

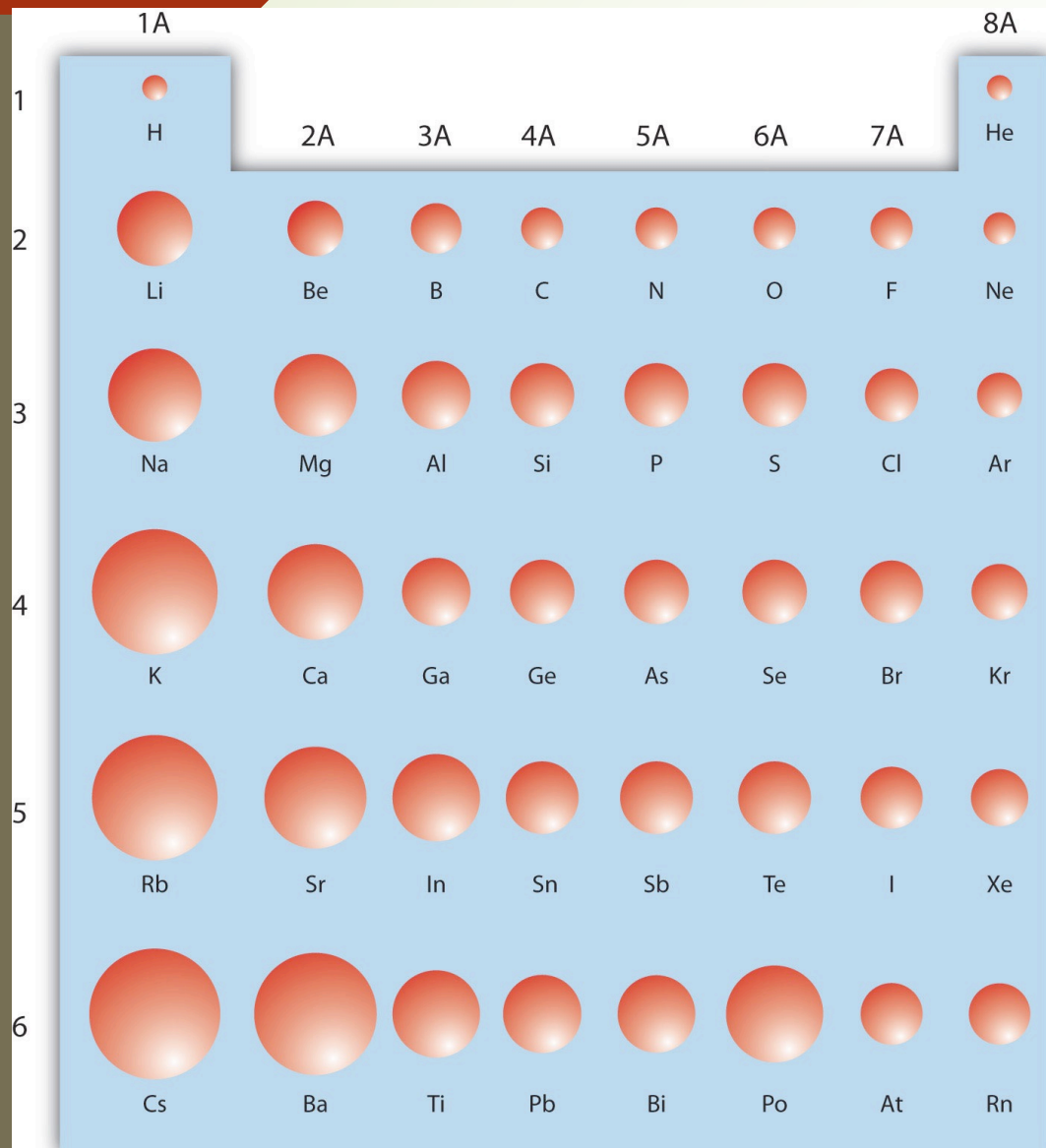
# Chemical Energy

**Objective: Learning about atoms, molecules and ions. Also, explore the formation of ionic and covalent compounds.**

Key concepts:

- ❖ Metals, Nonmetals and Semimetals
- ❖ Electronic Structure
- ❖ Lewis Dot Diagram
- ❖ Cations and Anions
- ❖ Ionic Bonds
- ❖ Covalent Bonds.

# Review: Periodic trends



## Atomic Radius:

- Elements in the **same row (periods)** show a smooth trend toward smaller diameter (size) as you move from left to right.
- Elements in the **same column (group)** get bigger as you go down, due to additional electron shell.

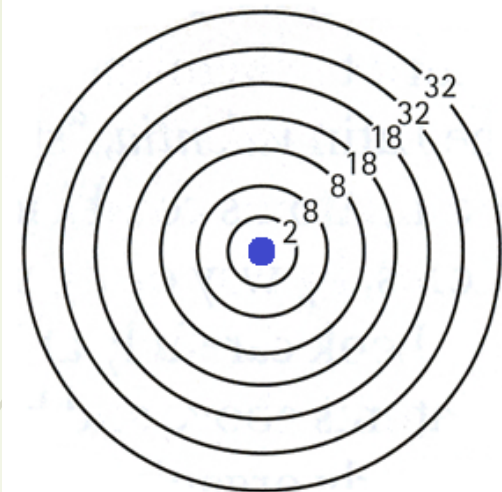
## Chemical properties:

- Elements in the **same column (group or family)** have similar properties, such as how they react (or not) and what kinds of compounds they form.

The original arrangement of the Periodic Table came from organizing elements by their properties, which repeated in a regular pattern with a particular period (frequency). We now understand this to be a result of **the organization of electron energy levels** given by quantum mechanics.

**The highest level is called the valence energy shell.**

# Review: Periodic trends



	1A	2A																								3A	4A	5A	6A	7A	8A		
1	H																														He	2	
2	Li	Be																								B	C	N	O	F	Ne	8	
3	Na	Mg																														8	
4	K	Ca	Sc																													18	
5	Rb	Sr	Y																													18	
6	Cs	Ba	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	32
7	Fr	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	113	114	115	116	117	118	32

- Electrons fill distinct energy “shells” and each row corresponds to the filling of a shell.
- A new period starts when one of these shells is filled.

**Counting valence electrons:** There are never more than 8 valence electrons. *Octet Rule!*

**We do not count the 10 extra electrons in rows 4 and 5 or the 24 extra electrons in rows 6 and 7.** They are not active in bonding because they have a slightly lower energy level than the 8 electrons associated with groups 1, 2, 13, 14, 15, 16, 17, and 18.

Groups **3A**, **4A**, **5A**, **6A**, **7A**, and **8A** identify the number of valence electrons for all elements in those columns. Groups with a “B” label have either 2 or 3 valence electrons.

# Atoms and Molecules

**ATOM:** the basic particle of matter. Example: Ne = one atom of neon.

**ELEMENT:** simplest pure substance, **only one type of atom present.** Examples: Fe = one atom of iron, O<sub>2</sub> = two atoms of oxygen.

**COMPOUND:** pure substance formed from the chemical combination of **two or more elements in a definite ratio.** Examples: H<sub>2</sub>O = water, H<sub>2</sub>SO<sub>4</sub> = sulfuric acid, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub> = sugar.

**MOLECULE:** neutral particle composed of **two of more atoms** chemically combined that **act as an independent unit.** Examples: H<sub>2</sub>O = water (a compound), O<sub>2</sub> = oxygen (an element).

Formulas	Element or Compound	Atom or Molecule
He		
Br <sub>2</sub>		
HCl		
CO		
Co		
Xe <sub>2</sub>		

# Metals, Nonmetals and Semimetals

Main group elements

Group 1

Period—1

Atomic number

Symbol

Metals

Semimetals

Nonmetals

Alkali metals

Alkaline earths

Halogens

Noble gases

Transition elements

Main group elements

1	Transition elements										Main group elements						
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
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	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup			
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		

# Metals, Nonmetals and Semimetals

## Metals

- Located to left of zigzag line
- At room temperature, all except Hg are solid
- Lustrous, malleable & ductile
- Good conductors of heat & electric
- Tend to donate electrons (oxidation) to form **cations** (acquire positive charge)

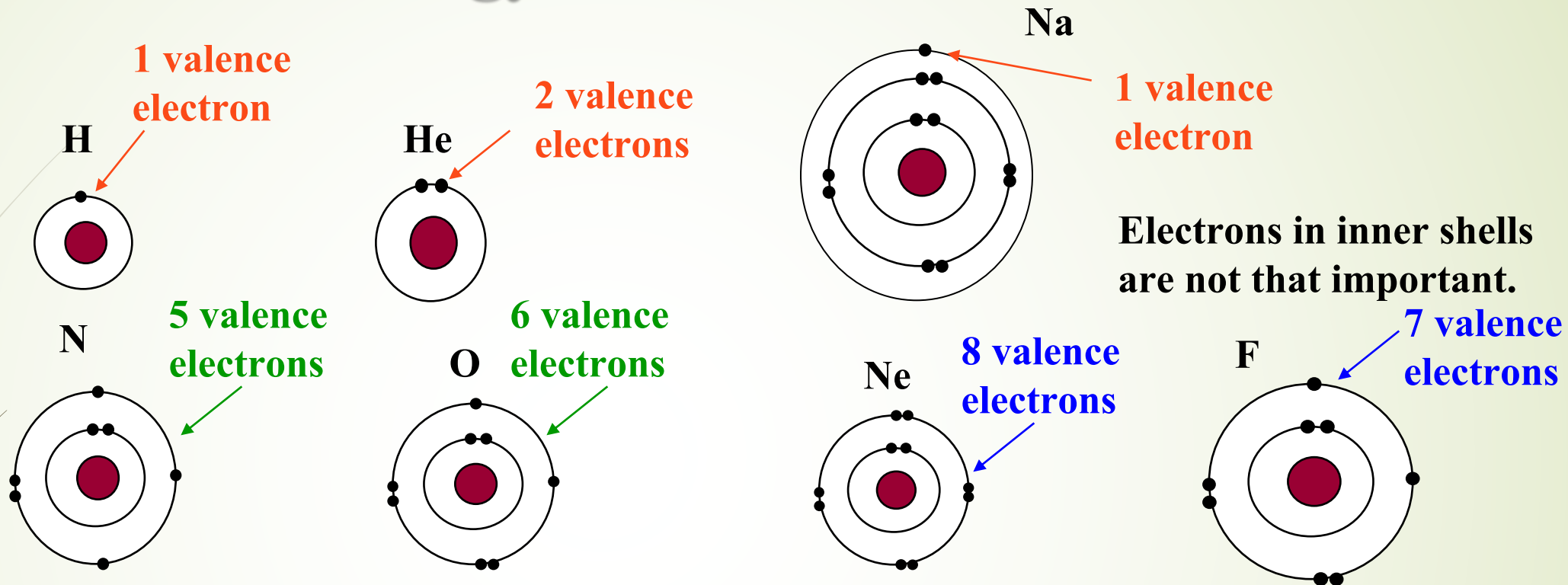
## Nonmetals

- Located to right of zigzag line
- At room temperature, 11 gases, 1 liquid, 5 solids
- Poor conductors of heat & electricity
- Tend to gain electrons (reduction) to form **anions** (acquire negative charge)

## Semimetals (Metalloids)

- Located along zigzag line
- All are solids: B, Si, As, Te, At, Ge, Sb
- Most of their physical properties resemble nonmetals.
- Several of the metalloids are semiconductors, which conduct electricity under special circumstances (Si, Ge).

# Electron Energy Levels & Valence Electrons



For each element, provide number of occupied electron energy levels and number of valence electrons

Elements	Number of energy levels (shells)	Valence electrons in outermost shell
Na		
Ne		
F		

# Metals and Nonmetals

Metals lose electrons (cations)

Non-metals gain electrons (anions)

**Common Ionic States of the Elements**

+1 ←										↓																			
1 H <sup>+</sup> HYDROGEN											5 B BORON	6 C CARBON	7 N <sup>3-</sup> NITROGEN	8 O <sup>2-</sup> OXYGEN	9 F <sup>-</sup> FLUORINE	10 Ne NEON													
+2		Atomic number										+3		-3		-2		-1											
3 Li <sup>+</sup> LITHIUM	4 Be <sup>2+</sup> BERYLLIUM											13 Al <sup>3+</sup> ALUMINIUM	14 Si SILICON	15 P <sup>3-</sup> PHOSPHORUS	16 S <sup>2-</sup> SULFUR	17 Cl <sup>-</sup> CHLORINE	18 Ar ARGON												
11 Na <sup>+</sup> SODIUM	12 Mg <sup>2+</sup> MAGNESIUM											19 K <sup>+</sup> POTASSIUM	20 Ca <sup>2+</sup> CALCIUM	21 Sc <sup>3+</sup> SCANDIUM	22 Ti <sup>3+</sup> TITANIUM	23 V <sup>3+</sup> VANADIUM	24 Cr <sup>2+</sup> CHROMIUM	25 Mn <sup>2+</sup> MANGANESE	26 Fe <sup>2+</sup> IRON	27 Co <sup>2+</sup> COBALT	28 Ni <sup>2+</sup> NICKEL	29 Cu <sup>+</sup> COPPER	30 Zn <sup>2+</sup> ZINC	31 Ga <sup>3+</sup> GALLIUM	32 Ge <sup>4+</sup> GERMANIUM	33 As <sup>3-</sup> ARSENIC	34 Se <sup>2-</sup> SELENIUM	35 Br <sup>-</sup> BROMINE	36 Kr KRYPTON
37 Rb <sup>+</sup> RUBIDIUM	38 Sr <sup>2+</sup> STRONTIUM	39 Y <sup>3+</sup> YTRIUM	40 Zr <sup>4+</sup> ZIRCONIUM	41 Nb <sup>3+</sup> NI OBIUM	42 Mo <sup>6+</sup> MOLYBDENUM	43 Tc <sup>7+</sup> TECHNETIUM	44 Ru <sup>3+</sup> RUTHENIUM	45 Rh <sup>3+</sup> RHODIUM	46 Pd <sup>2+</sup> PALLADIUM	47 Ag <sup>+</sup> SILVER	48 Cd <sup>2+</sup> CADMIUM	49 In <sup>3+</sup> INDIUM	50 Sn <sup>2+</sup> TIN	51 Sb <sup>3+</sup> ANTIMONY	52 Te <sup>2-</sup> TELLURIUM	53 I <sup>-</sup> IODINE	54 Xe XENON												
55 Cs <sup>+</sup> CESIUM	56 Ba <sup>2+</sup> BARIUM	71 Lu <sup>3+</sup> LUTETIUM	72 Hf <sup>4+</sup> HAFNIUM	73 Ta <sup>5+</sup> TANTALUM	74 W <sup>6+</sup> TUNGSTEN	75 Re <sup>7+</sup> RHENIUM	76 Os <sup>4+</sup> OSMIUM	77 Ir <sup>4+</sup> IRIDIUM	78 Pt <sup>2+</sup> PLATINUM	79 Au <sup>+</sup> GOLD	80 Hg <sup>2+</sup> MERCURY	81 Tl <sup>+</sup> THALLIUM	82 Pb <sup>2+</sup> LEAD	83 Bi <sup>3+</sup> BISMUTH	84 Po <sup>2+</sup> POLONIUM	85 At <sup>-</sup> ASTATINE	86 Rn RADON												
87 Fr <sup>+</sup> FRANCIUM	88 Ra <sup>2+</sup> RADIUM	103 Lr <sup>3+</sup> LAWRENCIUM																											
			57 La <sup>3+</sup> LANTHANUM	58 Ce <sup>3+</sup> CERIUM	59 Pr <sup>3+</sup> PRASEODYMIUM	60 Nd <sup>3+</sup> NEODYMIUM	61 Pm <sup>3+</sup> PROMETHIUM	62 Sm <sup>2+</sup> SAMARIUM	63 Eu <sup>2+</sup> EUROPIUM	64 Gd <sup>3+</sup> GADOLINIUM	65 Tb <sup>3+</sup> TERBIUM	66 Dy <sup>3+</sup> DYSPROSIUM	67 Ho <sup>3+</sup> HOLMIUM	68 Er <sup>3+</sup> ERBIUM	69 Tm <sup>3+</sup> THULIUM	70 Yb <sup>3+</sup> YTTERIUM													
			89 Ac <sup>3+</sup> ACTINIUM	90 Th <sup>4+</sup> THORIUM	91 Pa <sup>4+</sup> PROTACTINIUM	92 U <sup>4+</sup> URANIUM	93 Np <sup>5+</sup> NEPTUNIUM	94 Pu <sup>4+</sup> PLUTONIUM	95 Am <sup>3+</sup> AMERICIUM	96 Cm <sup>3+</sup> CURIUM	97 Bk <sup>3+</sup> BERKELIUM	98 Cf <sup>3+</sup> CALIFORNIUM	99 Es <sup>3+</sup> EINSTEINIUM	100 Fm <sup>3+</sup> FERMIUM	101 Md <sup>2+</sup> MEHNDELIUM	102 No <sup>2+</sup> NOBELIUM													

To get the stability of an inert gas,

- **metals** tend to \_\_\_\_\_ electrons to empty the valence shell
- **nonmetals** tend to \_\_\_\_\_ electrons to fill the valence shell.



# Formation of Ions

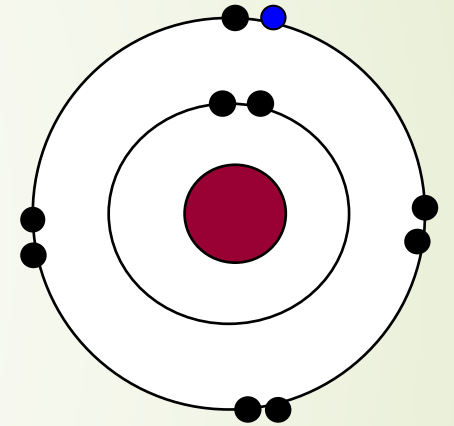
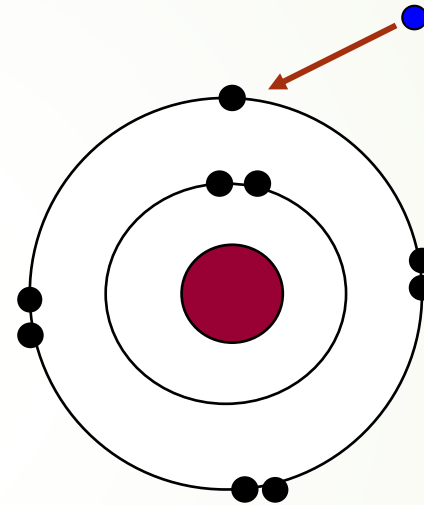
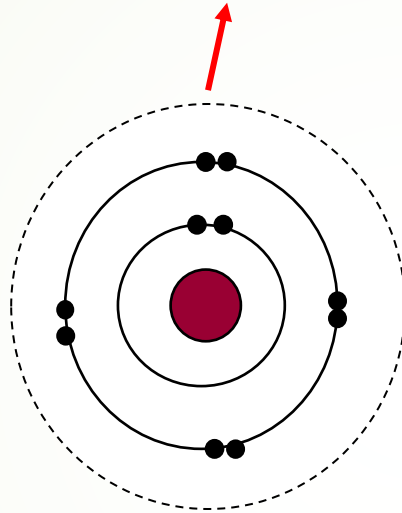
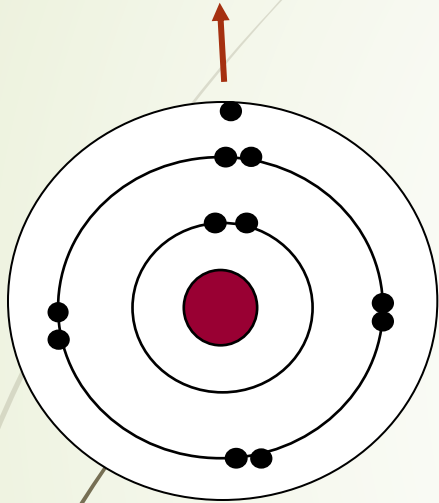
## Positive ion or cation

## Negative ion or anion

lose the electron

vacant valance shell

gain an electron



Na atom

Na ion net charge +1

F atom

F ion net charge -1

Na – 11 protons  
11 electrons

Na<sup>1+</sup> - 11 protons  
10 electrons

F – 9 protons  
9 electrons

F<sup>1-</sup> – 9 protons  
10 electrons

# Lewis “dot” diagram

Chemists use simple diagrams to show an atom’s valence electrons and how they transfer. These diagrams have two advantages over the electron shell diagrams.

- First, they show only valence electrons.
- Second, instead of having a circle around the chemical symbol to represent the electron shell, they have up to eight dots around the symbol; each dot represents a valence electron.

For example, the representation for sodium: Na•

unpaired  
electron

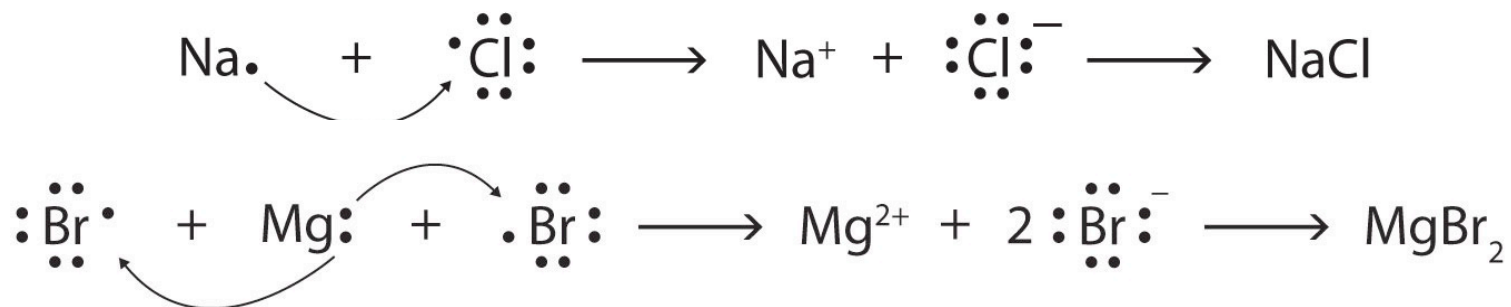


paired  
electrons

And the representation for chlorine:



The transfer of electrons can be illustrated easily with Lewis diagrams:

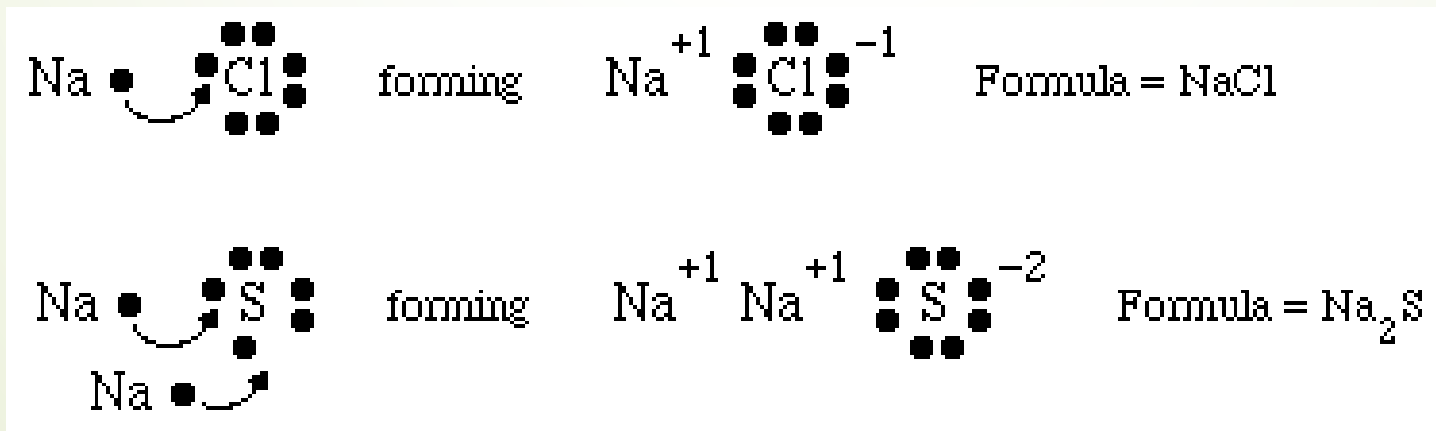


\*\*\*In the final formula, the dots are omitted.

# IONIC BONDING

- The metal atoms give up their valence electrons and become positive ions (cations) with all inner energy levels filled.
  - The non-metal atoms gain extra valence electrons and become negative ions (anions) with all energy levels filled.
  - **The ions stay distinct** and are electrically attracted (bonded) to form Ionic compounds.
- \*\*By convention, the lowest whole-number ratio of the ions is used in ionic formulas. There are exceptions for certain ions, such as  $\text{Hg}_2^{2+}$

Example: A **metal atom** has valence electrons **stolen** by a **non-metal**. The **positive and negative ions** are then **attracted** to each other by the **coulomb force** to form an ionic compound.



# IONIC BONDING

- Additionally, some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called **polyatomic ions**. **Polyatomic ions** have characteristic formulas, names, and charges. Some examples of **polyatomic ions** with their charges are shown below:

$\text{NH}_4^+$	ammonium	$\text{OCN}^-$	cyanate
$\text{H}_3\text{O}^+$	hydronium	$\text{MnO}_4^-$	permanganate
$\text{OH}^-$	hydroxide	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate ( $\text{OAc}^-$ , $\text{CH}_3\text{CO}_2^-$ )
$\text{CN}^-$	cyanide	$\text{CO}_3^{2-}$	carbonate
$\text{O}_2^{2-}$	peroxide	$\text{HCO}_3^-$	hydrogen carbonate, bicarbonate
$\text{N}_3^-$	azide	$\text{SO}_4^{2-}$	sulfate
$\text{NO}_3^-$	nitrate	$\text{SO}_3^{2-}$	sulfite
$\text{NO}_2^-$	nitrite	$\text{S}_2\text{O}_3^{2-}$	thiosulfate
$\text{ClO}_3^-$	chlorate	$\text{C}_2\text{O}_4^{2-}$	oxalate
$\text{ClO}_2^-$	chlorite	$\text{CrO}_4^{2-}$	chromate
$\text{ClO}^-$	hypochlorite	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
$\text{ClO}_4^-$	perchlorate	$\text{PO}_4^{3-}$	phosphate

# IONIC BONDING

Practice: Identify each compound as ionic or not ionic.

Compound	Ionic or not ionic
Na <sub>2</sub> O	
PCl <sub>3</sub>	
NH <sub>4</sub> Cl	
OF <sub>2</sub>	
N <sub>2</sub> O	

Show the “Lewis dot diagram” and electron transfer with an arrow.

Mg

F

Formula: \_\_\_\_\_

# A Guide to Naming Simple Ionic Compounds

Identify:

- Cation name
- Anion name

Can the cation have more than one possible charge?

yes

no

Use Stock system name of cation + name of anion.  
or  
Use stem of cation name + -ic/-ous + name of anion.

Examples:

- $\text{FeCl}_2$       iron(II) chloride or ferrous chloride
- $\text{CuSO}_4$       copper(II) sulfate or cupric sulfate
- $\text{Cr}_2\text{O}_3$       chromium(III) oxide or chromic oxide

Use name of cation + name of anion.

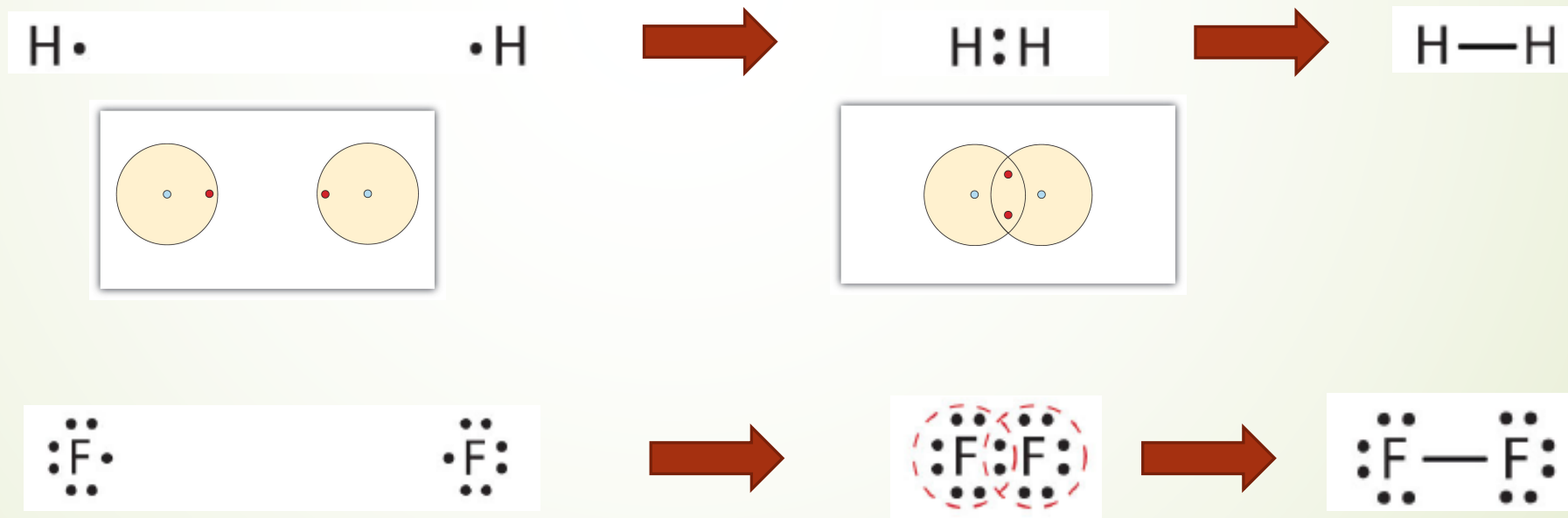
Examples:

- $\text{KBr}$       potassium bromide
- $\text{NaNO}_3$       sodium nitrate
- $(\text{NH}_4)_2\text{S}$       ammonium sulfide

# COVALENT BONDING

- An unpaired electron in one atom is **shared** with an unpaired electron in another atom. Each atom is happy when it thinks it has a full set of **8 valence electrons** even though some electrons actually belong to a different atom and are just being shared with it.
- **Specific formula resulting in molecules.** The formula is determined by the number of **unpaired valence electrons**.

This is where the **Lewis dot diagrams**, which show us how the valence electrons are organized, are most useful. We put atoms with unpaired electrons next to each other and then put a line between the electrons to show the bonded pair.



# COVALENT BONDING

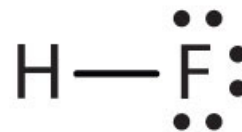
**Covalent Bonds between Different Atoms:** Consider a molecule composed of one hydrogen atom and one fluorine atom.



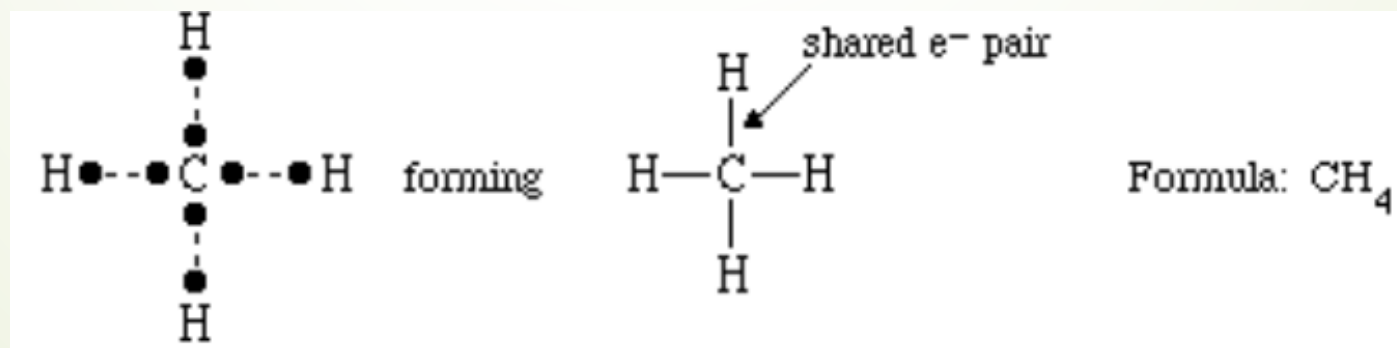
Each atom needs one additional electron to complete its valence shell. By each contributing one electron, they make the following molecule:



or



Larger molecules are constructed in a similar fashion, with some atoms participating in more than one covalent bond. For example, methane ( $\text{CH}_4$ ), with one carbon atom and four hydrogen atoms, can be represented as follows:





# How Many Covalent Bonds Are Formed?

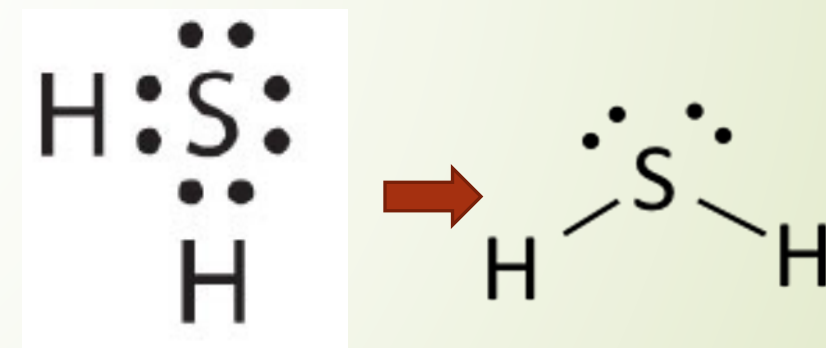
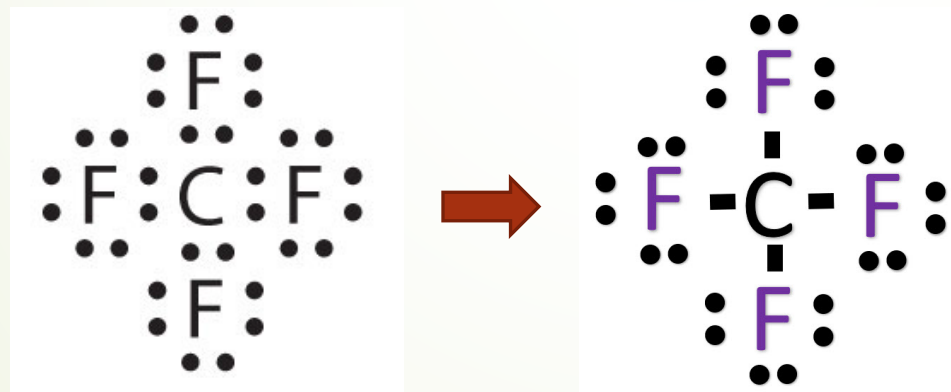
1																							
																		4	3	2	1		
																		4	3,5	2,6	1		
																					2	1	

*In molecules, there is a pattern to the number of covalent bonds that different atoms can form. Each block with a number indicates the number of covalent bonds formed by that atom in neutral compounds.*

# COVALENT BONDING

- Always start by drawing the dot diagram for the atom at the center of the group.  
*(We tell you what that is)*
- The formula tells you how many of the other atoms need to be added to find a partner for every unpaired atom.
- A dash connects paired electron dots to show the bond.

Example:



# Multiple Covalent Bonds

- Two pairs of electrons shared between two atoms make a **double bond** between the atoms, which is represented by a double dash:



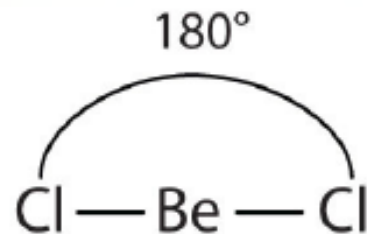
- Some molecules contain **triple bonds**, covalent bonds in which *three* pairs of electrons are shared by two atoms. A simple compound that has a triple bond is acetylene ( $\text{C}_2\text{H}_2$ ), whose Lewis diagram is as follows:



# Molecular Shape: VSEPR Theory

Unlike ionic compounds, covalent molecules are discrete units with specific 3-D shapes. The shape of a molecule is determined by the fact that covalent bonds, which are composed of negatively charged electrons, tend to repel one another. This concept is called the **Valence Shell Electron Pair Repulsion (VSEPR)** theory.

For example, the two covalent bonds in  $\text{BeCl}_2$  stay as far from each other as possible, ending up  $180^\circ$  apart from each other. The result is a *linear molecule*:



The three covalent bonds in  $\text{BF}_3$  repel each other to form  $120^\circ$  angles in a plane, in a shape called *trigonal planar*:

