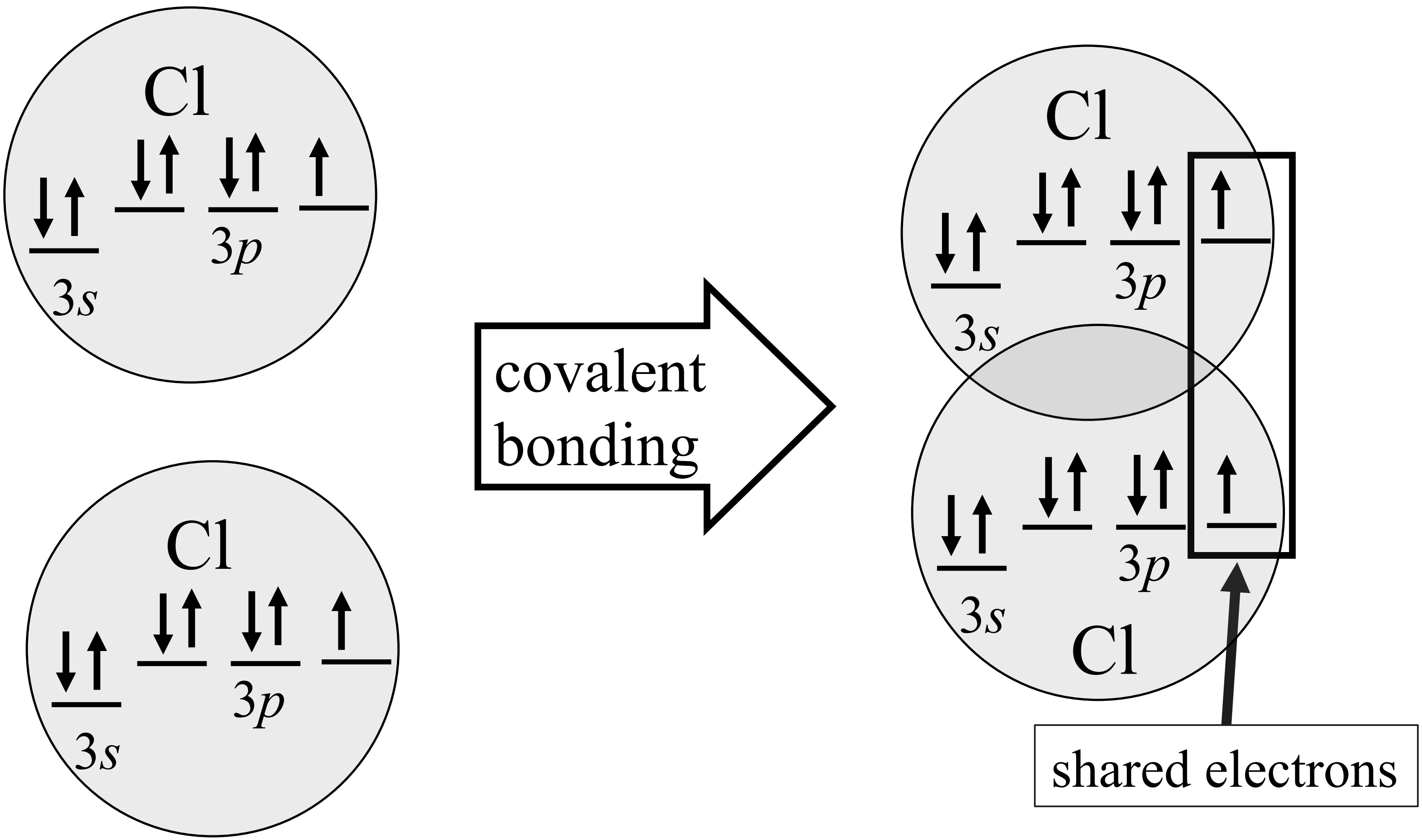


line bond structure



The Octet Rule in the Formation of Molecules

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



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

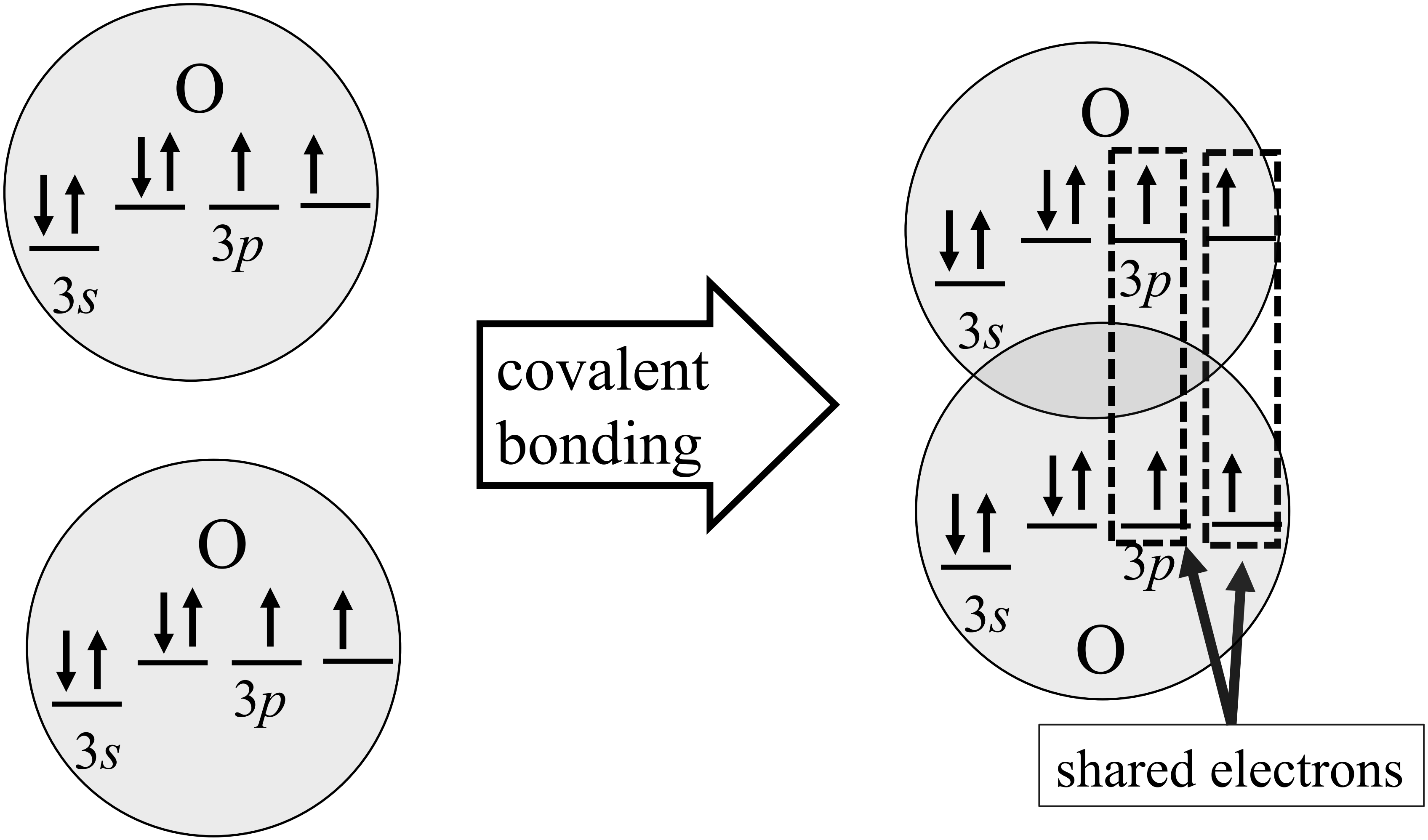
You try it:

Draw the ***line bond structure*** for Cl_2.

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

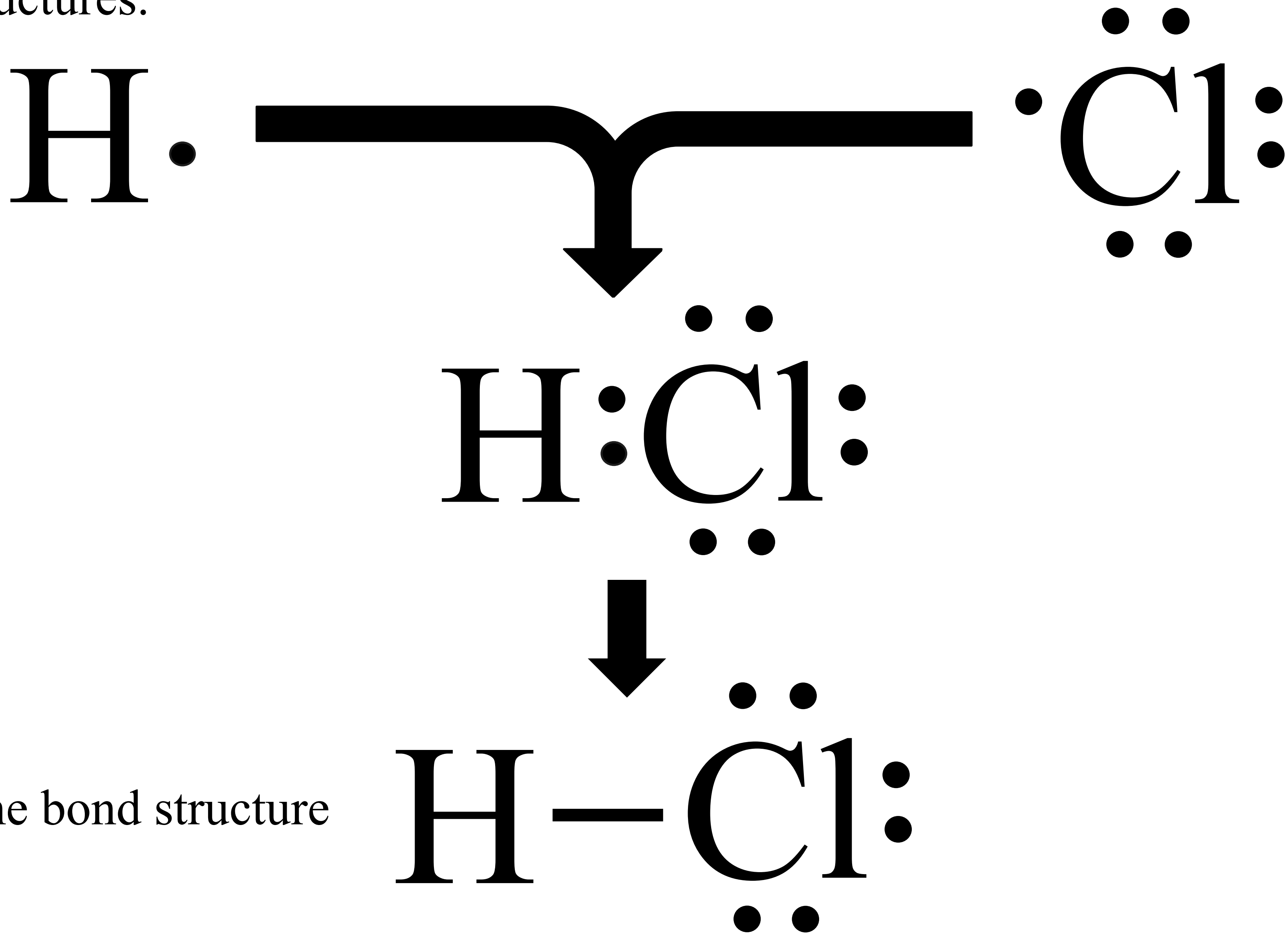
The Octet Rule in the Formation of Molecules

Let's do **oxygen gas (O_2)**.



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

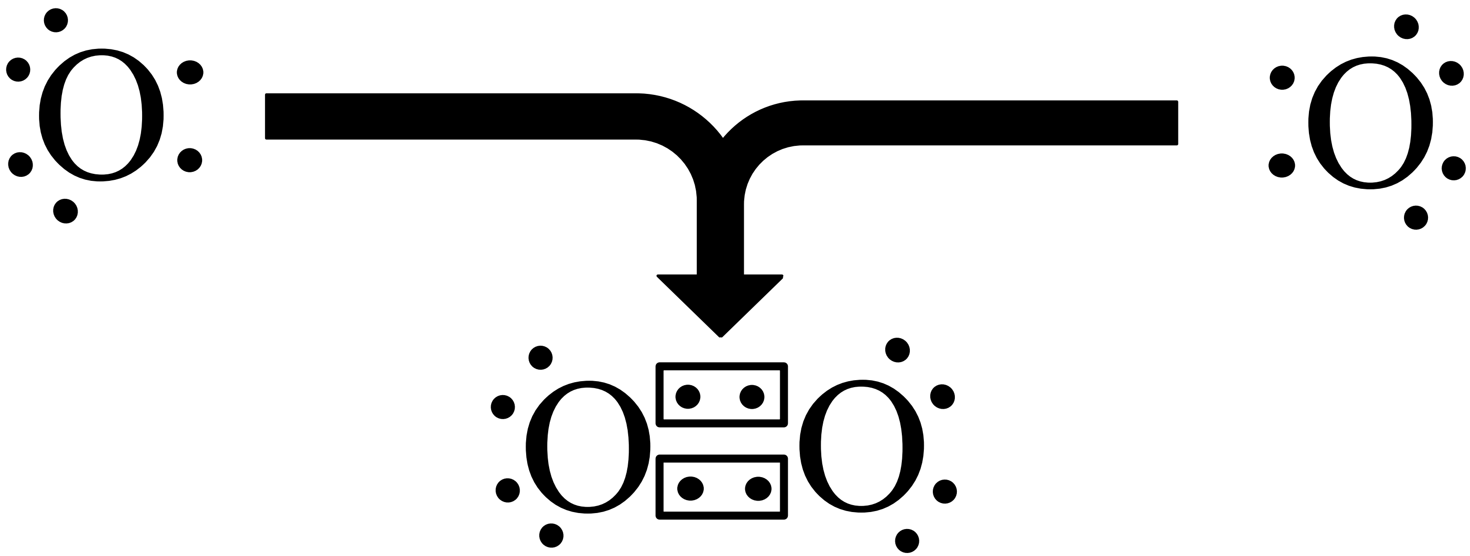
The HCl covalent bond can also be illustrated using electron dot structures.



line bond structure

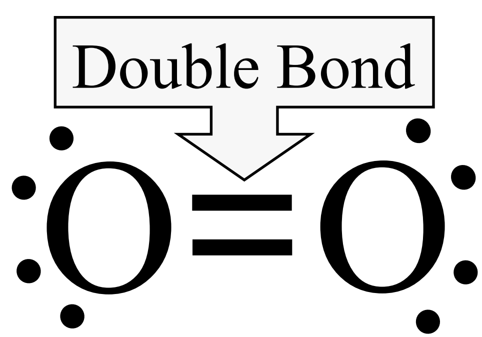
Let’s draw the line bond structure for oxygen gas (**O_2**).

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



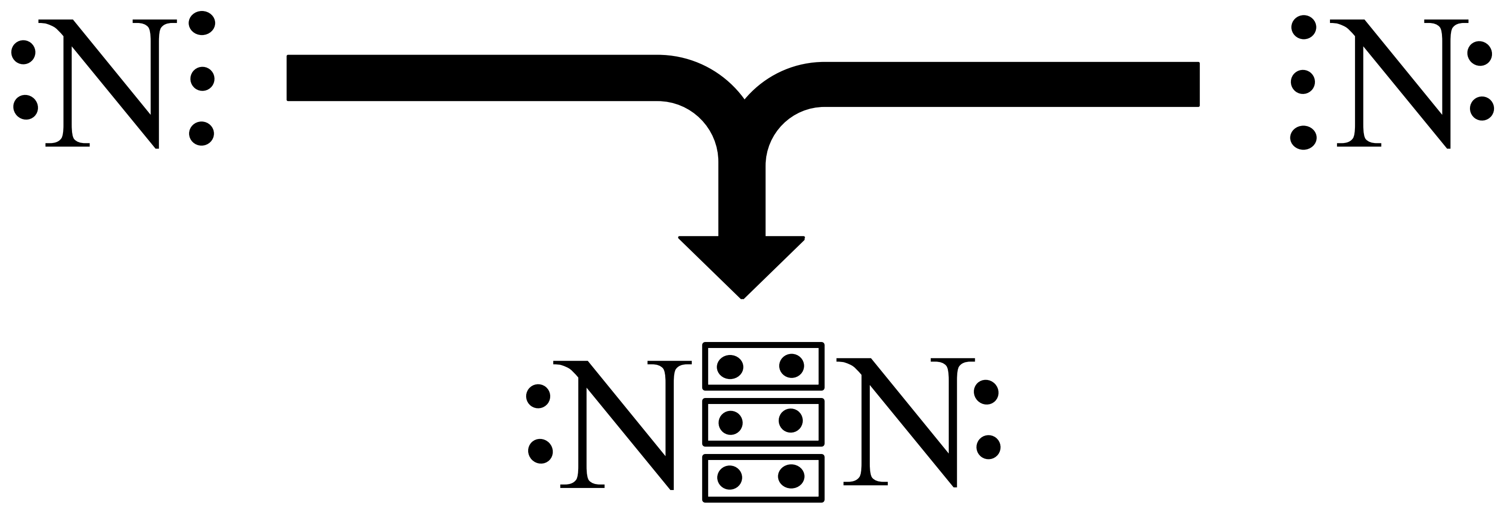
We use lines to represent ***shared*** electron pairs.

When atoms are bonded with **2 pairs** of electrons it is called a ***double bond***.



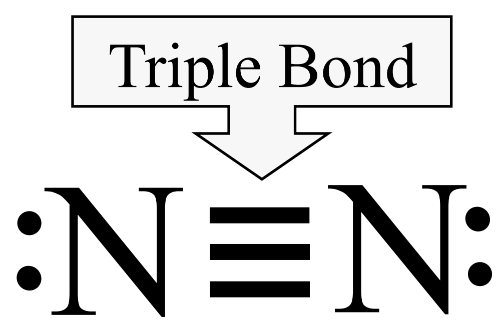
Let's draw the line bond structure for nitrogen gas (**N_2**)

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



We use lines to represent *shared* electron pairs.

When atoms are bonded with **3 pairs** of electrons it is called a ***triple bond***.



Naming Binary Covalent Compounds

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

1 **H**			Metals (Green)			Nonmetals (Blue)			Metalloids (Red)								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**									
				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**
				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**

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

Covalent Bonding: Molecular Compounds

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

Covalent Bonding: Molecular Compounds

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

When discussing molecules we use a ***molecular formula*** 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–Ö–H (with two lone pairs on O)	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

Binary **covalent compounds** contain only ***two elements*** (the "**bi-**" prefix indicates "***two***").

- Examples of binary covalent compounds are HCl, H_2O, and 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

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

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

- **Exception:** If there is just one atom of the ***first*** element in the formula, do not use **mono-** for the ***first element in the name***.
  - Example: 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: Name the following compound CCl_4

- 1) List the name of the first element in the formula.
- 2) List the second element and add the –ide suffix.
- 3) Use Greek prefixes to indicate the number of each atom in the formula.
  - Exception: do not use mono- for the first element in the name.

~~mono~~carbon tetrachloride

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

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

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 <u>tetra</u>fluoride**.

$$N_2F_4$$

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*:

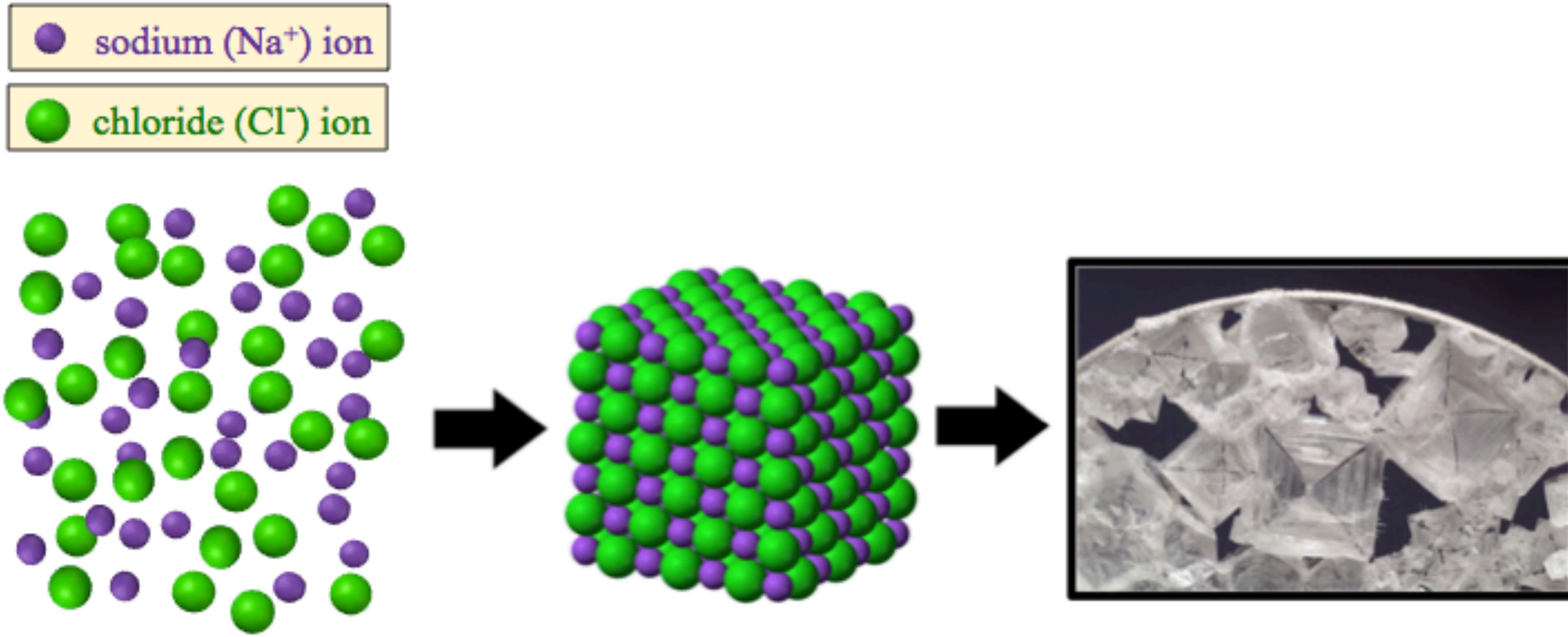
- Glucose ($C_6H_{12}O_6$)
- Acetone (C_3H_6O)

Ionic Bonding

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

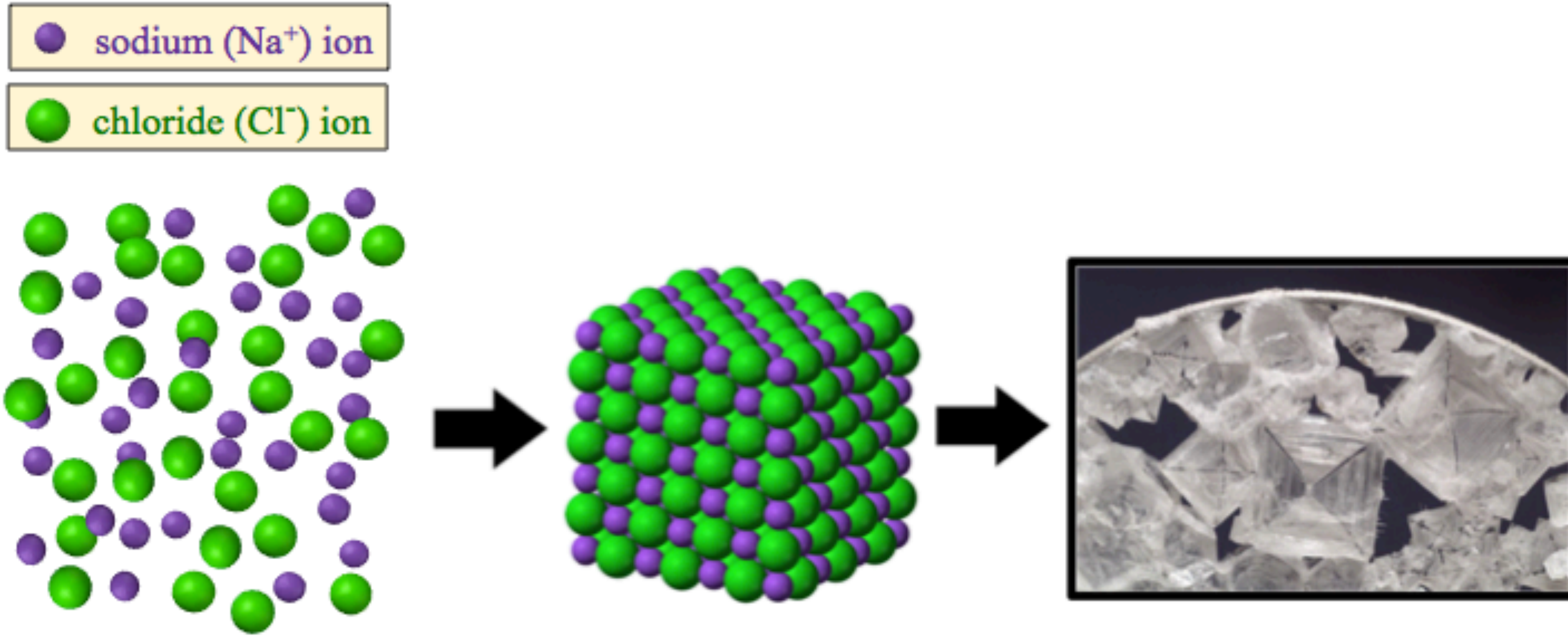
- Compounds composed of ions are called **ionic compounds**.

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.

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



We call this structure a ***crystal*** or ***crystal lattice***.

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.

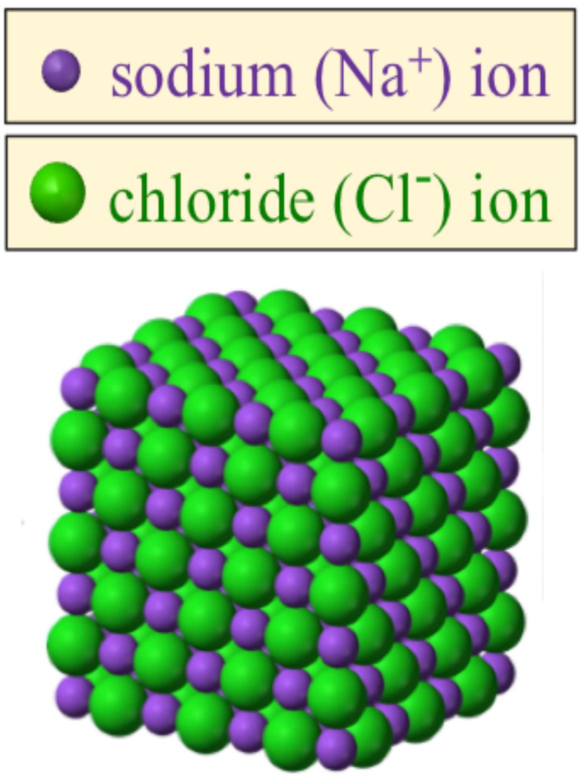
Ionic bonding (ionic compounds) results from:

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

Ionic Compounds

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

- 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 ratio in the crystal

For every sodium ion, there is one chloride ion!

The charges add 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 ***formula units*** **(**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**.

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

Calcium ions have a charge of 2+

Fluoride ions have a charge of 1-

They combine in a ***1-to-2*** ratio in the crystal

For *every* calcium ion, there are ***two*** 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.

$Mg_{\boxed{?}}Cl_{\boxed{?}}$

Understanding Check:

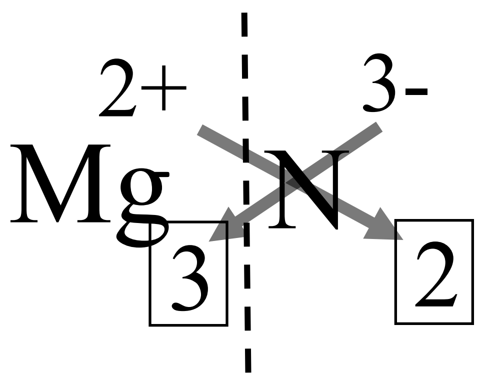
Write the ***formula unit*** for the compound formed by combining potassium and oxide ions.

$K_{?}O_{?}$

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

$Mg_{\boxed{?}}N_{\boxed{?}}$

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-}

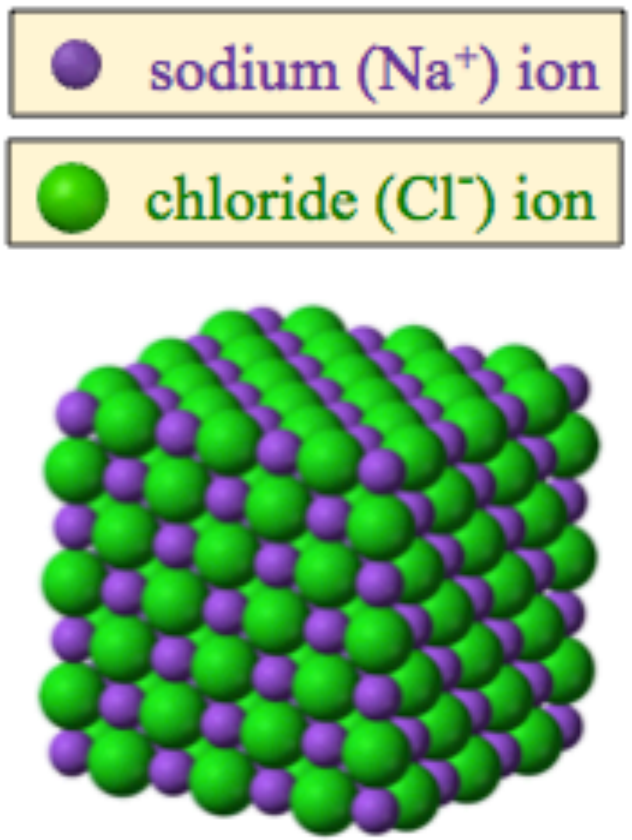
Sn^{4+} and S^{2-}

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

Formula Unit vs. Molecular Formula

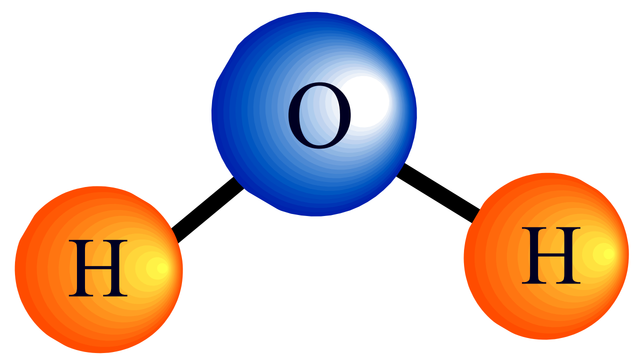
Formula Unit =
Lowest **RATIO** of ions

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**.

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

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.

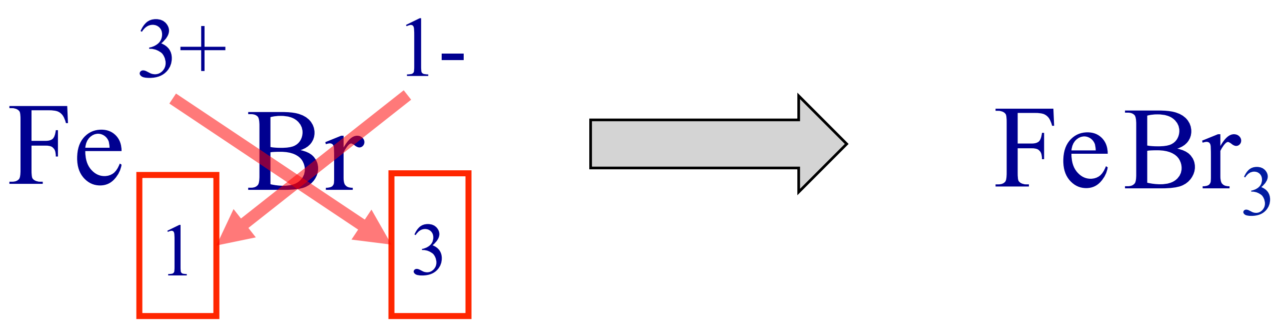
For ***polyatomic ions***:
When the ***subscript*** for a **polyatomic ion** is ***greater than 1***, the polyatomic ion formula is written in parenthesis and the subscript is written after/outside of the parenthesis.

Example:

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

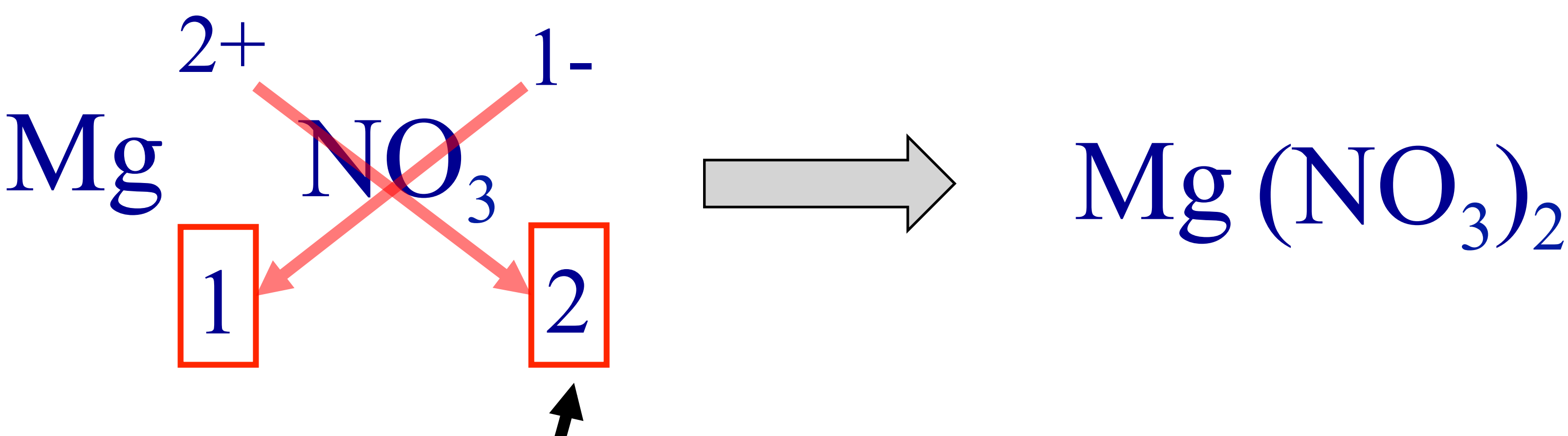
You will know the **charge** for the metals t*hat occur with various charges* because the charge will be written in the compound's name in **Roman numerals**.

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

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

Method for Writing the Names of Ionic Compounds

1. Write the ***cation*** name first, then the ***anion*** 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:

$MgCl_2$

Name the metal ion first.

Name the anion next.

magnesium chloride

Example:

Name the following compound:

$CuBr_2$

Name the metal ion first.

Name the anion next.

magnesium chloride

Example:

Name the following compound:

$$\overset{2+}{Cu}\overset{1-}{Br_2}$$

Name the metal ion first

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

Name the anion next.

What is the charge of the bromide ion?

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	__________________

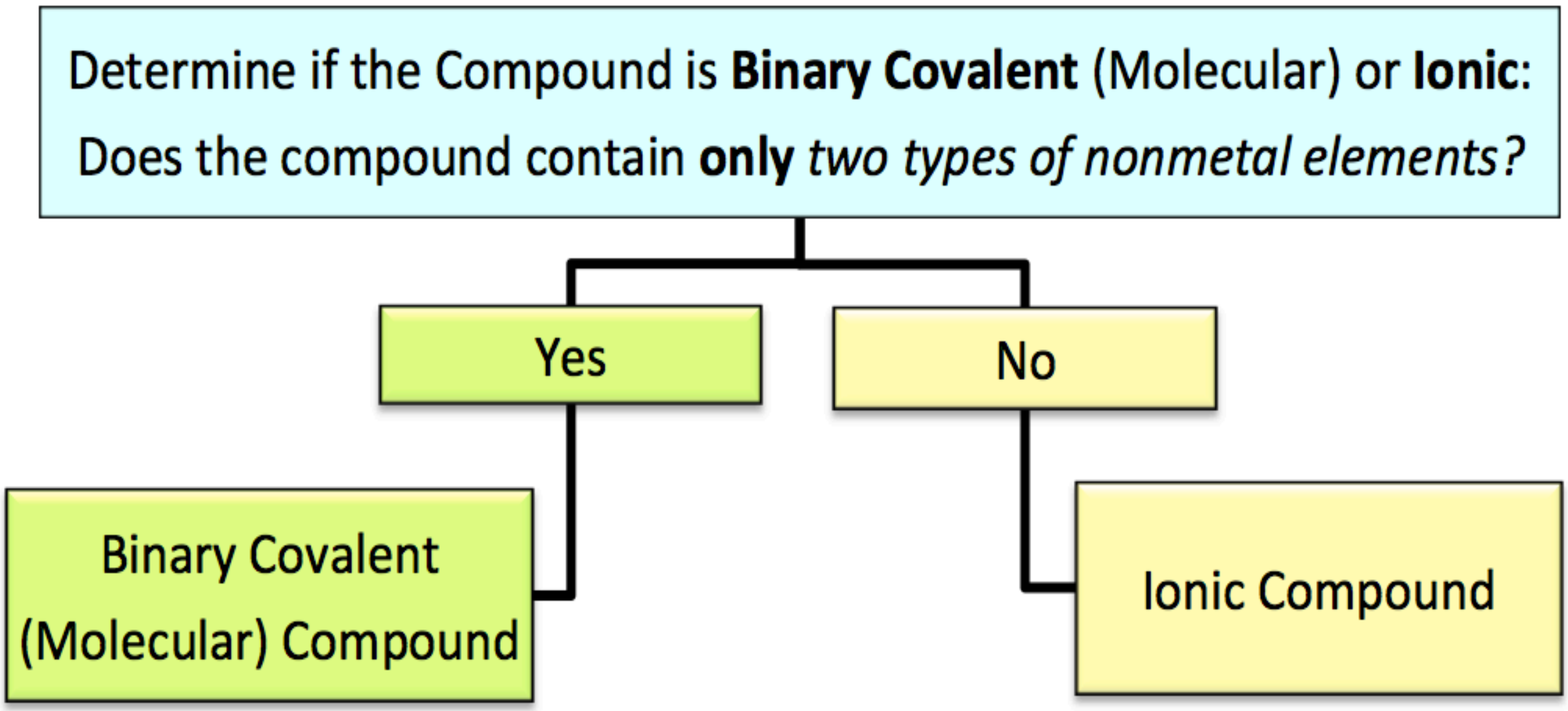
Name the following ionic compounds

NaCl

ZnI_2

Al_2O_3

Naming Compound Summary



Given the Molecular Formula, Write the Name

Determine if the Compound is **Binary Covalent (Molecular)** or **Ionic**:

Does the compound contain **only** *two types of nonmetal elements*?

Yes

No

Binary Covalent (Molecular) Compound:

1) List the name of the first element in the formula.
2) List the second element and add the "ide" suffix.
3) Use Greek prefixes to indicate the number of each atom in the formula.
   - Exception: **Do not use** mono- for the first element in the name.
   - The ***o*** or ***a*** at the end of the Greek prefix is dropped when the element name begins with a vowel.

Ionic Compound:

1) Write the cation name first, then name the anion.
   - Monoatomic anions use the "ide" suffix
2) If the cation is one of the metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.

Given the Name, Write the Molecular Formula

Determine if the Compound is **Binary Covalent (Molecular)** or **Ionic**:

Does the compound contain **only** ***<u>two</u>*** types of ***nonmetal*** *elements*?

Yes

No

Binary Covalent (Molecular) 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 molecule's name in Greek prefixes
    - NOTE: If there is no Greek prefix in front of the first element in the name that implies the number is 1.

Ionic Compound:

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion 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.
    - *IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. Example* $Mg(NO_3)_2$

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**

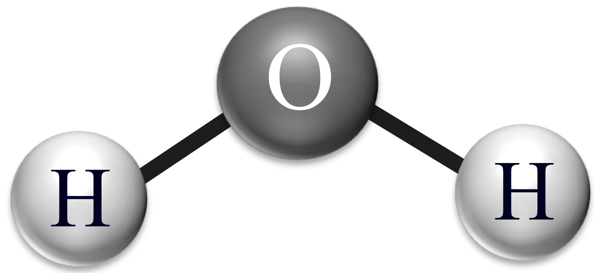
Molar Mass of Covalent Compounds (Molecules)

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

- The *molar mas*s 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** ***atoms*** in the molecule.

Example: Let's calculate the molar mass of H_2O.

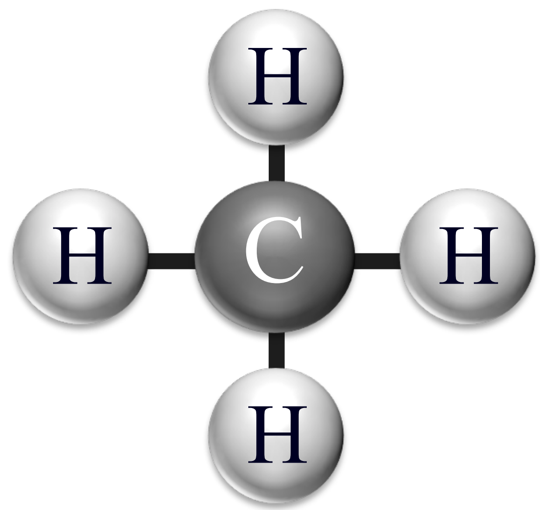


Atom	# of Atoms	Atomic Molar Mass	Total
oxygen	1	x 16.00 g/mole	16.00 g/mole
hydrogen	2	x 1.01 g/mole	2.02 g/mole
		Molar Mass of H_2O =	**18.02 g/mole**

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

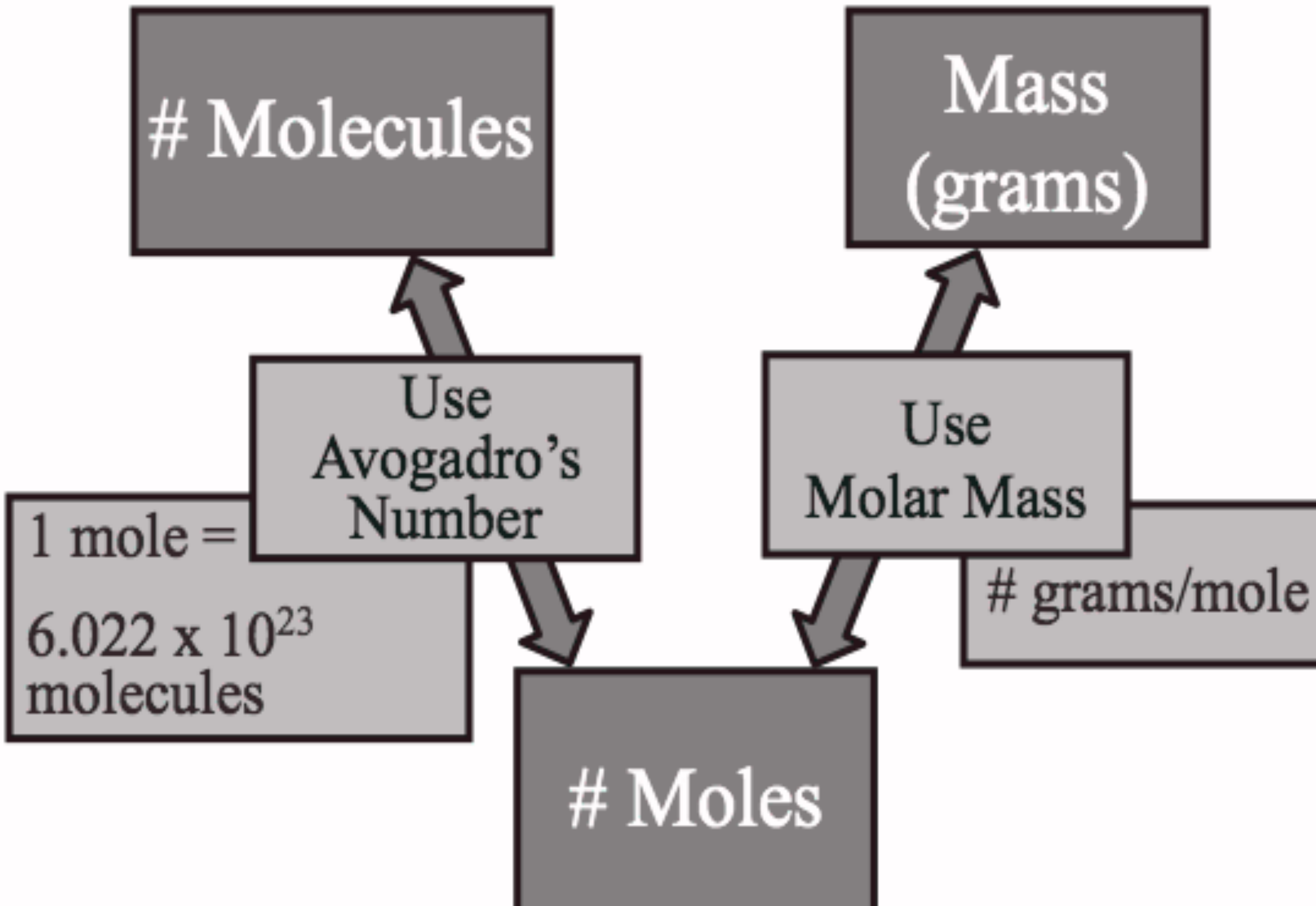
Understanding Check:

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.



Mass-Mole Conversion Example:

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:

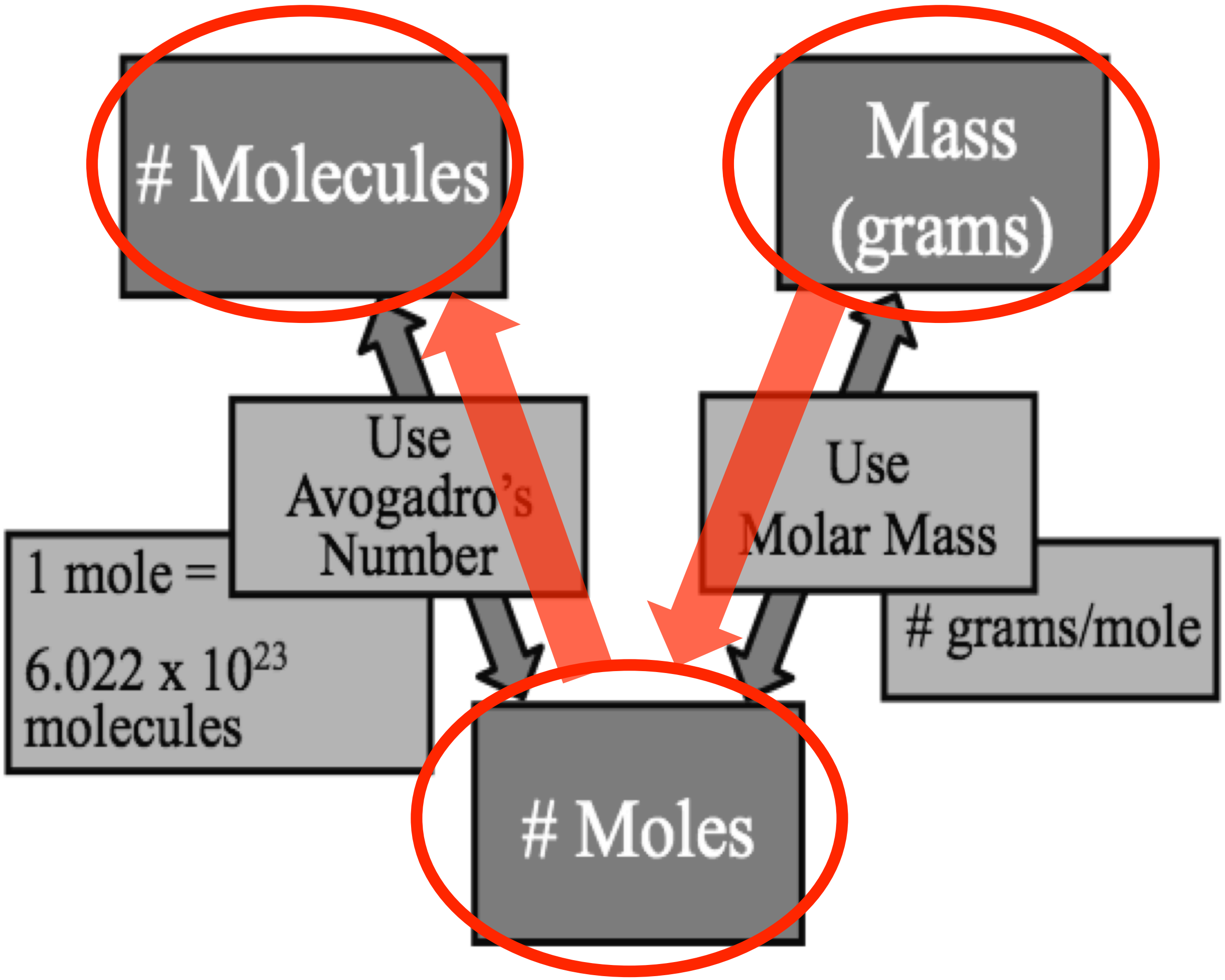
$$\left(\frac{\text{1 mole } CH_4}{\text{16.05 grams}}\right) \longleftrightarrow \left(\frac{\text{16.05 grams}}{\text{1 mole } CH_4}\right)$$

Conversion Factors

$$\frac{3.65 \text{ moles } CH_4}{} \times \frac{\textbf{16.05 grams } CH_4}{1 \text{ mole } CH_4} = 58.6 \text{ grams } CH_4$$

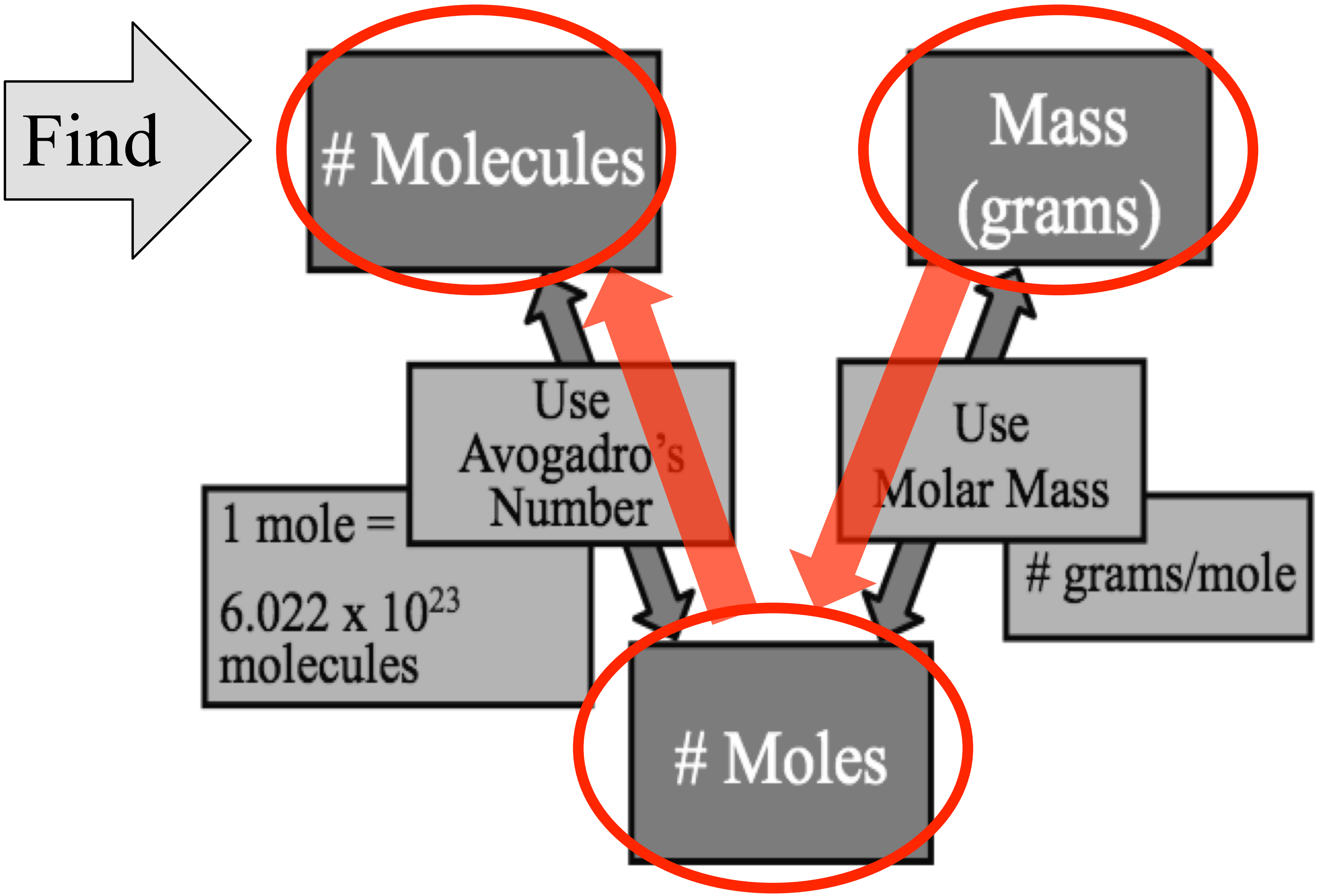
You have just learned how to convert between moles and mass of a compound and vice versa. We do a two-step calculation to convert between mass and number of molecules.

We can *convert between molecules and moles* since **Avogadro's Number applies to molecules**; one mole of a molecular compound contains 6.022 x 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 ***formula*** **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**.

continued on next slide

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

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:

Example: What is the molar mass of iron(II) phosphate, $Fe_{\mathbf{3}}(PO_4)_{\mathbf{2}}$?

$Fe_{\mathbf{3}}$

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

$(PO_4)_{\mathbf{2}}$

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

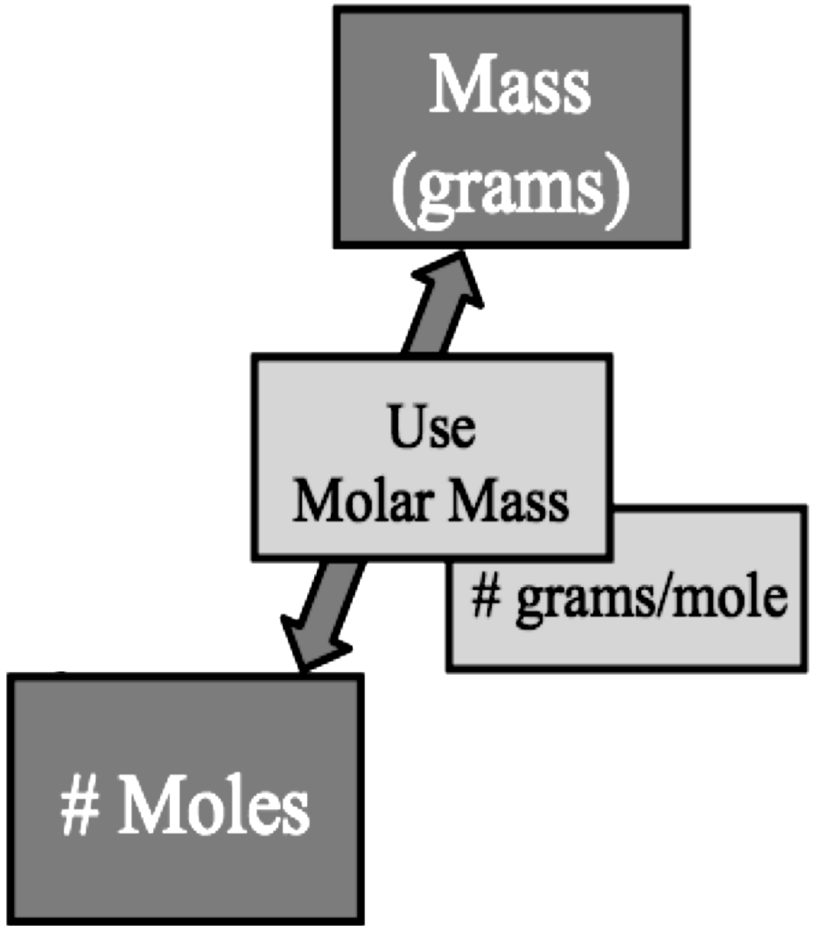
The molar mass of $Fe_{\mathbf{3}}(PO_4)_{\mathbf{2}}$ is **357.49** g/mole

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 $Mg(NO_3)_2$?