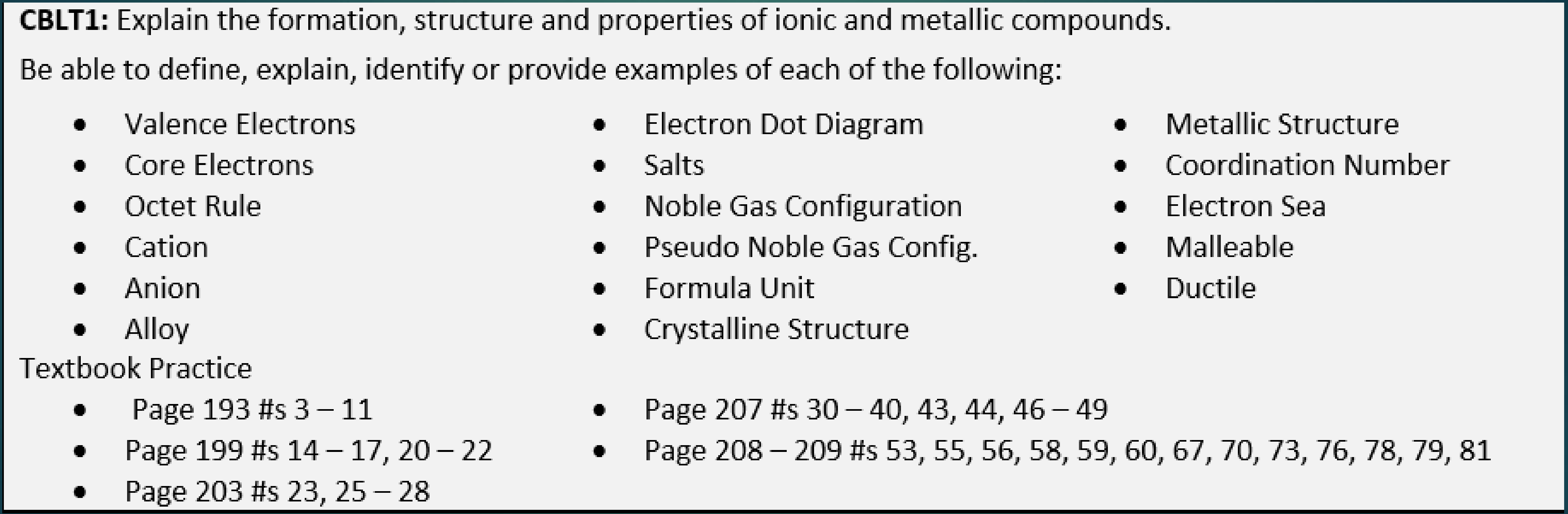
Chemical Bonding

IONIC, METALLIC AND COVALENT BONDS

Learning Target #1



CBLT1: Explain the formation, structure and properties of ionic and metallic compounds.

Be able to define, explain, identify or provide examples of each of the following:

- Valence Electrons
- Core Electrons
- Octet Rule
- Cation
- Anion
- Alloy
- Electron Dot Diagram
- Salts
- Noble Gas Configuration
- Pseudo Noble Gas Config.
- Formula Unit
- Crystalline Structure
- Metallic Structure
- Coordination Number
- Electron Sea
- Malleable
- Ductile

Textbook Practice

- Page 193 #s 3 – 11
- Page 199 #s 14 – 17, 20 – 22
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- Page 207 #s 30 – 40, 43, 44, 46 – 49
- Page 208 – 209 #s 53, 55, 56, 58, 59, 60, 67, 70, 73, 76, 78, 79, 81

Core and Valence Electrons

- The electrons are responsible for chemical properties of atoms are those in the **outer/highest** energy level (principle quantum number ***n***).
- **Valence e^-**: The ***s*** and ***p*** orbital electrons in the highest energy level.
- **Core e^-**: All the electrons in the energy levels below the highest.

Modeling Valence Electrons: Electron Dot Structures

- Electron dot structures are used to represent the valence electrons in a atom.
  - They are represented as dots on the top, bottom an sides of an element symbol.
- When creating an electron dot structure, place a dot at one of the sides of the element and continue placing dots around the symbol but not pairing them up until you have to (Hund's rule)
  - This is very important because the bonds that form incorporate paired and unpaired electrons in different ways.

Electron Dot Structure/Diagram

- Place first dot.
- Next dot is placed at an adjacent side.
- Continue the pattern.
- Never more than 8 dots.
- Use Periodic Table for the number of valence e^-.

Octet Rule

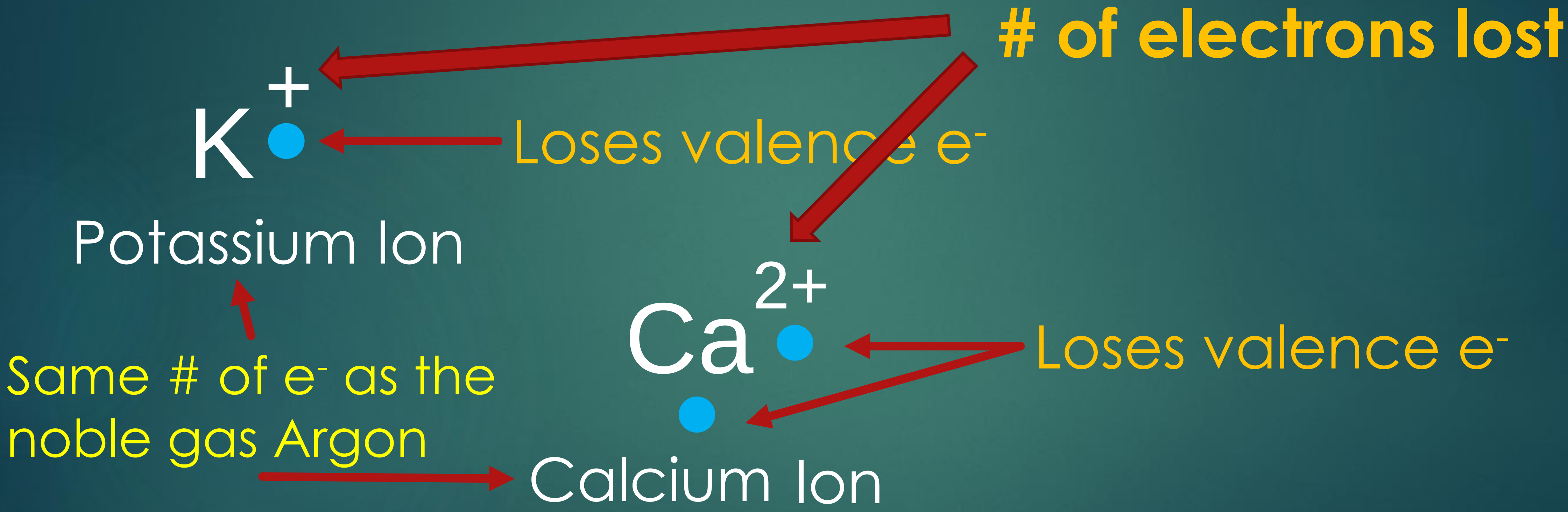
- When forming compounds, atoms tend to achieve a noble gas configuration; 8 e- in the outer level is the most stable.
  - There are exceptions, but we will only work with compounds that obey this rule.
- Metals lose valence electrons.
- Nonmetals gain (or share) one or more electrons to fill their highest energy level.

Pseudo Noble Gas Configuration

- Many transition metals would need to lose or gain many (like 6 or more) electrons to achieve a noble gas configuration – this is just not possible.
- Some transition metals will aim for a pseudo noble gas configuration, which has 18e$^-$ in the highest energy level.
  - The ***s***, ***p*** and ***d*** orbitals are filled.
  - $3s^2 3p^6 3d^{10}$ for a total of 18e$^-$.

Cation

- A positive ion, formed by losing one or more electrons.



# of electrons lost

K^+

Loses valence e^-

Potassium Ion

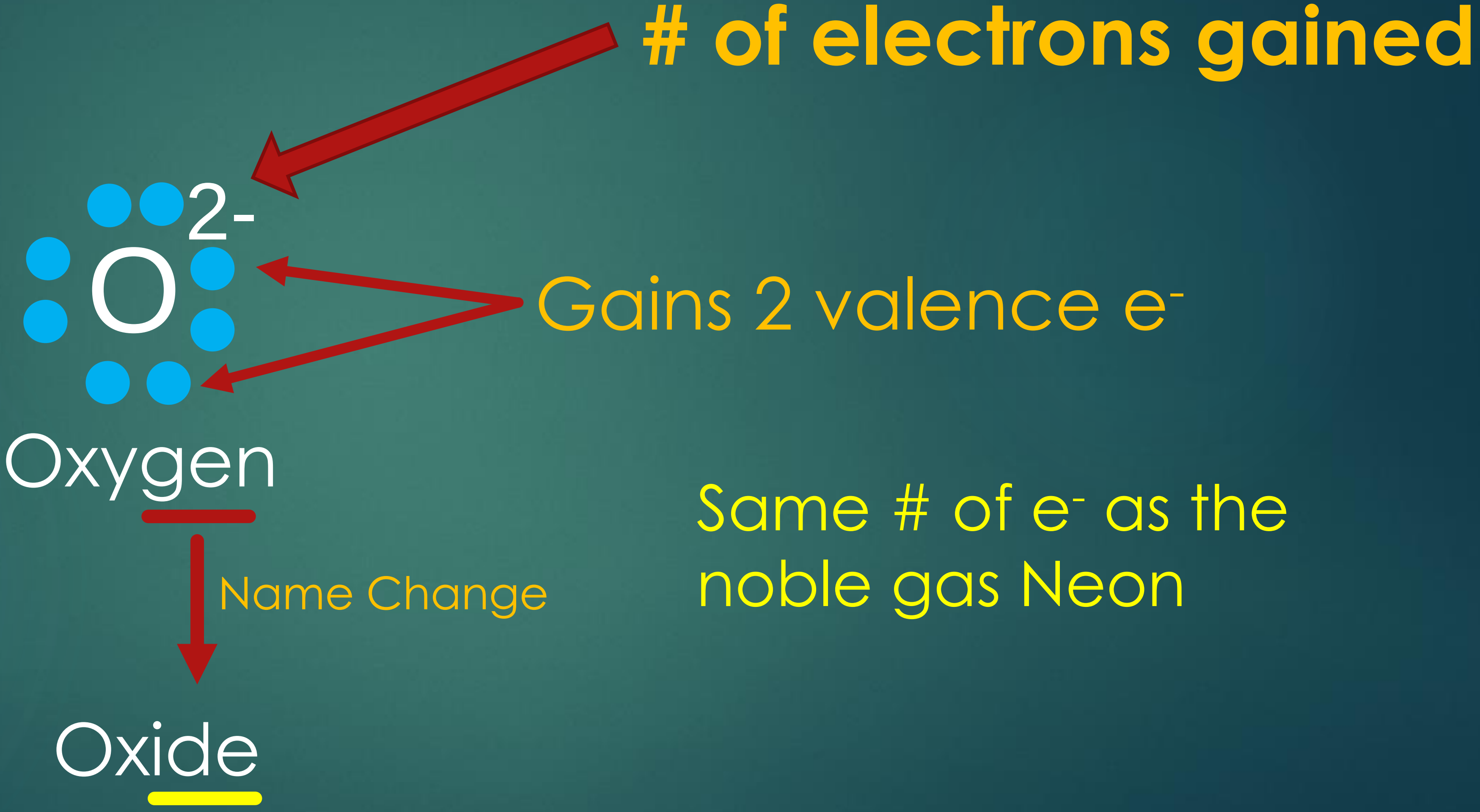
Same # of e^- as the noble gas Argon

Ca^{2+}

Loses valence e^-

Calcium Ion

Anion



# of electrons gained

O^{2-}

Gains 2 valence e^-

Oxygen

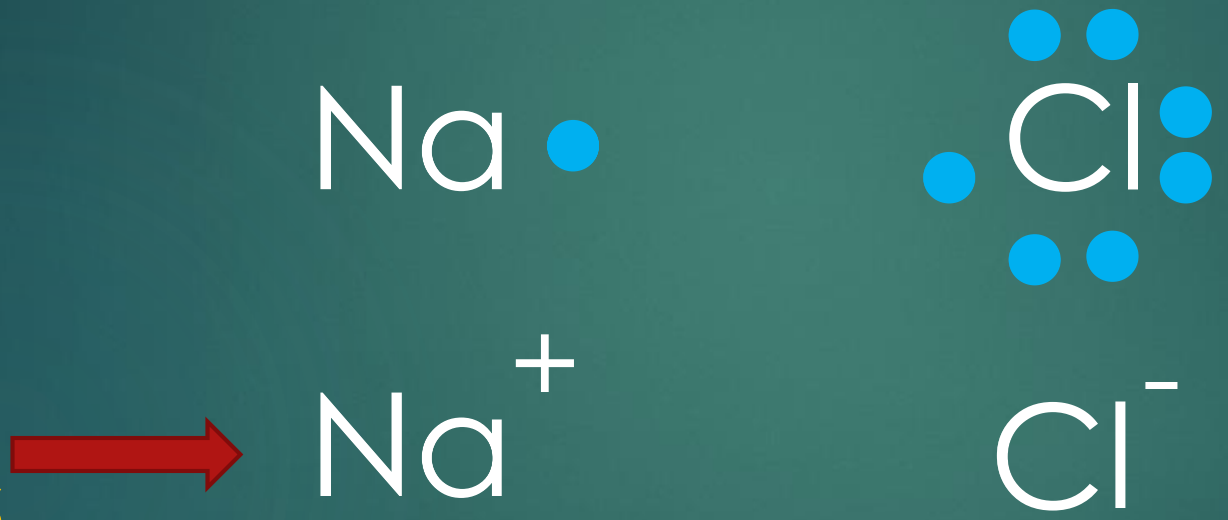
Name Change

Oxide

Same # of e^- as the noble gas Neon

Visualizing Electron Transfer

- In this case, you can see how the one e^- from Na can complete the octet for Cl.



Forms the ions

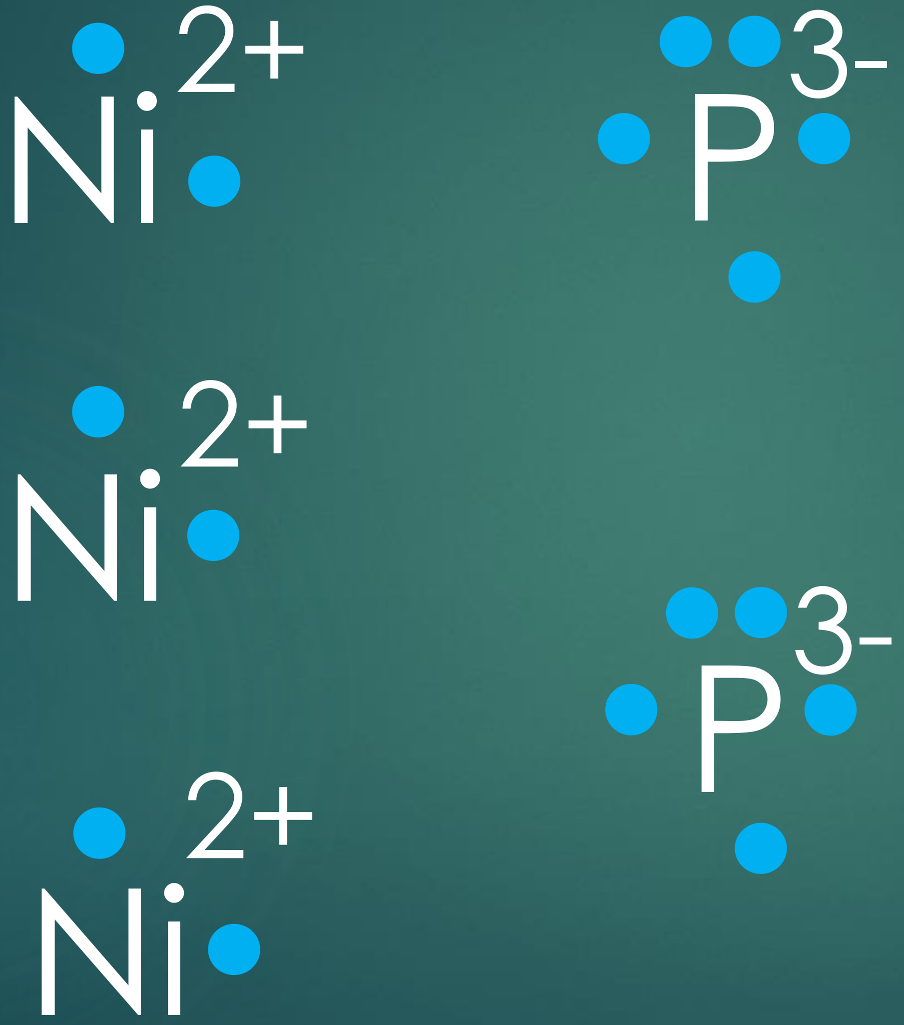
Formula Unit

NaCl

Sodium Chloride

Positive and negative charges hold the compound together!

Ionic Compound between
Nickel and Phosphorus



Ni has 2 valence e^-,
Phosphorus wants 3 more.

P still needs 1 more e^-!
P is full! Ni has one more!

Ni needs to give this e^-!
Same problem as before!

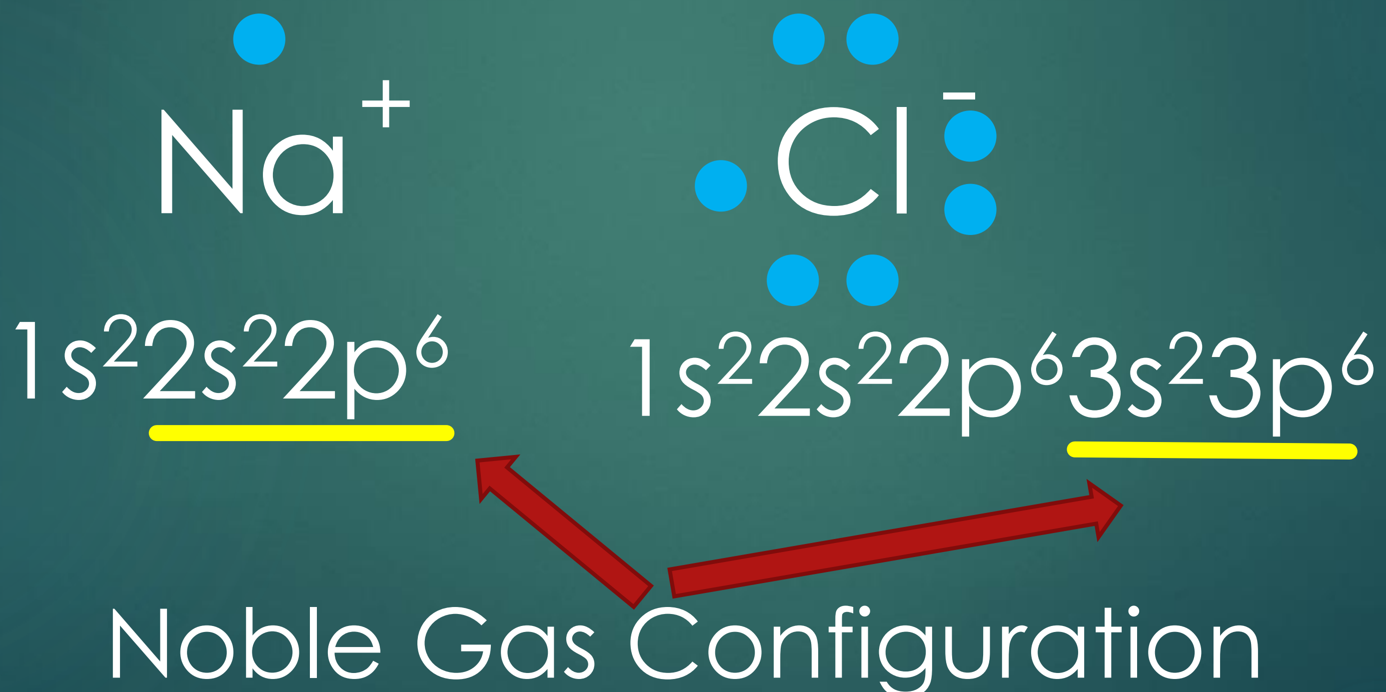
Formula Unit

Ni_3P_2

Nickel Phosphide

Ionic Compound

- Electrons are transferred to achieve a noble gas configuration (8e⁻ in the highest energy level).



Na^+ Cl^-

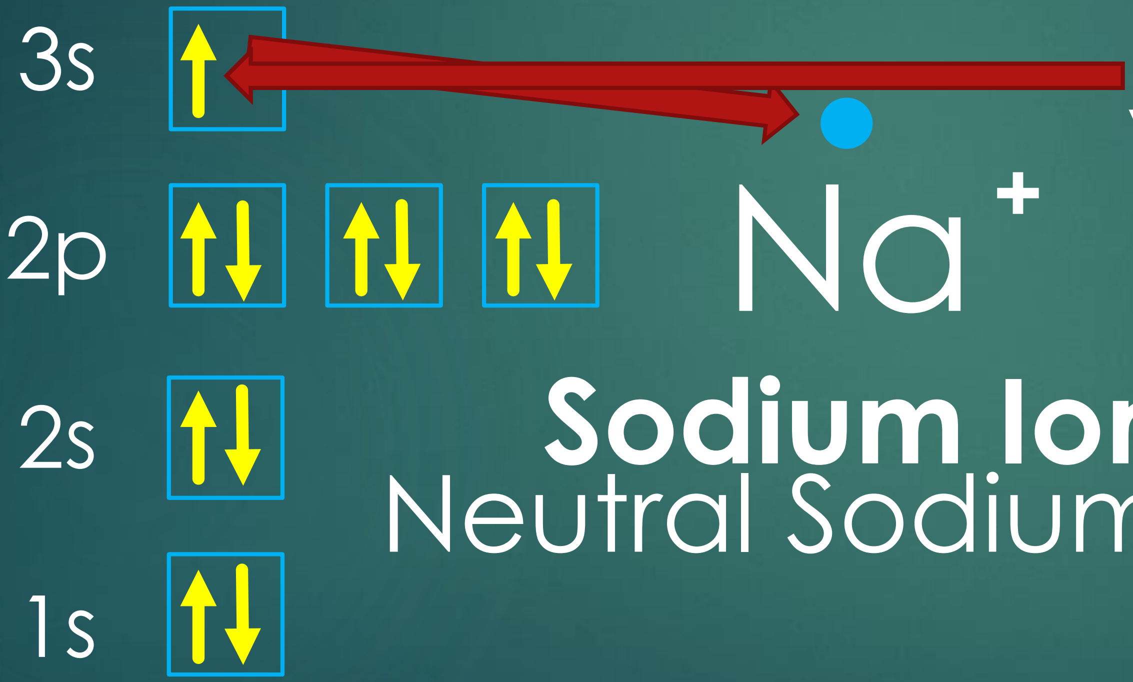
$1s^2\underline{2s^22p^6}$

$1s^22s^22p^6\underline{3s^23p^6}$

Noble Gas Configuration

Cations

- The name for atoms (or compounds) that have a positive charge due to the loss of one or more electrons.



Na^+

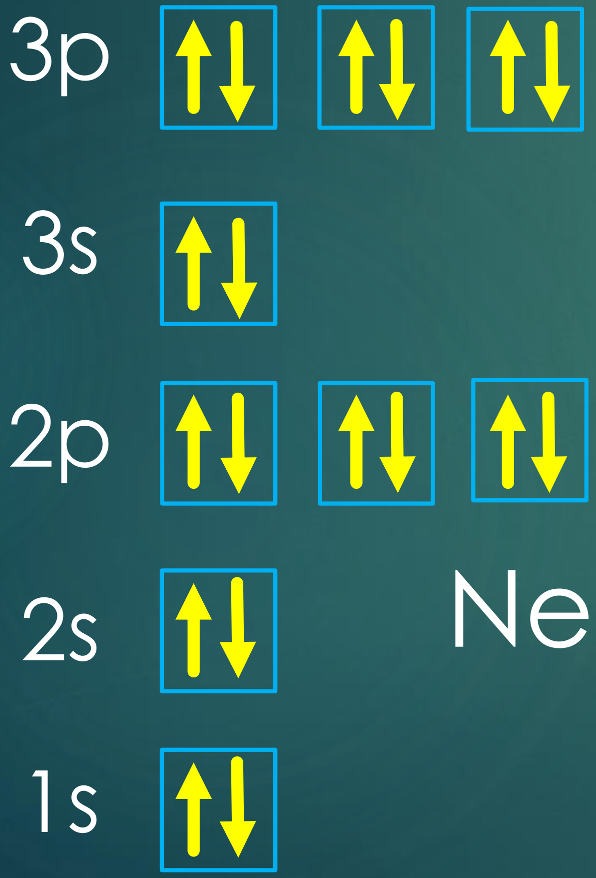
Loss of valence e^-

Sodium Ion

Neutral Sodium Atom

Anions

- The name for atoms (or compounds) that have a negative charge due to the gain of one or more electrons.

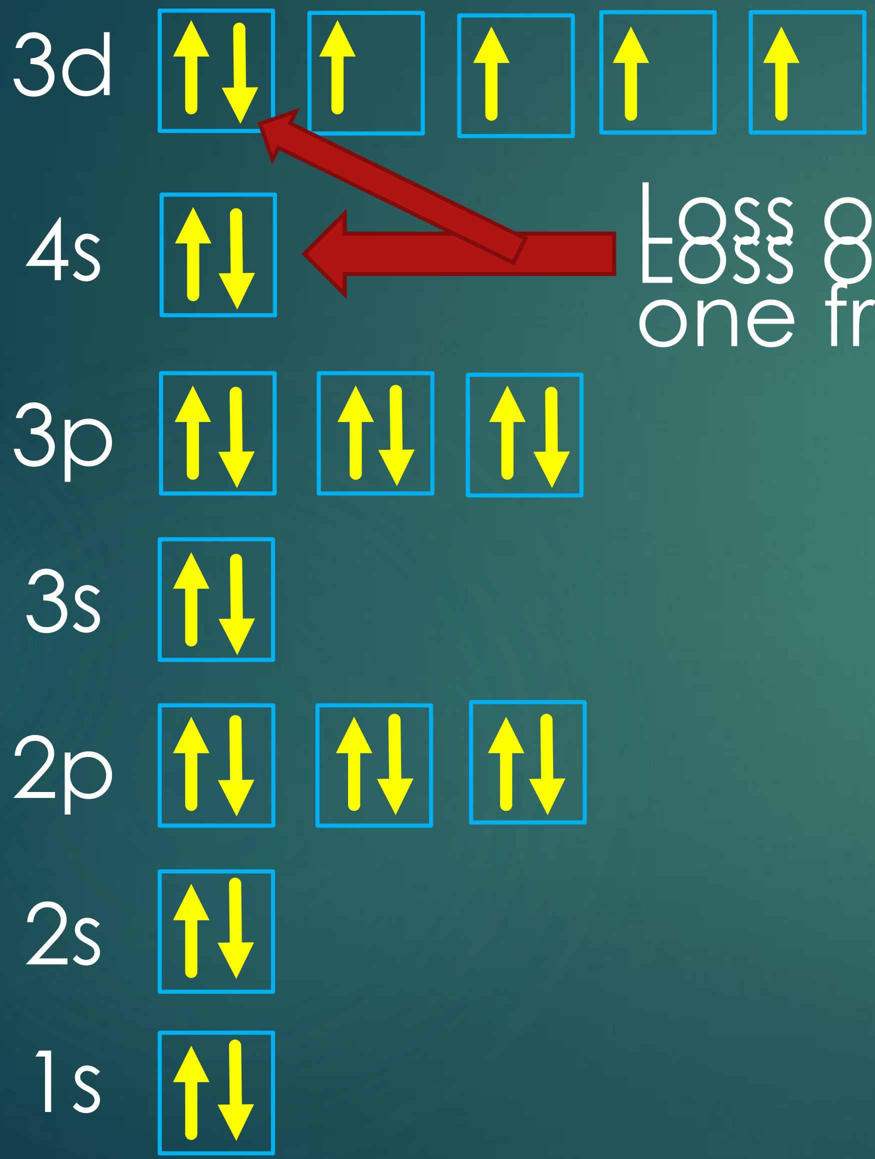


Gain 2 valence e^-

Sulfide

Neutral Sulfur Atom

Cations With More Than One Ion



Half-filled orbital is a lower energy/more stable configuration

Loss of Valence e$^-$ and one from the ***3d*** orbital

Loss of Valence e$^-$

Fe^{3+}

Predict Tin's Possible Ionic Charges

Pseudo noble gas configuration

Sn^{2+}

Sn^{4+}

Filled orbital is a lower energy/more stable configuration

Ionic Compounds

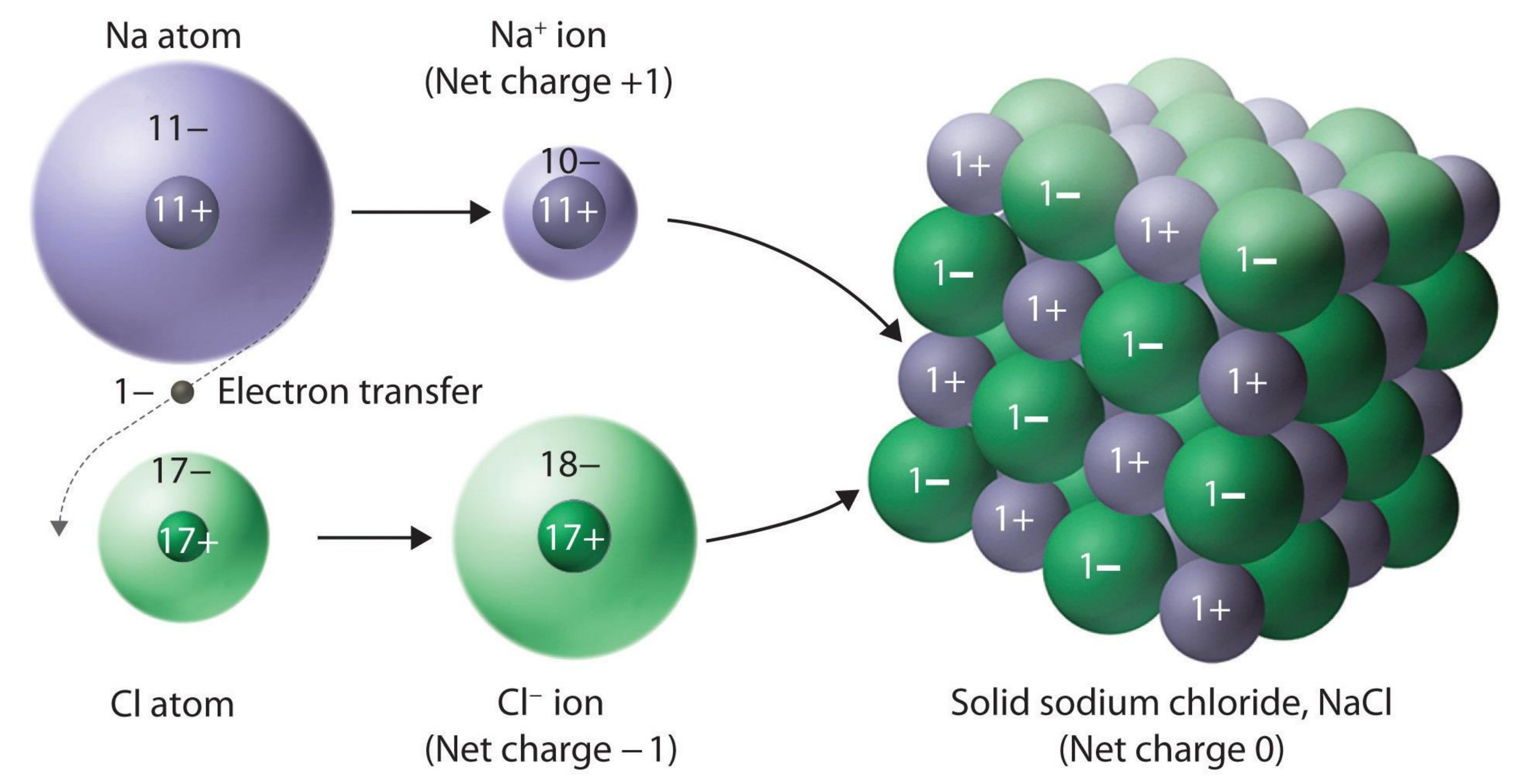
- Are compounds composed of cations and anions.
  - Also called ***salts***.
- Electrically neutral (charges are present, but they balance out.
- Cations transfer electrons to the anions, the electrostatic force between the ions creates the ***ionic bond***.
- Simplest ratio of ions is called the ***formula unit***.

Formula Unit vs Chemical Formula

- A ***chemical formula*** communicates the kinds and number of atoms in the smallest part of a substance.
  - Ionic compounds don't have a "smallest part"
- Ionic compounds have a ***formula unit***, the lowest whole-number ratio of ions to have a neutral charge.
- NaCl is a 1:1 ratio
- Fe_2O_3 is a 2:3 iron-to-oxygen ratio

Structure of Ionic Compounds

- Exist as a crystalline solid – a regular repeating arrangement of ions.

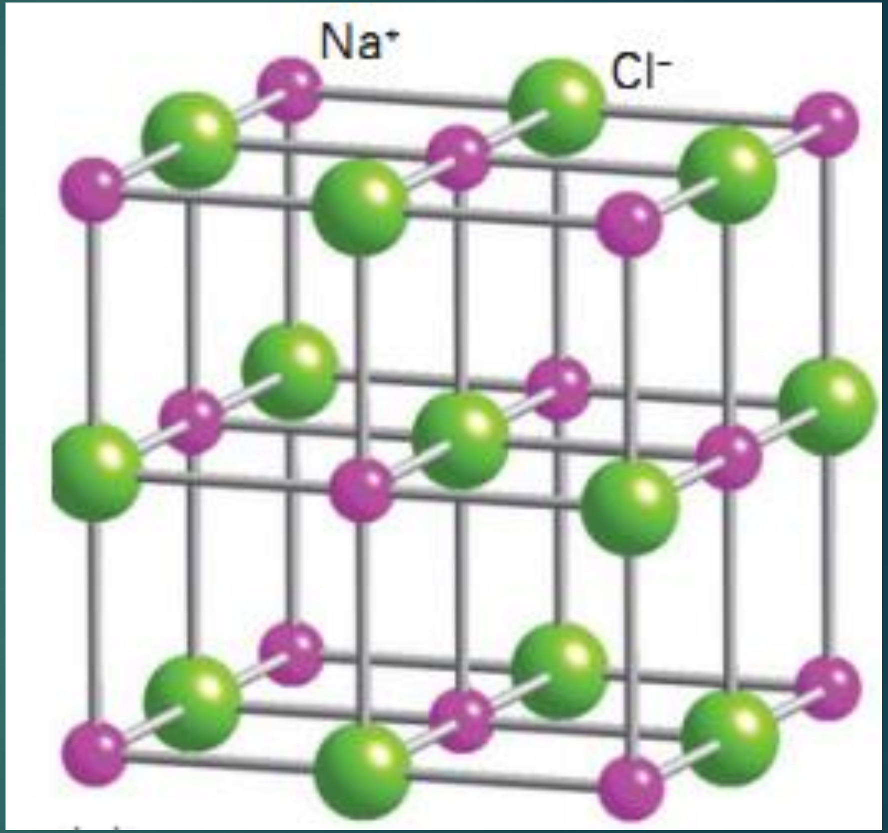


Properties of Ionic Compounds

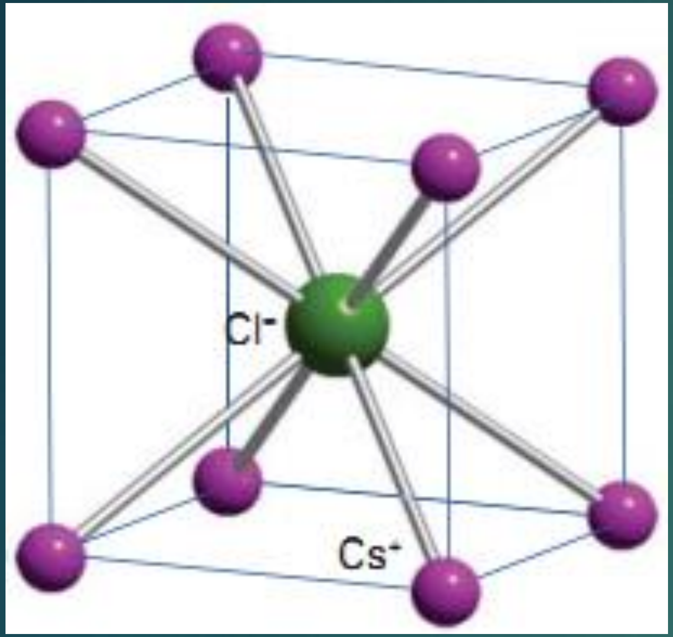
- Crystalline solids at room temperature.
- The alternating of + and – ions results in a very stable structure.
- The electrostatic force between the ions is very strong.
- Brittle when an external force is applied.
- Very high melting points.
- Conduct electric current when melted or dissolve.

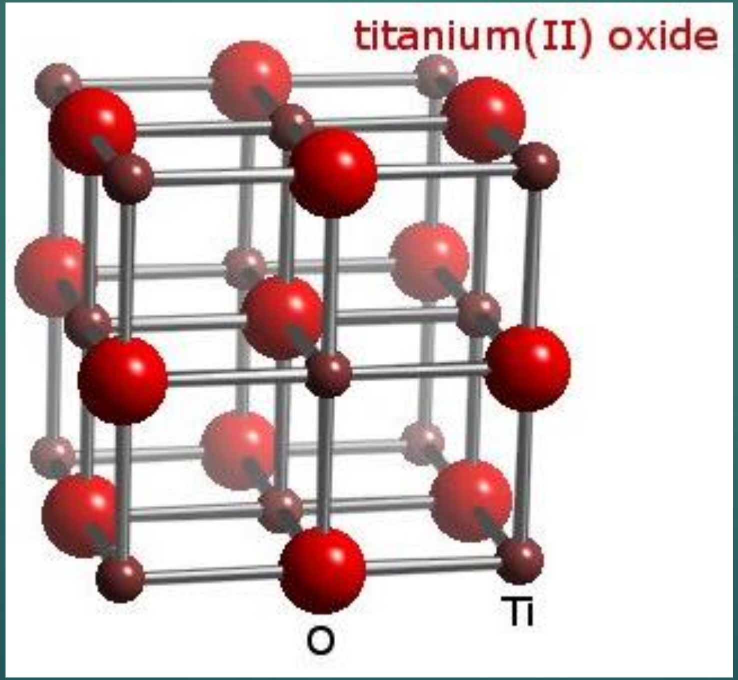
Crystalline Structure

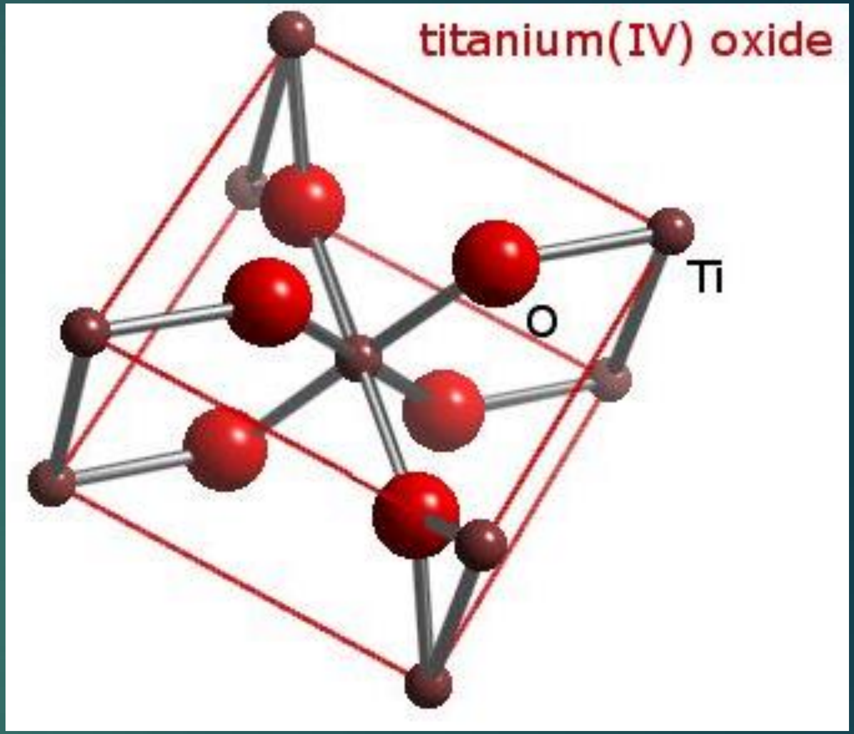
- The arrangement of ions in ionic compounds is an alternating of + and – ions, but the varying sizes of the atoms the number of electrons involved lead to varying internal structures.
- Internal structures are summarized by an ion's ***coordination number***. That is the number of ions of opposite charge that surround it. In sodium chloride, they each have a coordination number of 6.



Other Crystalline Structures

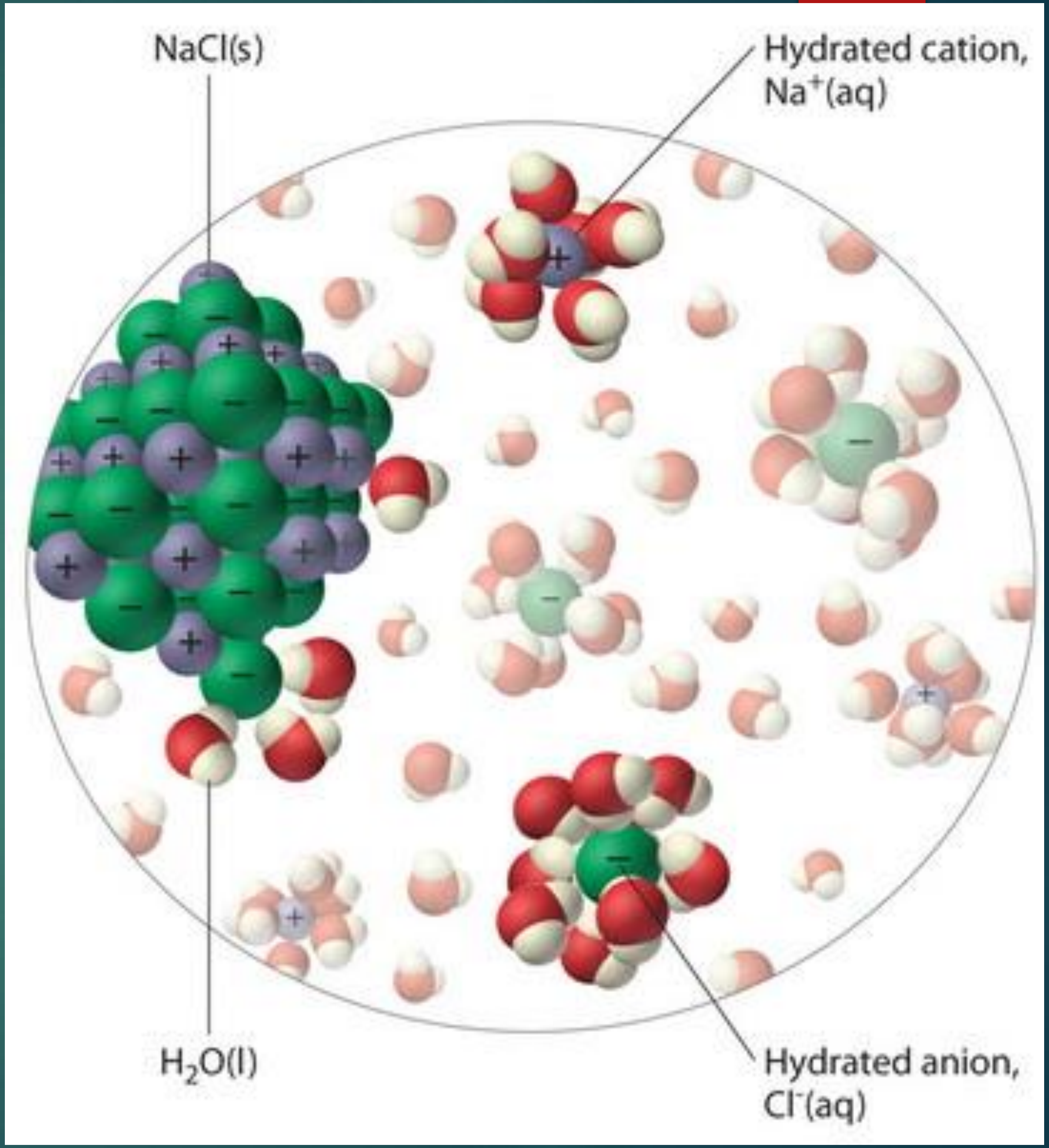






Conductivity When Melted and Dissolved

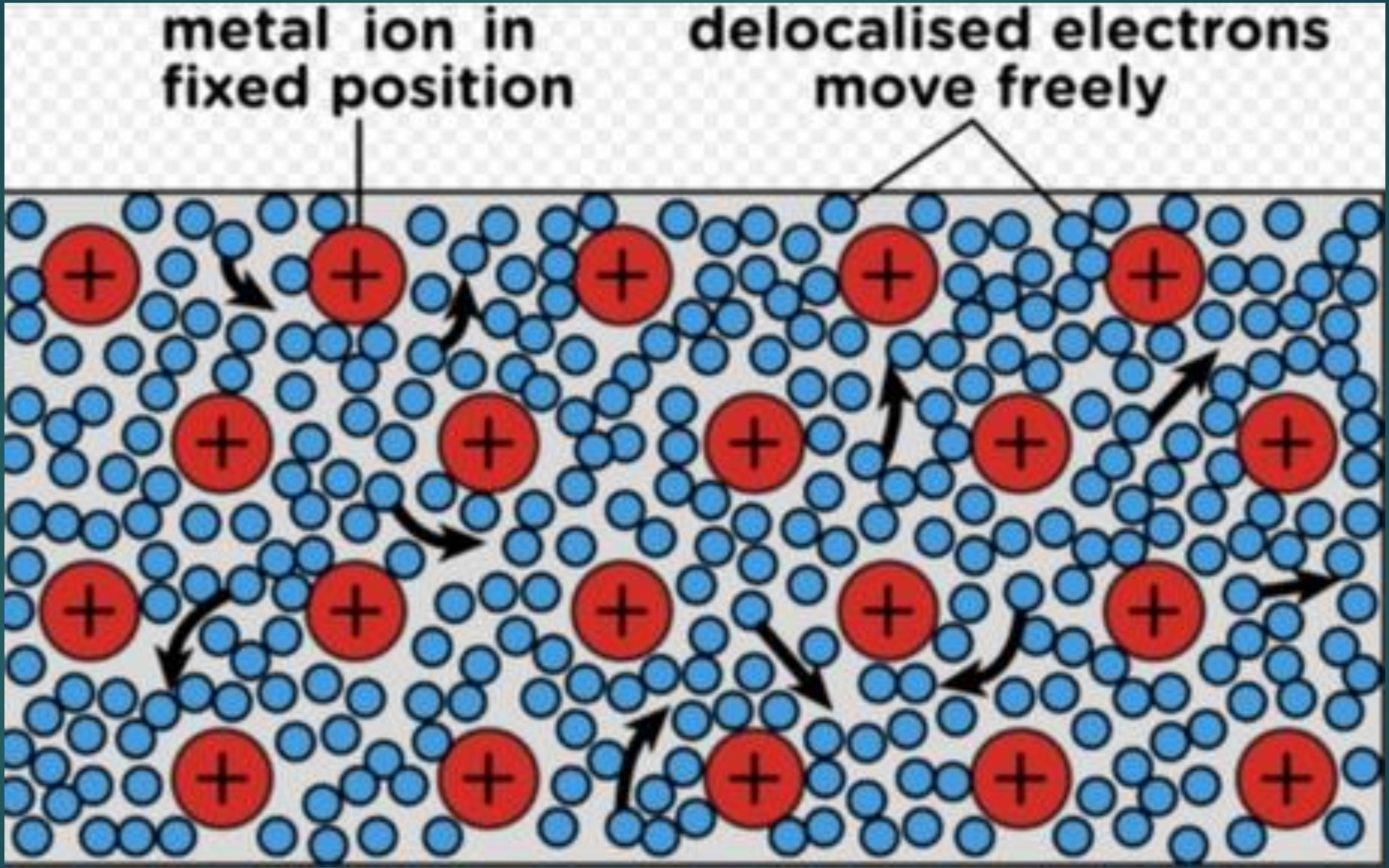
- Ionic compound separate into their ions when placed in water or melted.
- The ions are then capable of conducting an electric current as they orient within an electric field.



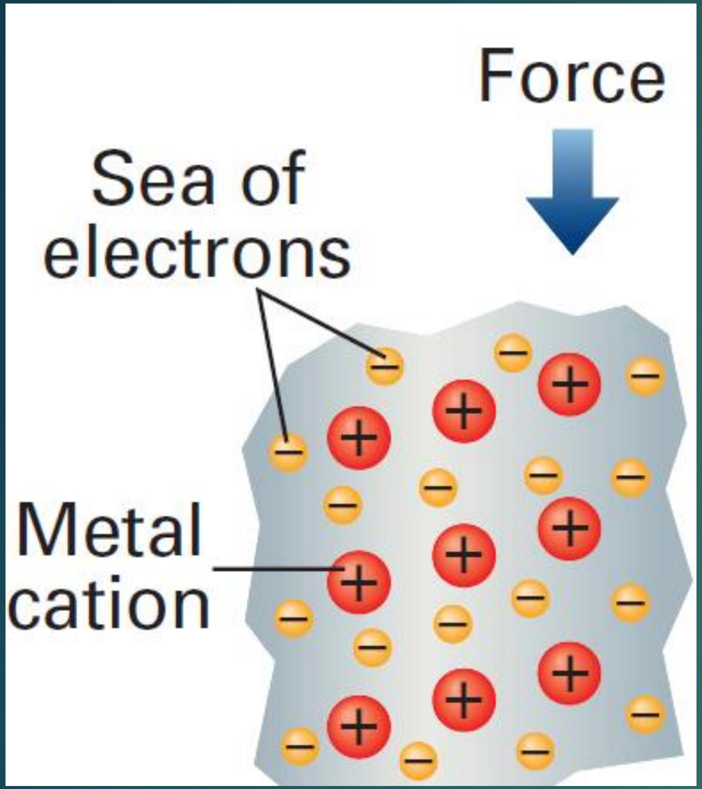
Bonding in Metals

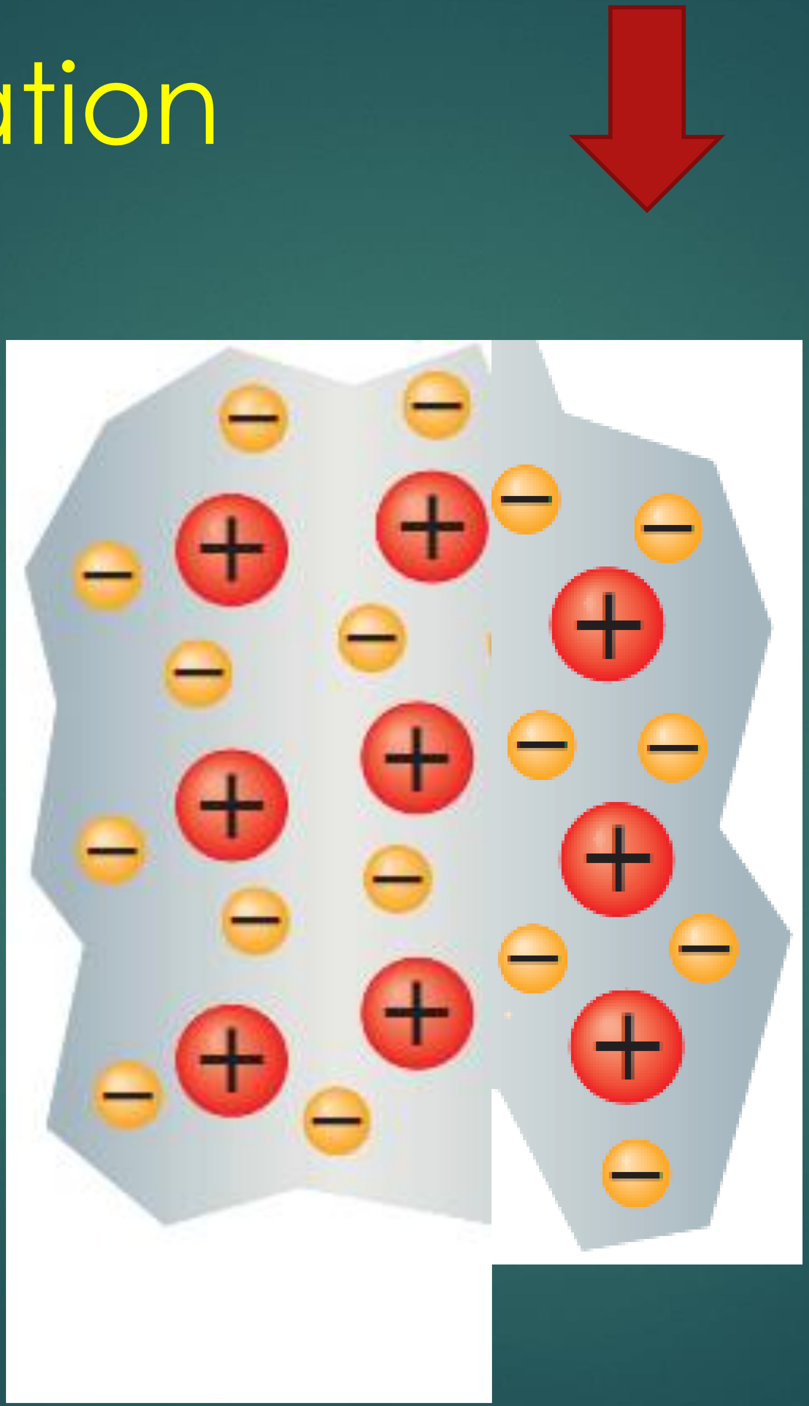
- Valence e- in metals can be modeled as a *sea of electrons*.
  - They are mobile and change positions as needed.
- Explains the physical properties of metals.
  - High melting points (strong electrostatic bonds).
  - Malleable and ductile.

Sea of Electrons

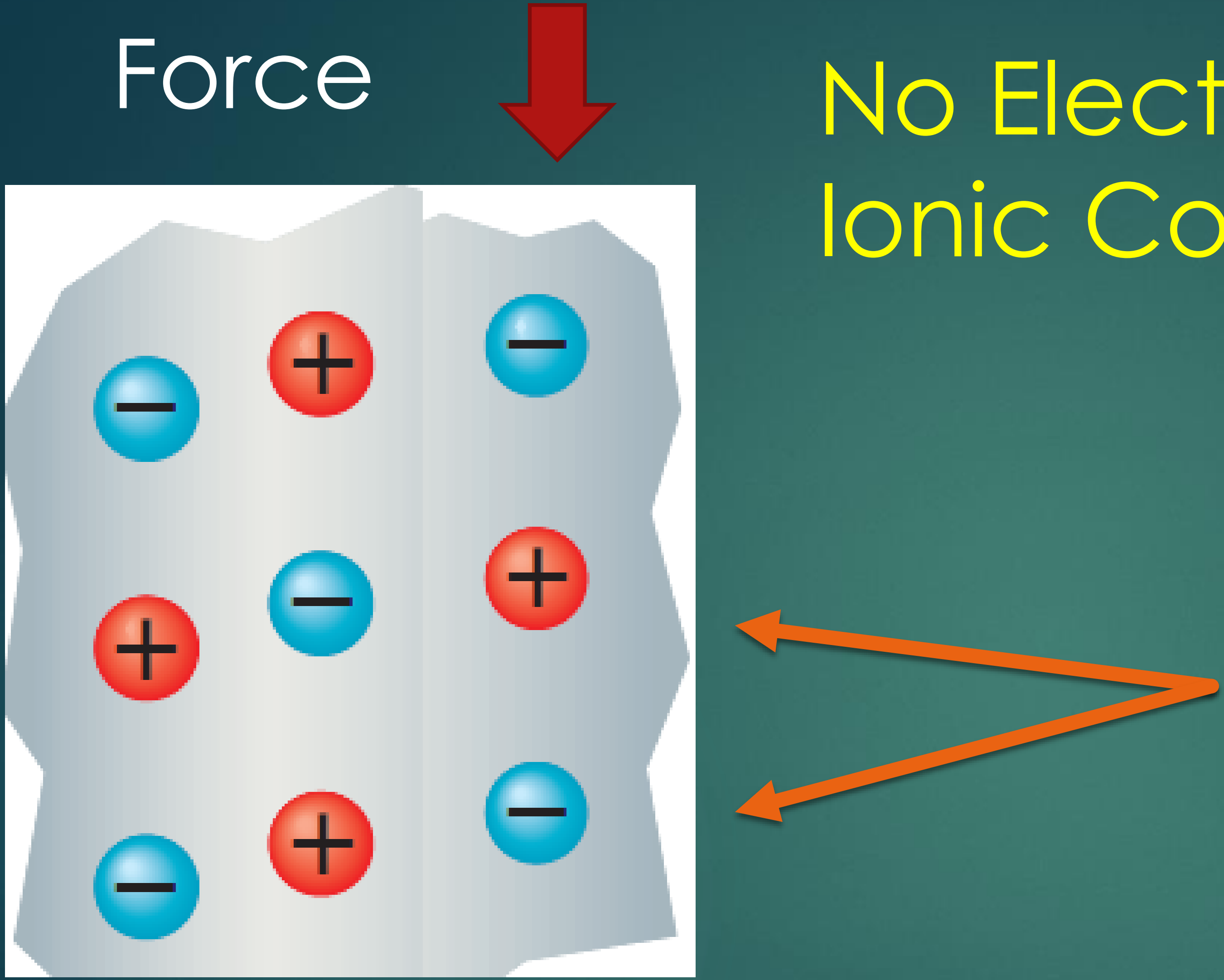


Electron Insulation





No Electron Insulation: Ionic Compounds

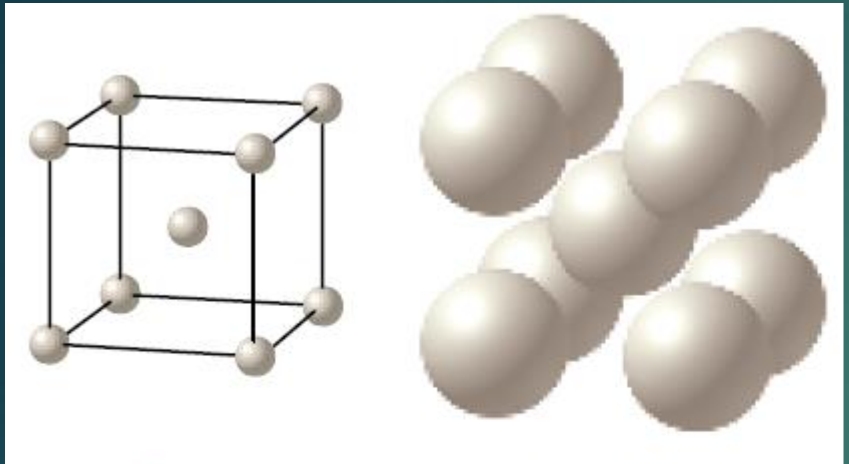


Strong repulsion from like charges

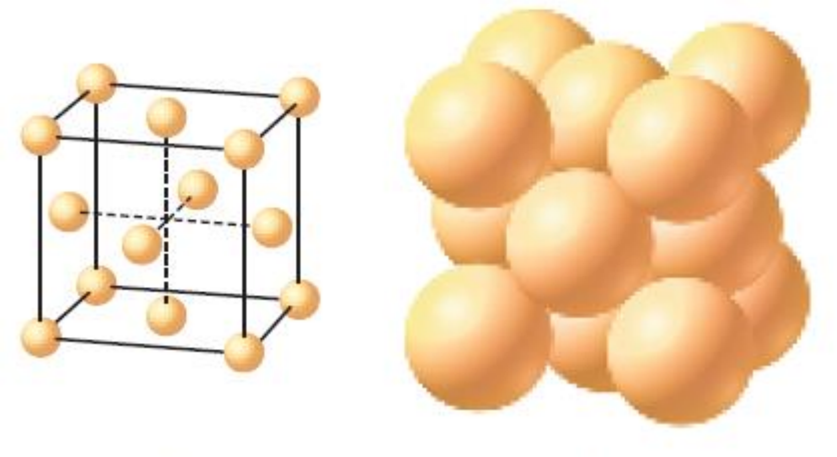
Crystalline Structure of Metals

- Metal atoms are arranged in very compact and orderly patterns.
  - Body-centered cubic
  - Face-centered cubic
  - Hexagonal close-packed

Crystalline Structure of Metals

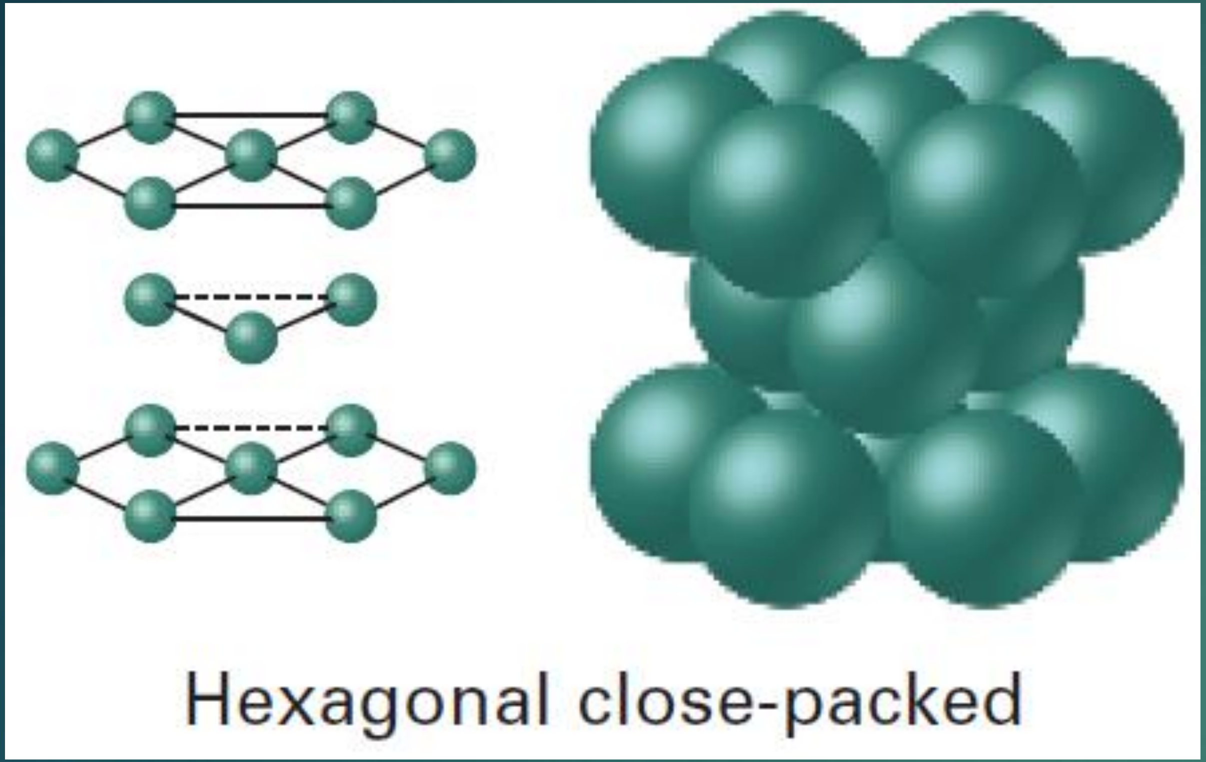


Body-centered cubic



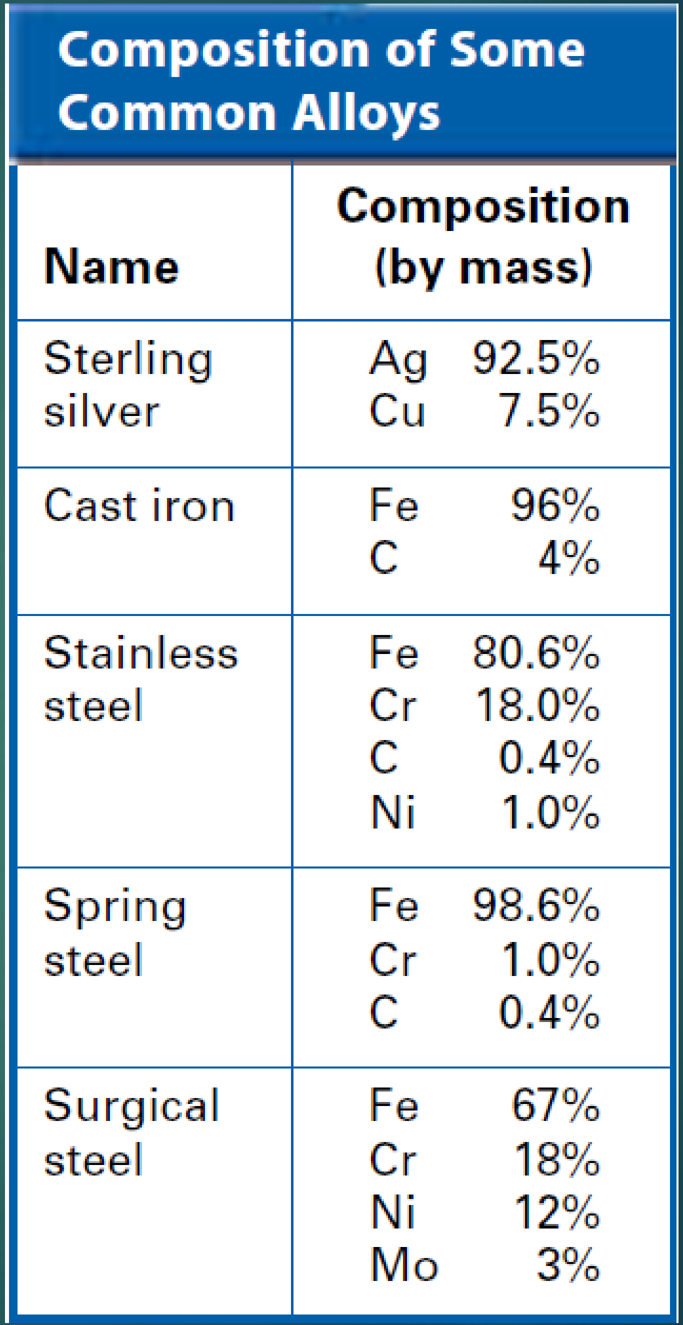
Face-centered cubic

Crystalline Structure of Metals


Hexagonal close-packed

Alloys

- Alloys are physical mixtures composed of two or more elements, at least one of which is a metal.
- Important because their properties are often superior to those of their components.



Composition of Some Common Alloys

Name	Composition (by mass)
Sterling silver	Ag 92.5% Cu 7.5%
Cast iron	Fe 96% C 4%
Stainless steel	Fe 80.6% Cr 18.0% C 0.4% Ni 1.0%
Spring steel	Fe 98.6% Cr 1.0% C 0.4%
Surgical steel	Fe 67% Cr 18% Ni 12% Mo 3%

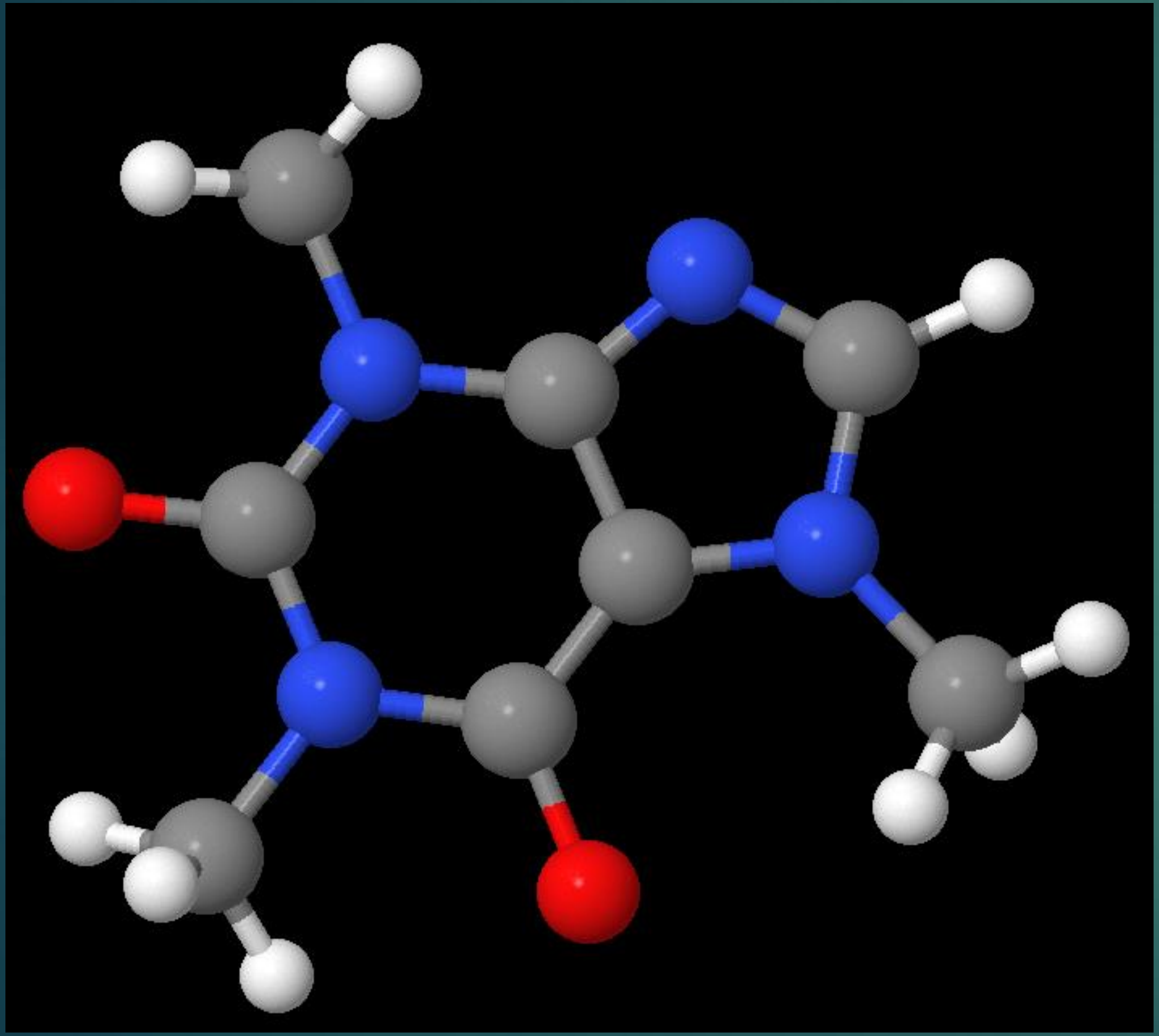
Shorthand e^- Configuration

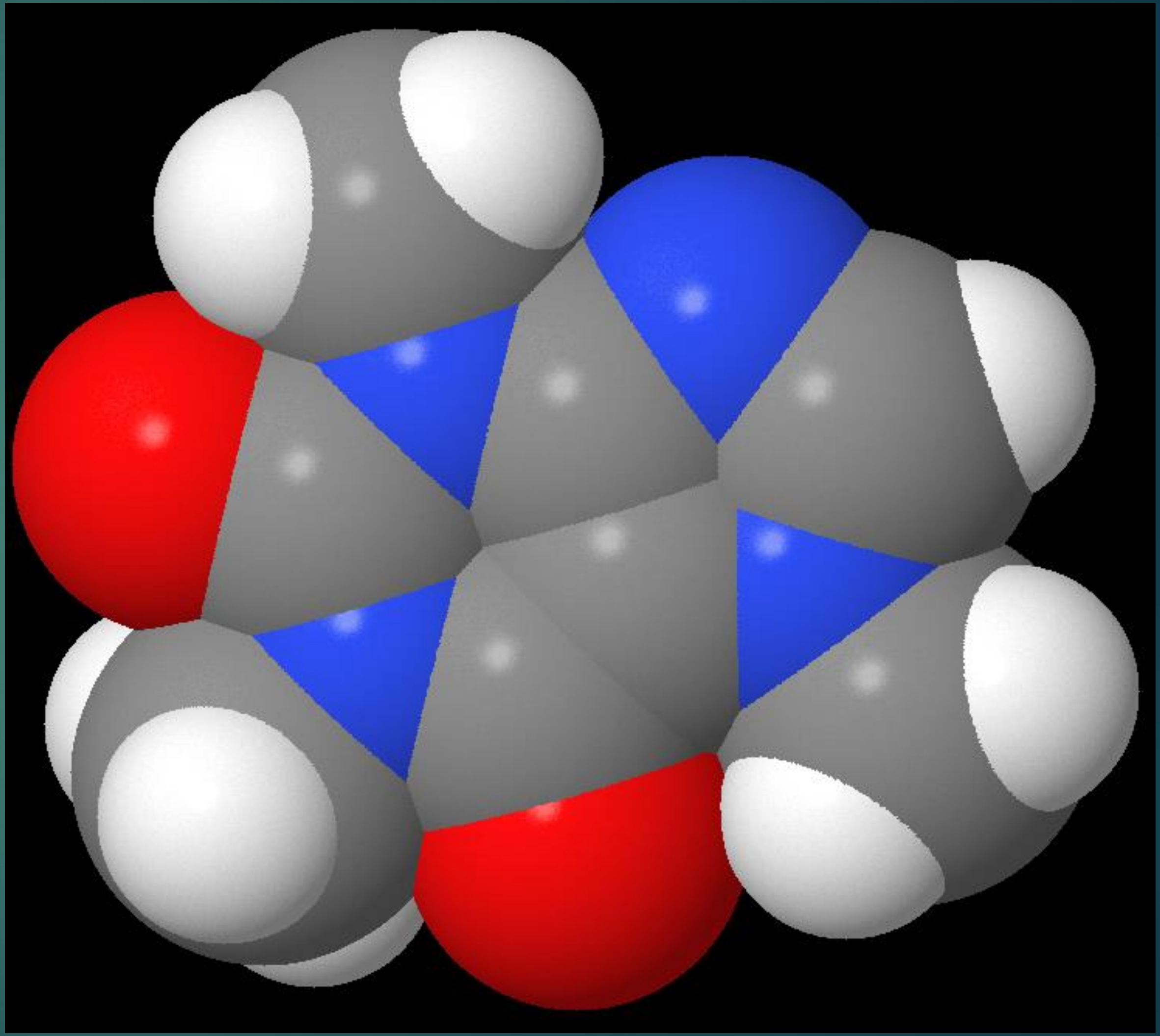
- The first part of the electron configuration is represented by the preceding noble gas in square brackets. Ex. $[Kr]5s^2 4d^8$ would be for Palladium.

$$Pd = \underbrace{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6}_{\text{Krypton's } e^- \text{ Configuration}} 5s^2 4d^8$$

$$Pd = [Kr]5s^2 4d^8$$

Molecular Compounds



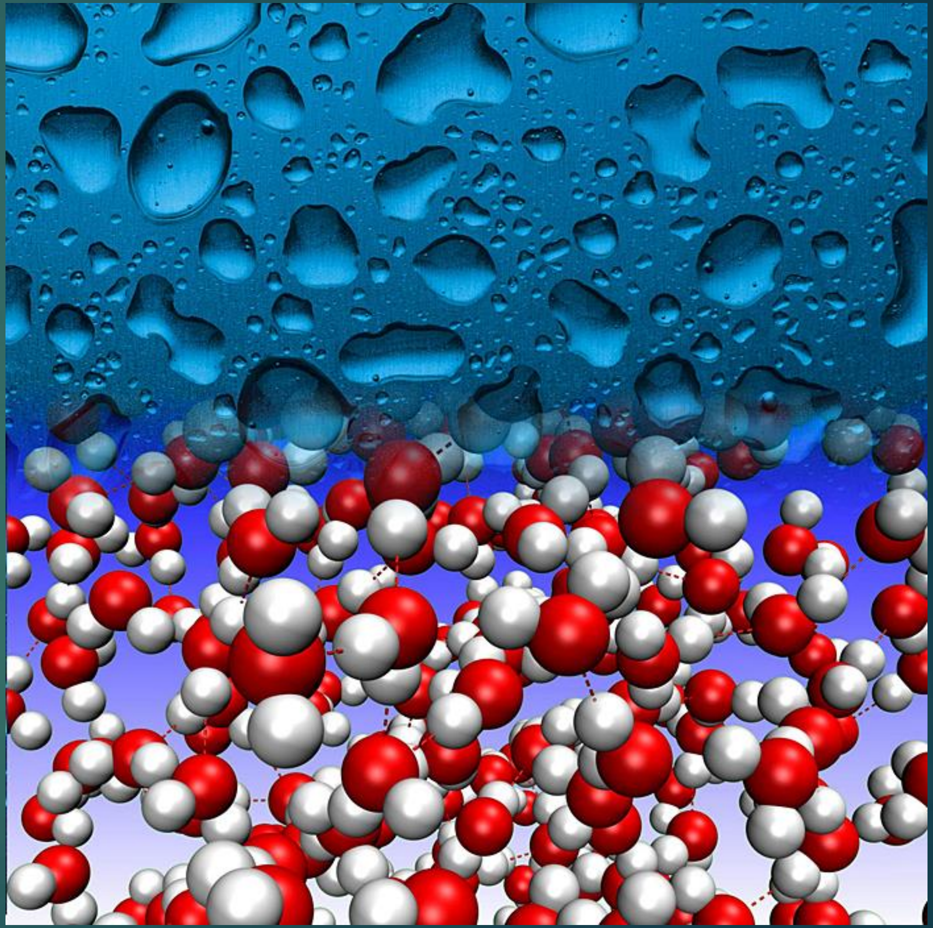


Molecular Compounds

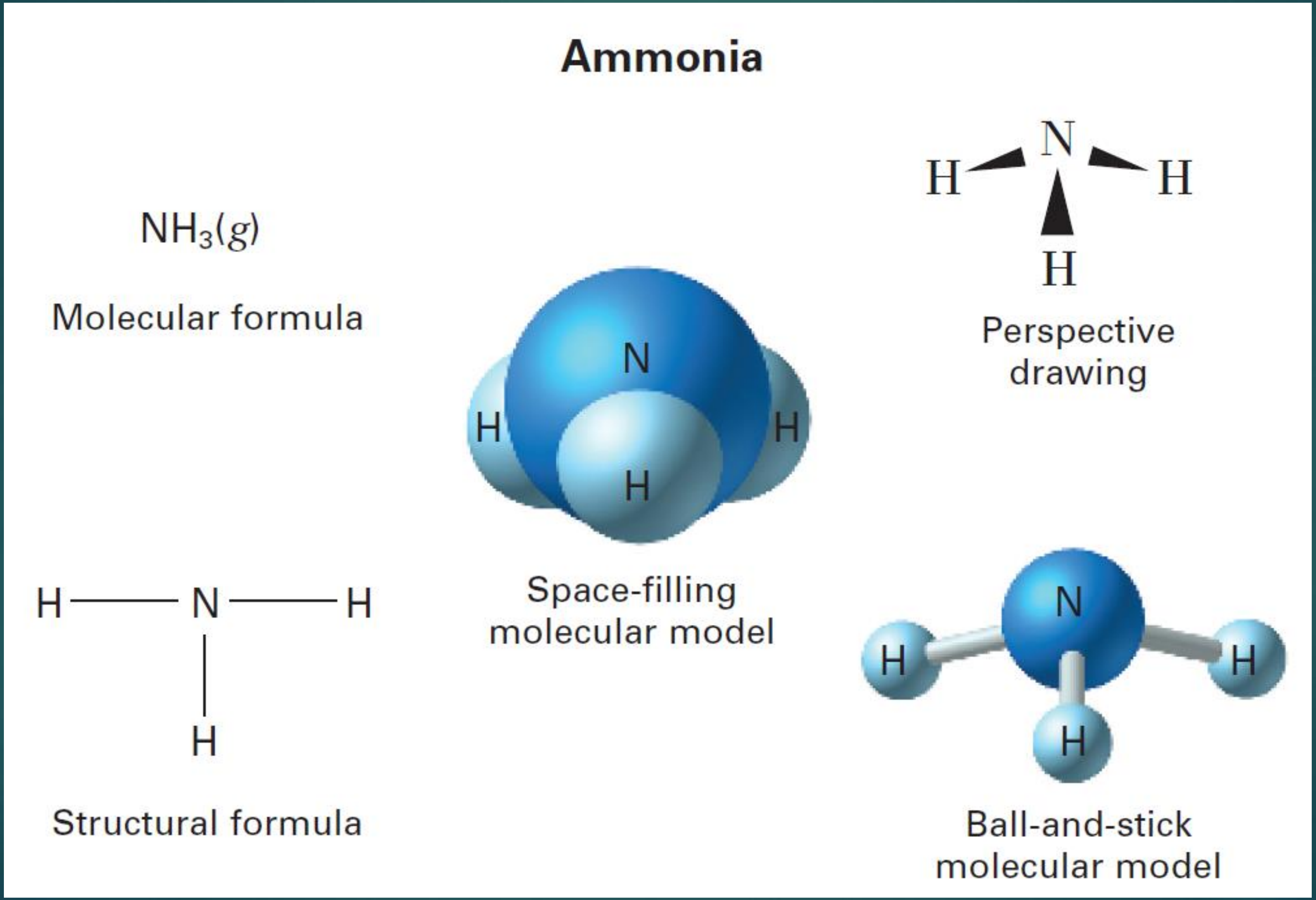
- Formed through the sharing of electrons between non-metal atoms.
- A bond formed through the sharing of electrons is called a covalent bond.
- Exist as individual molecules with a chemical formula.
  - CH_4, CO_2, N_2O, CO

Molecular Compounds

- Relatively low melting and boiling points.
  - Many are liquids or gases at room temperatures.
  - Exist as individual molecules.
- Commonly represented with a molecular formula (how many of each atom). CH_4, N_2O, H_2O



Representing Molecular Compounds

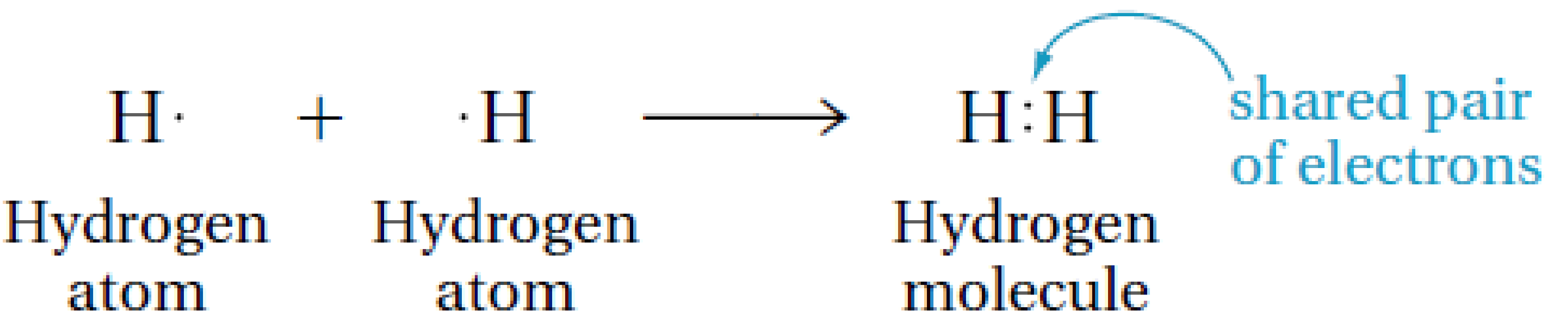


Covalent Bonds

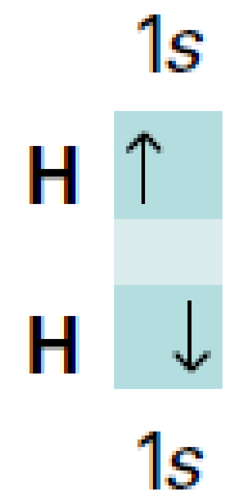
- A covalent bond is formed when electrons are shared between two non-metals.
- Occurs so that atoms can attain the electron configuration of a noble gas (octet rule).
- Single covalent bond
- Double covalent bond (second strongest)
- Triple covalent bond (strongest)
- Coordinate covalent bond

Single Covalent Bond

- Two atoms held together by share one pair of electrons.
  - Each atom supplies one electron for a total of two.

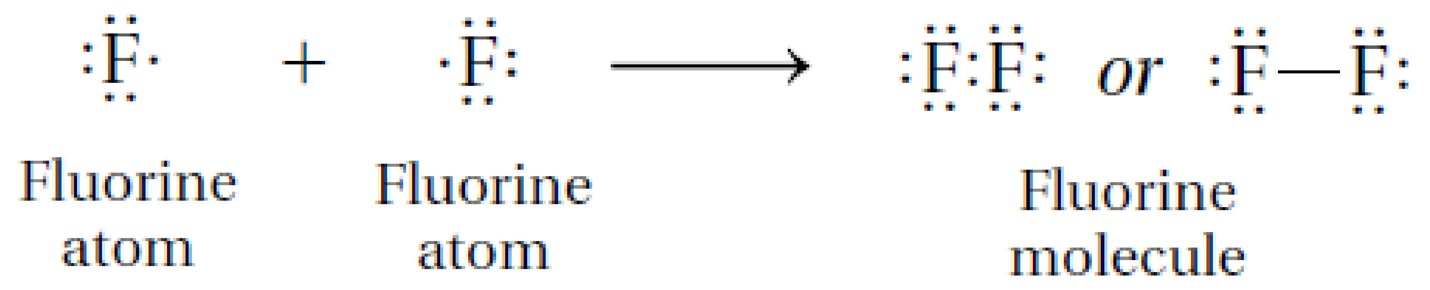


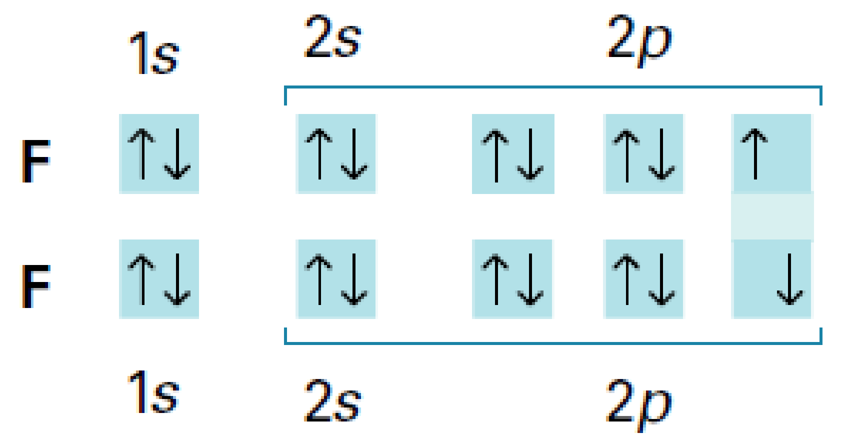
Hydrogen molecule

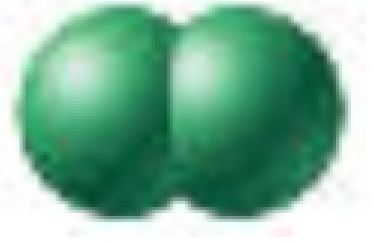


Single Covalent Bond: Structural Formula

- Fluorine needs one more electron to complete the octet; which it can get from another fluorine atom.
- Structural formula: Covalent bonds are represented as dashed lines.

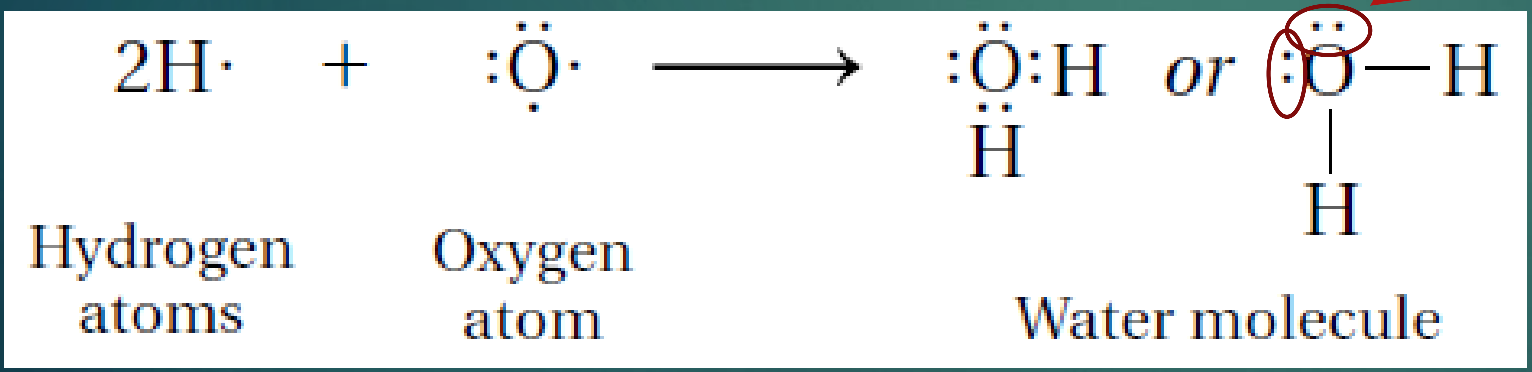


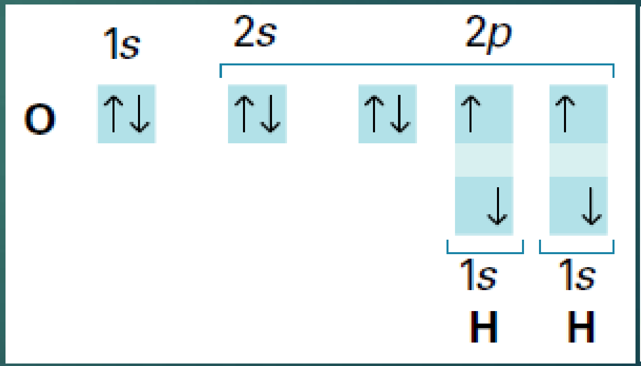



Fluorine molecule

Unshared Pairs of Electrons

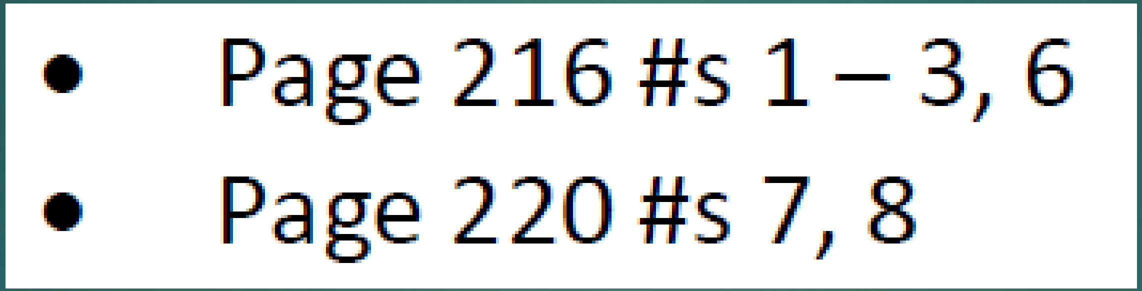
- Atoms want to have 8 electrons, but they don't all have to partake in covalent bonding.
- Such pairs of electrons are called unshared pairs.
- Influence other aspects of the molecule like its shape and interaction with its brethren molecules.






Water molecule

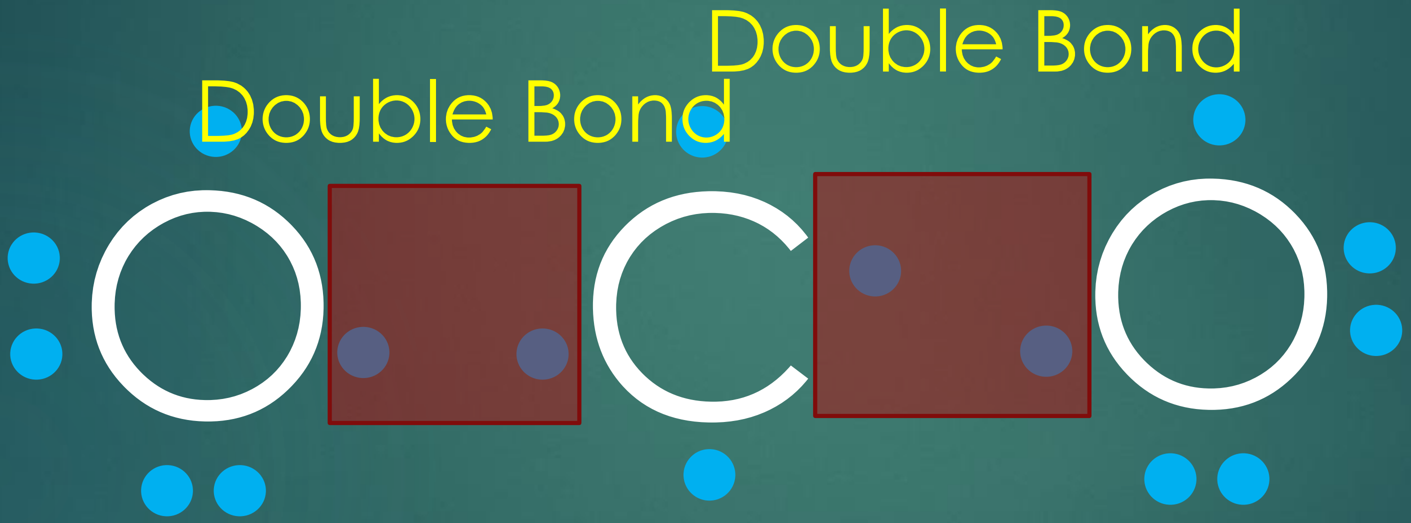
Review



- Page 216 #s 1 – 3, 6
- Page 220 #s 7, 8

Double Covalent Bond

- Atoms share two electrons each, for a total of four, creating two pairs of electrons. Ex: CO_2



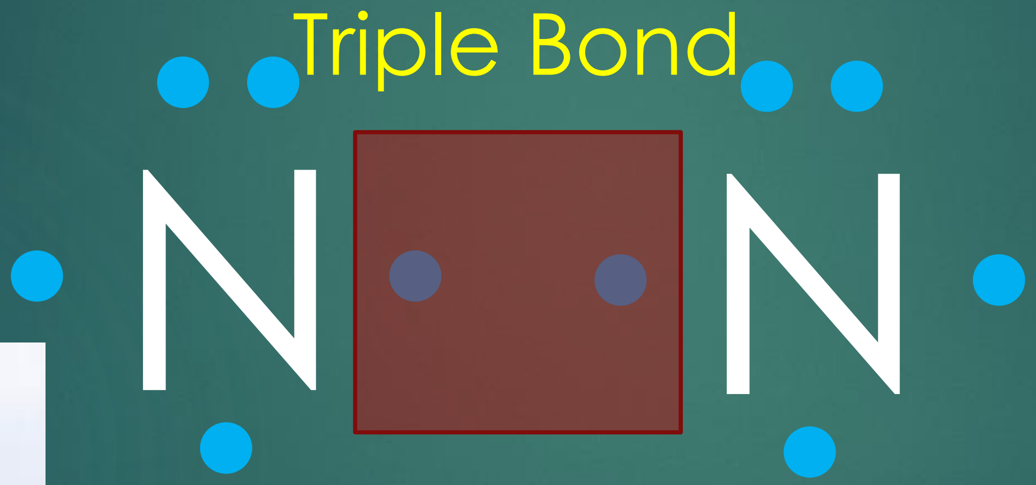
Double Bond

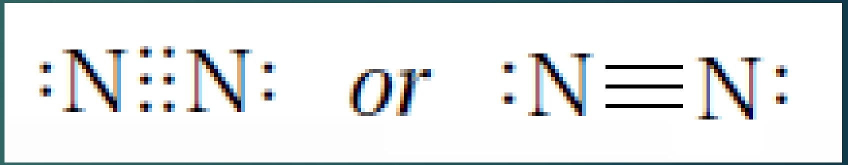
Double Bond

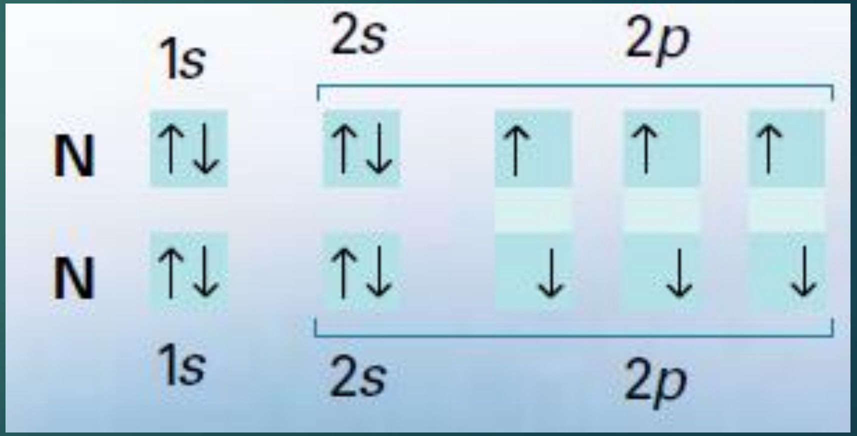
- Remember, only unpaired electrons will be contributed from each atom to form a bond.

Triple Covalent Bond

- Two atoms each provide three electrons, for a total of six, to create the triple bond.




:N⋮⋮N: *or* :N≡N:

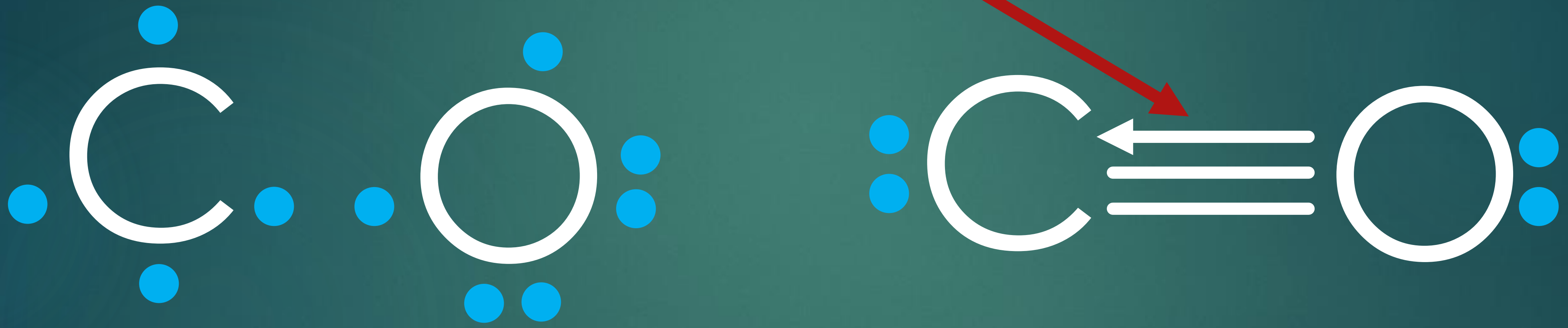


	1s	2s	2p		
N	↑↓	↑↓	↑	↑	↑
N	↑↓	↑↓	↓	↓	↓
	1s	2s	2p		


Nitrogen molecule

Coordinate Covalent Bond

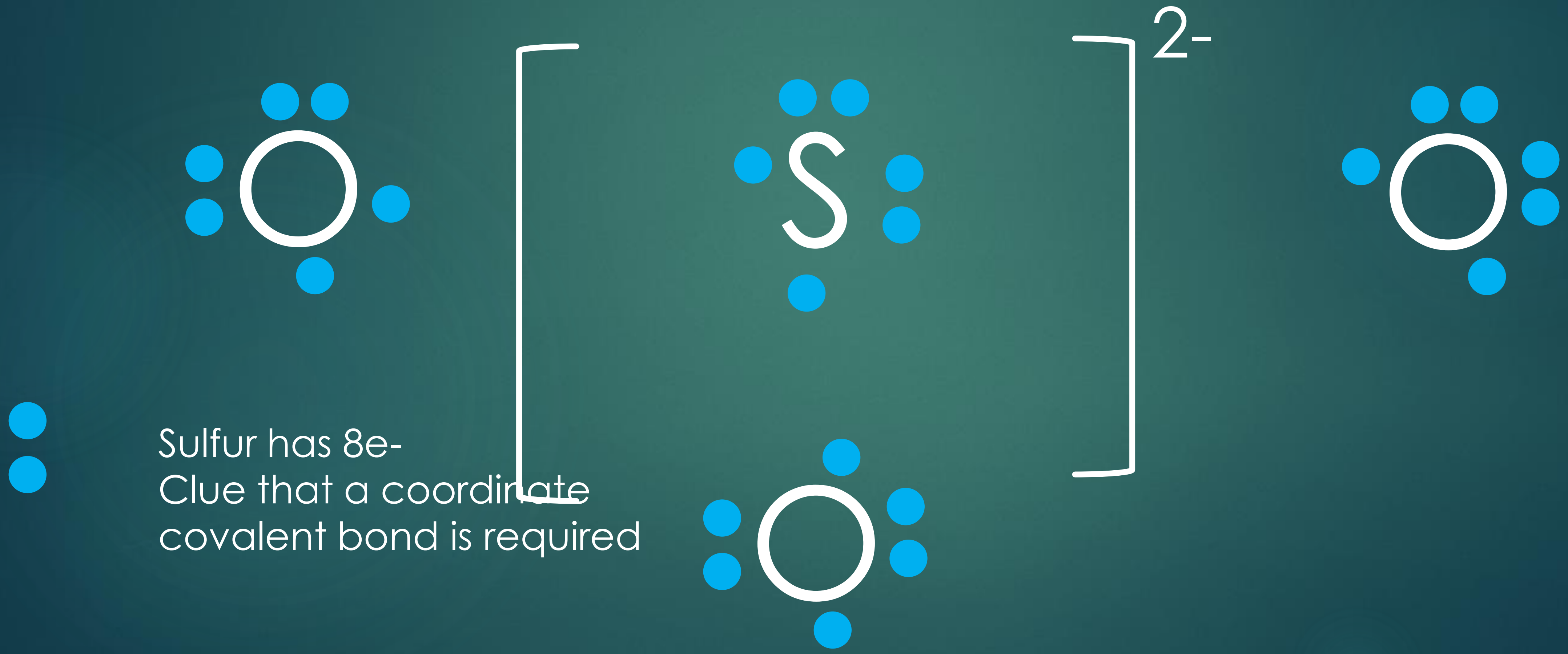
- A bond in which one atom contributes both bonding electrons required for a covalent bond.



- Oxygen has no more unpaired electrons.
- However, being awesome, it will share a full pair!

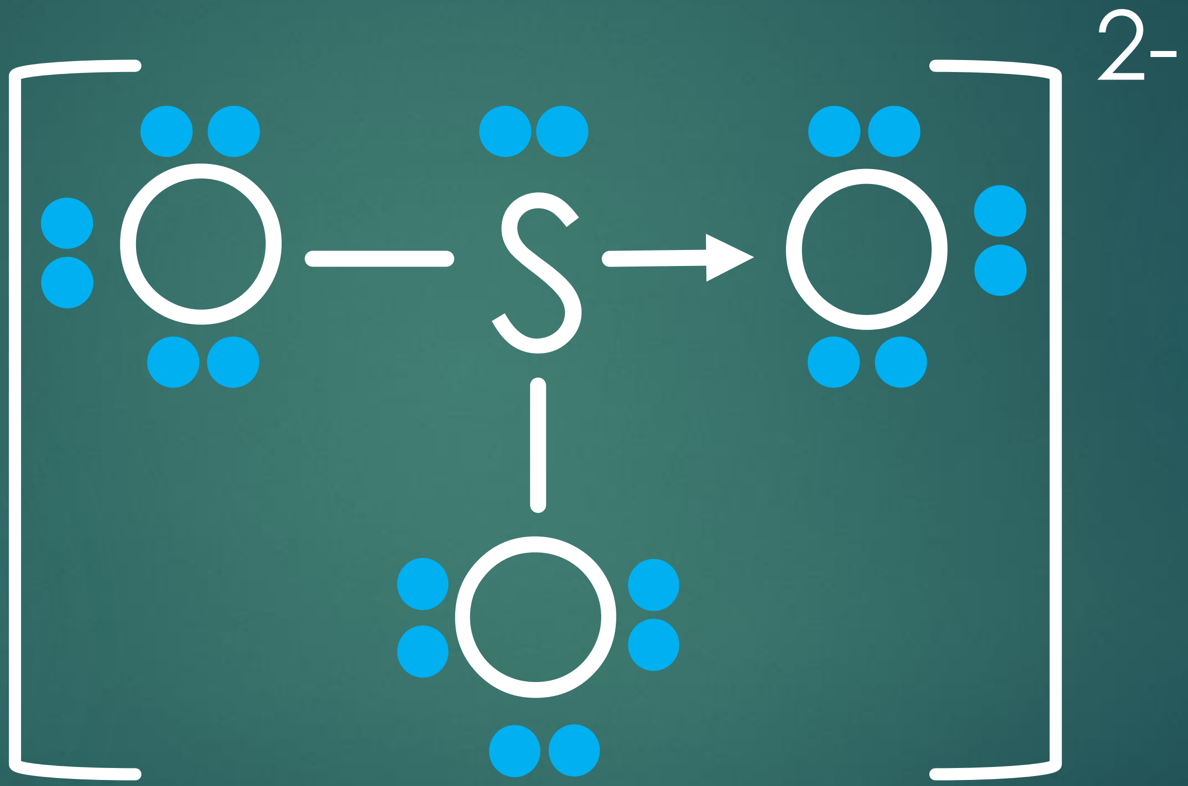
Polyatomic Ions

- Atoms that covalently bond together to have an ionic charge. Ex: SO_3^{2-}, often have a coordinate covalent bond.



Sulfur has 8e-
Clue that a coordinate covalent bond is required

Polyatomic Ion: SO_3^{2-} Structural Diagram



Structural Diagram Algorithm

- Formula to calculate how many bonds are needed:
  - Have = Add up all the valence electrons including any electrons gained or lost by a polyatomic ion.
  - Want = Count up the total # of electrons to achieve a noble gas configuration (2 for hydrogen, 8 for all other elements).
  - Subtract the Want - Have; then divide by two (each bond is two electrons)
  - The answer is how many bonds to draw.
  - Fill in the rest of the valence electrons, don't forget any extra electrons if it is a polyatomic ion.

Example: Structural Diagram for HCN (Cyanic Acid)

- Given that C is the central atom.
- N – has 5 valence electrons, wants 8.
- C – has 4 valence electrons, wants 8.
- H – has 1 valence electron, wants 2.
- Have = 5 + 4 + 1 = 10
- Want = 8 + 8 + 2 = 18
- Math: (Want – Have)/2
  - = (18 – 10)/2 = 4 bonds required

Structural Diagram for HCN

- The molecule requires 4 bonds

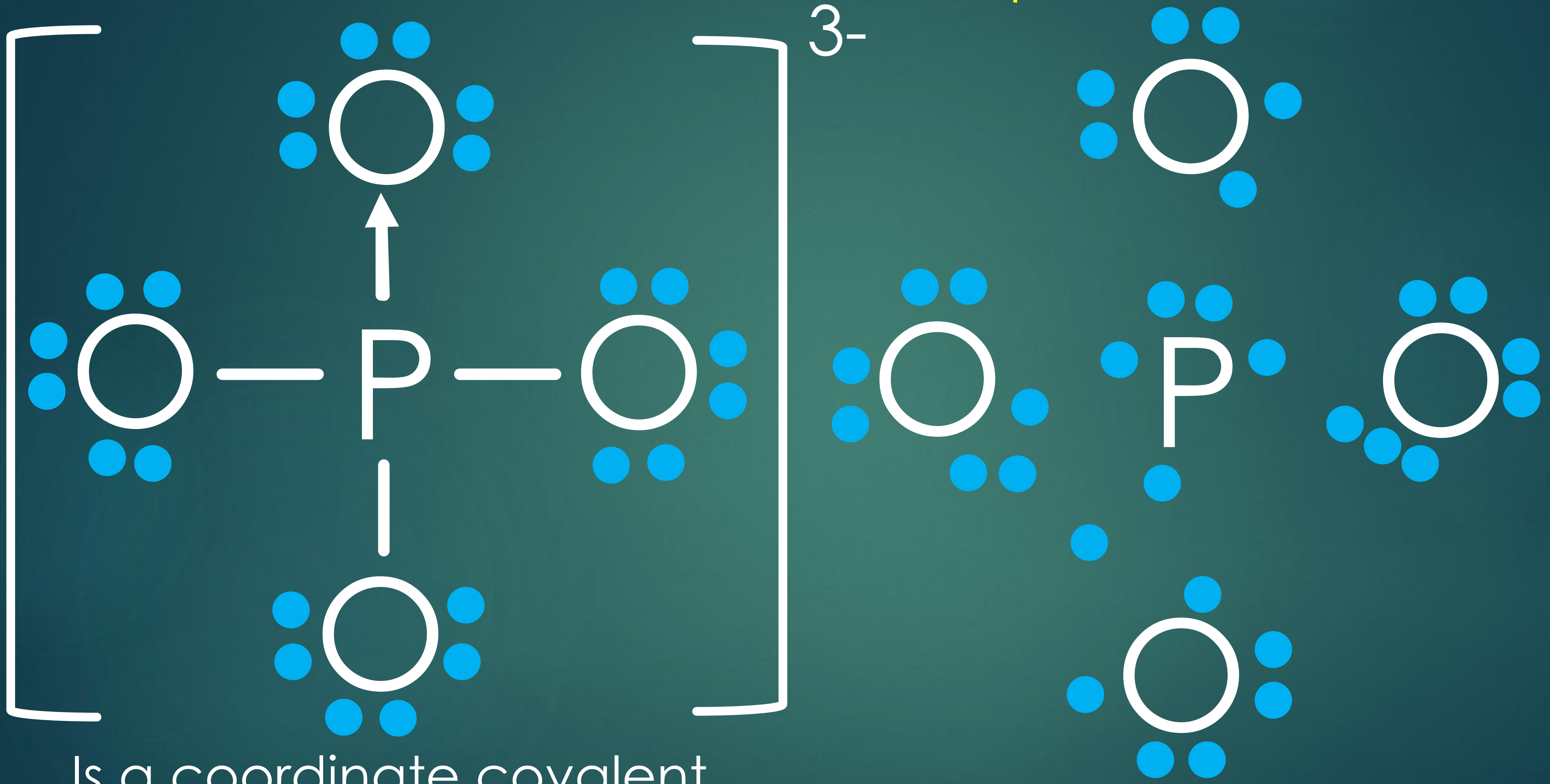
$$\mathrm{H-C\equiv N:}$$

- H can’t have more than two electrons
- Other bonds must be between C & N
- Add remaining electrons
- Check for coordinate covalent bonds

Structure Diagram for PO_4^{3-}

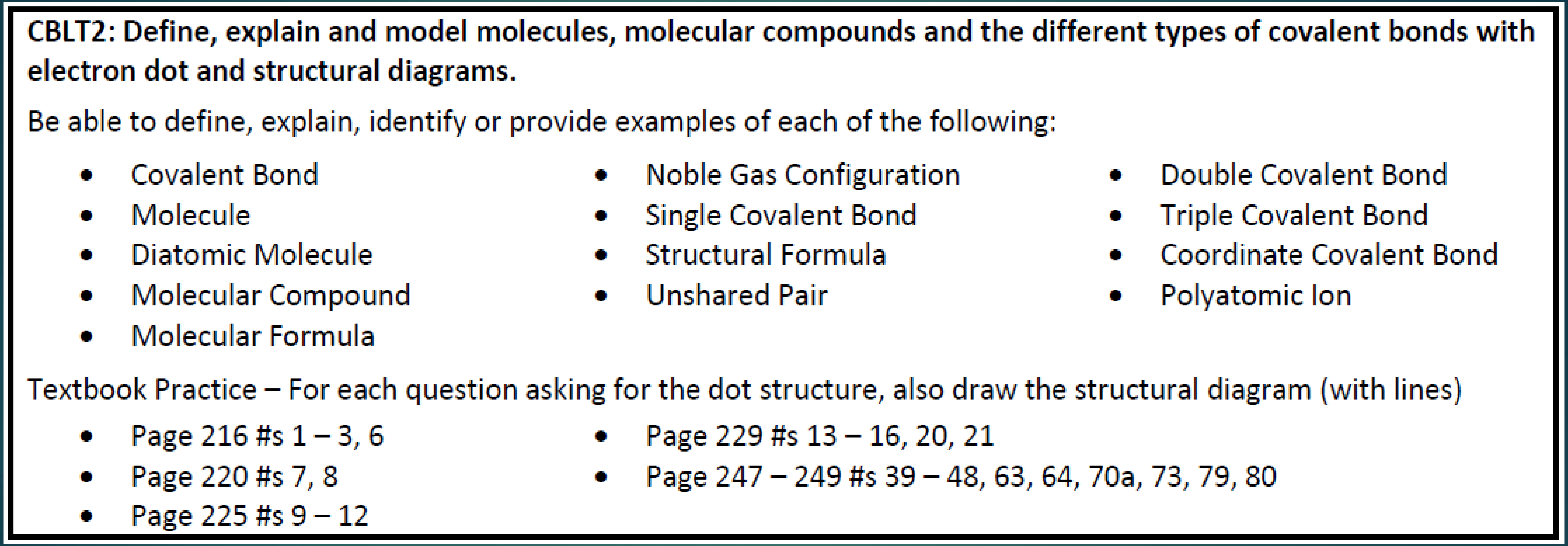
- Given that P is the central atom (surrounded by all O).
- P – has 5 valence electrons, wants 8
- O – has 6 valence electrons x 4, wants 8 x 4
- Anion Charge – 3 additional electrons
- Have = 5 + 6x4 + 3= 32
- Want = 8 + 8x4 = 40
- Math: (Want – Have)/2
  - = (40 – 32)/2 = 4 bonds required

Structure Diagram for PO_4^{3-}



Is a coordinate covalent bond necessary?

Learning Target Review



CBLT2: Define, explain and model molecules, molecular compounds and the different types of covalent bonds with electron dot and structural diagrams.

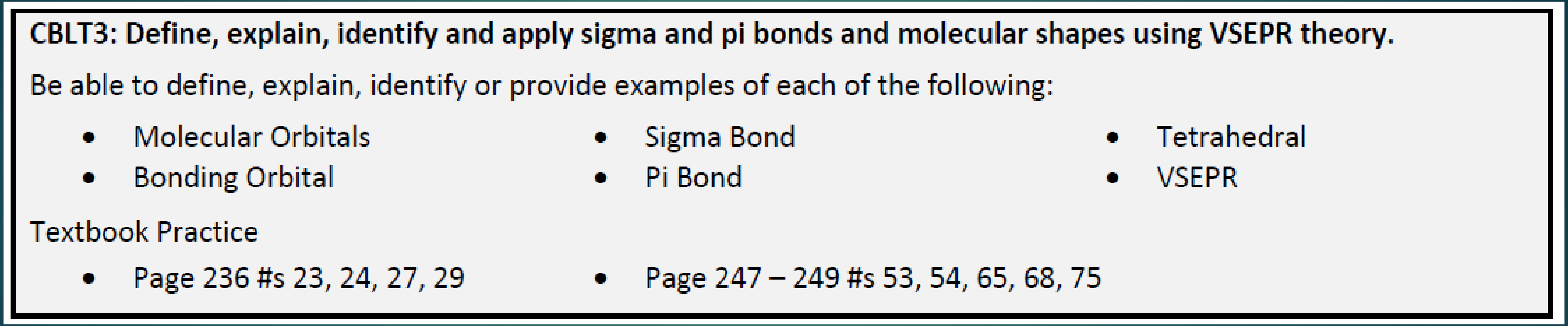
Be able to define, explain, identify or provide examples of each of the following:

- Covalent Bond
- Molecule
- Diatomic Molecule
- Molecular Compound
- Molecular Formula
- Noble Gas Configuration
- Single Covalent Bond
- Structural Formula
- Unshared Pair
- Double Covalent Bond
- Triple Covalent Bond
- Coordinate Covalent Bond
- Polyatomic Ion

Textbook Practice – For each question asking for the dot structure, also draw the structural diagram (with lines)

- Page 216 #s 1 – 3, 6
- Page 220 #s 7, 8
- Page 225 #s 9 – 12
- Page 229 #s 13 – 16, 20, 21
- Page 247 – 249 #s 39 – 48, 63, 64, 70a, 73, 79, 80

Molecular Compounds
Bonding Theories and Shapes



CBLT3: Define, explain, identify and apply sigma and pi bonds and molecular shapes using VSEPR theory.

Be able to define, explain, identify or provide examples of each of the following:

- Molecular Orbitals
- Bonding Orbital
- Sigma Bond
- Pi Bond
- Tetrahedral
- VSEPR

Textbook Practice

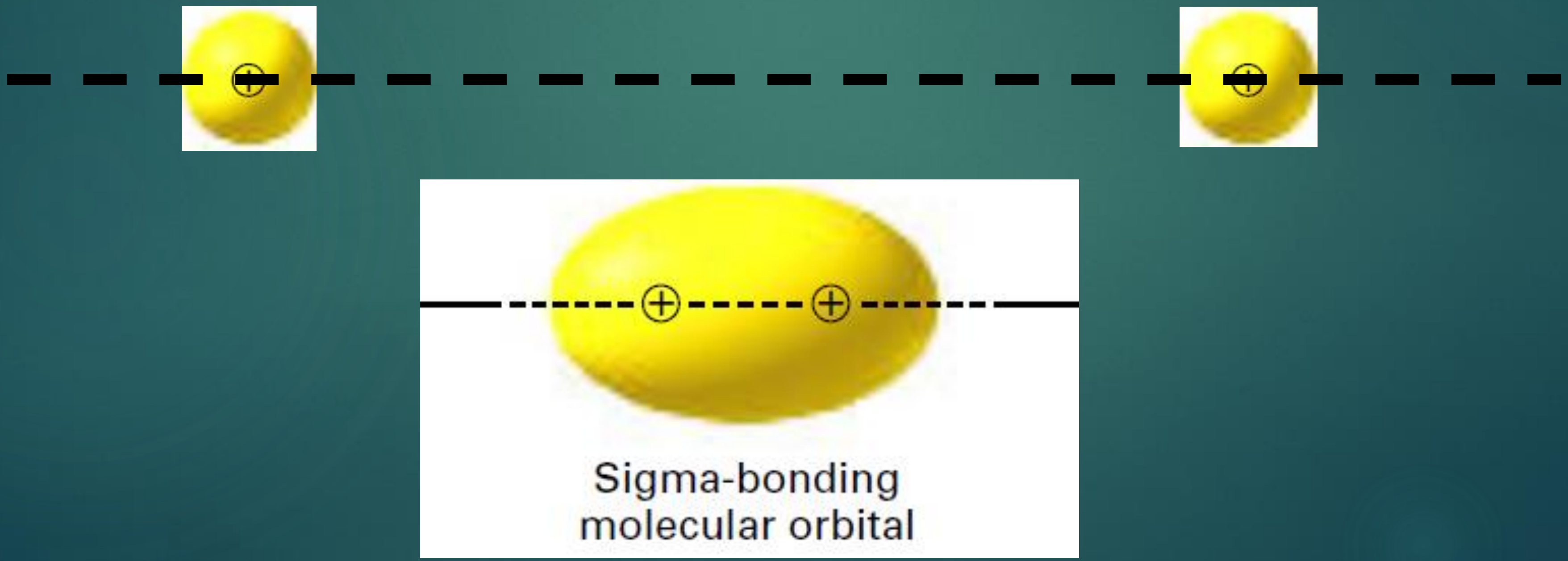
- Page 236 #s 23, 24, 27, 29
- Page 247 – 249 #s 53, 54, 65, 68, 75

Bonding Theories

- Covalent bonds are created when electron orbitals overlap. The overlapping orbitals create ***molecular orbitals***.
  - Molecular orbitals belong to the entire molecule, not just a particular atom.
  - A molecular orbital that can be occupied by two electrons of a covalent bond is called a ***bonding orbital***.
- There are two types of bonds that create molecular orbitals:
  - Sigma bonds, σ-bonds
  - Pi bonds, π-bonds

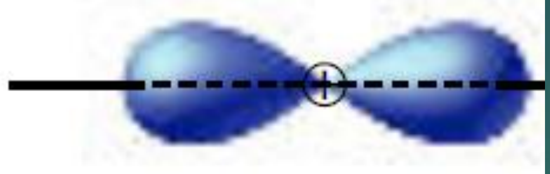
Sigma Bonds

- Occurs when two atomic orbitals combine to form a molecular orbital along the bonding axis connecting the two atomic nuclei.
- Can occur with two *s*-orbitals, two *p* or an *s* and *p*-orbital.

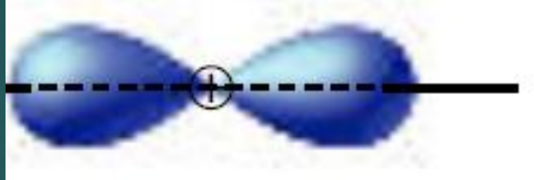


Bonding axis

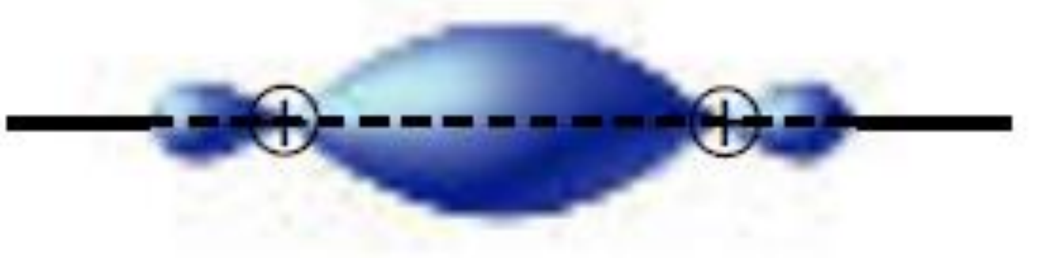
Sigma Bonds: With *p*-orbitals



p atomic orbital



p atomic orbital

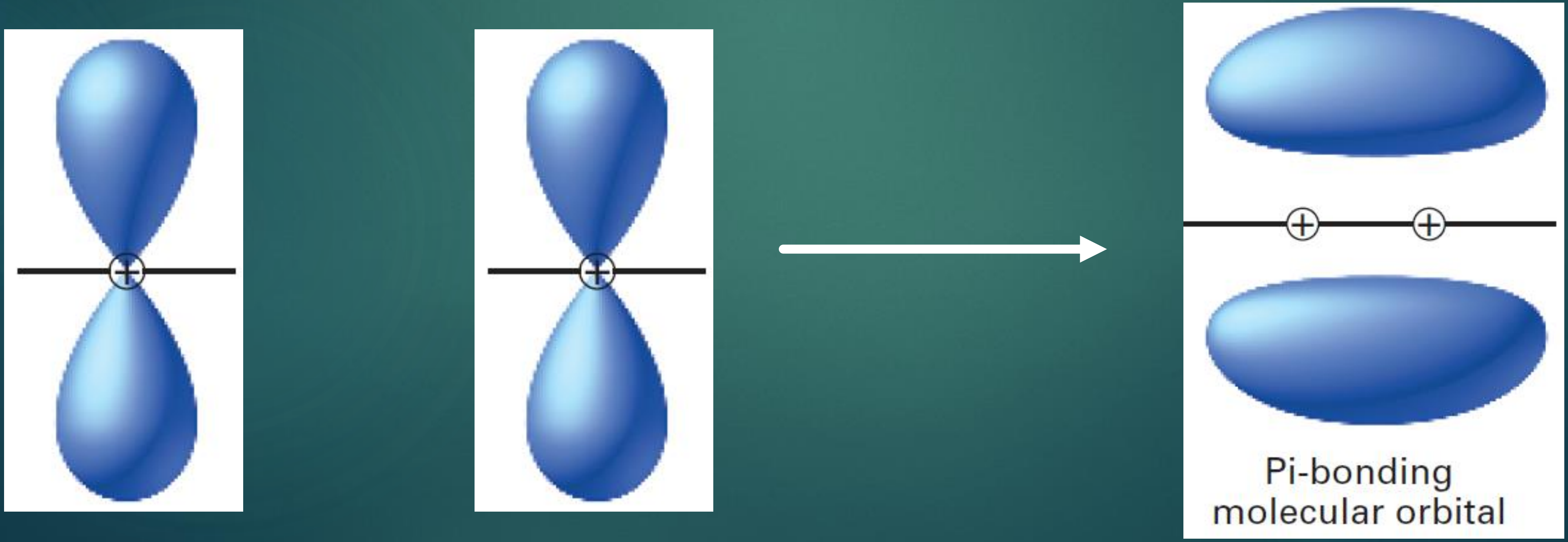


Sigma-bonding molecular orbital

Like atomic orbitals, they are regions with a high probability of finding an electron around the molecule.

Pi Bonds

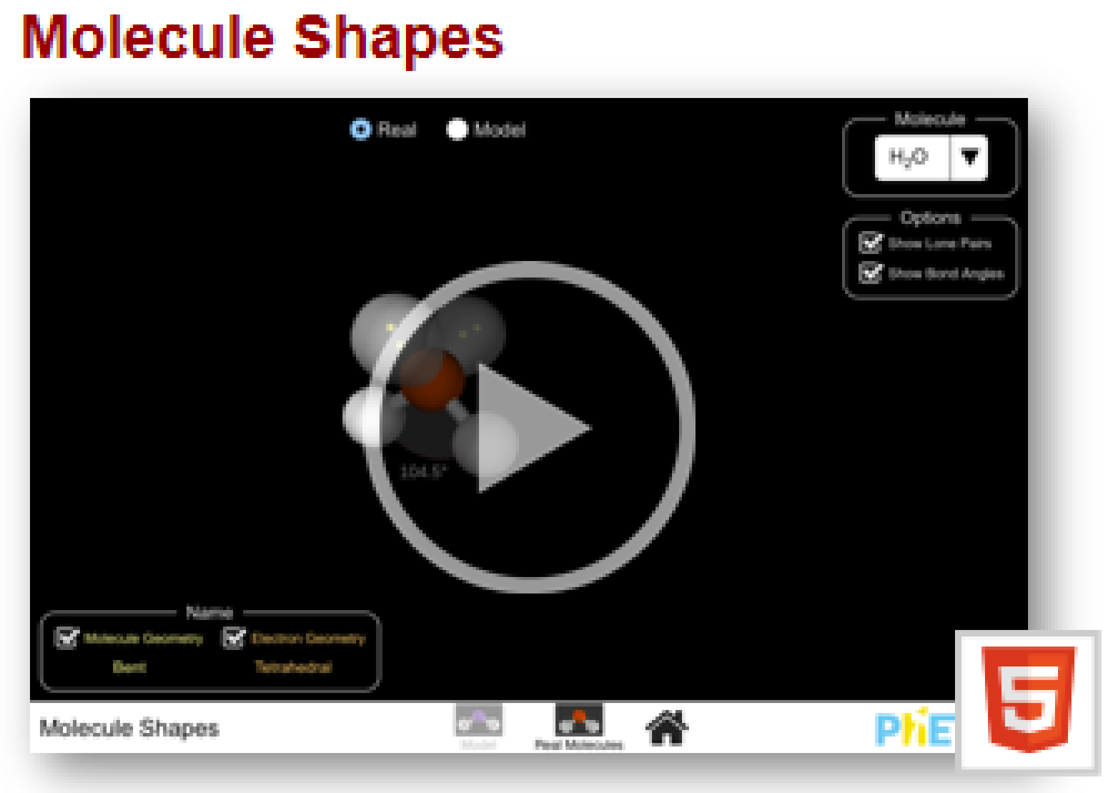
- The electrons are most likely to be found in bean-shaped regions above and below the bonding axis.
  - p-orbitals only.
  - Pi bonds overlap less than sigma bonds, so pi bonds tend to be weaker.



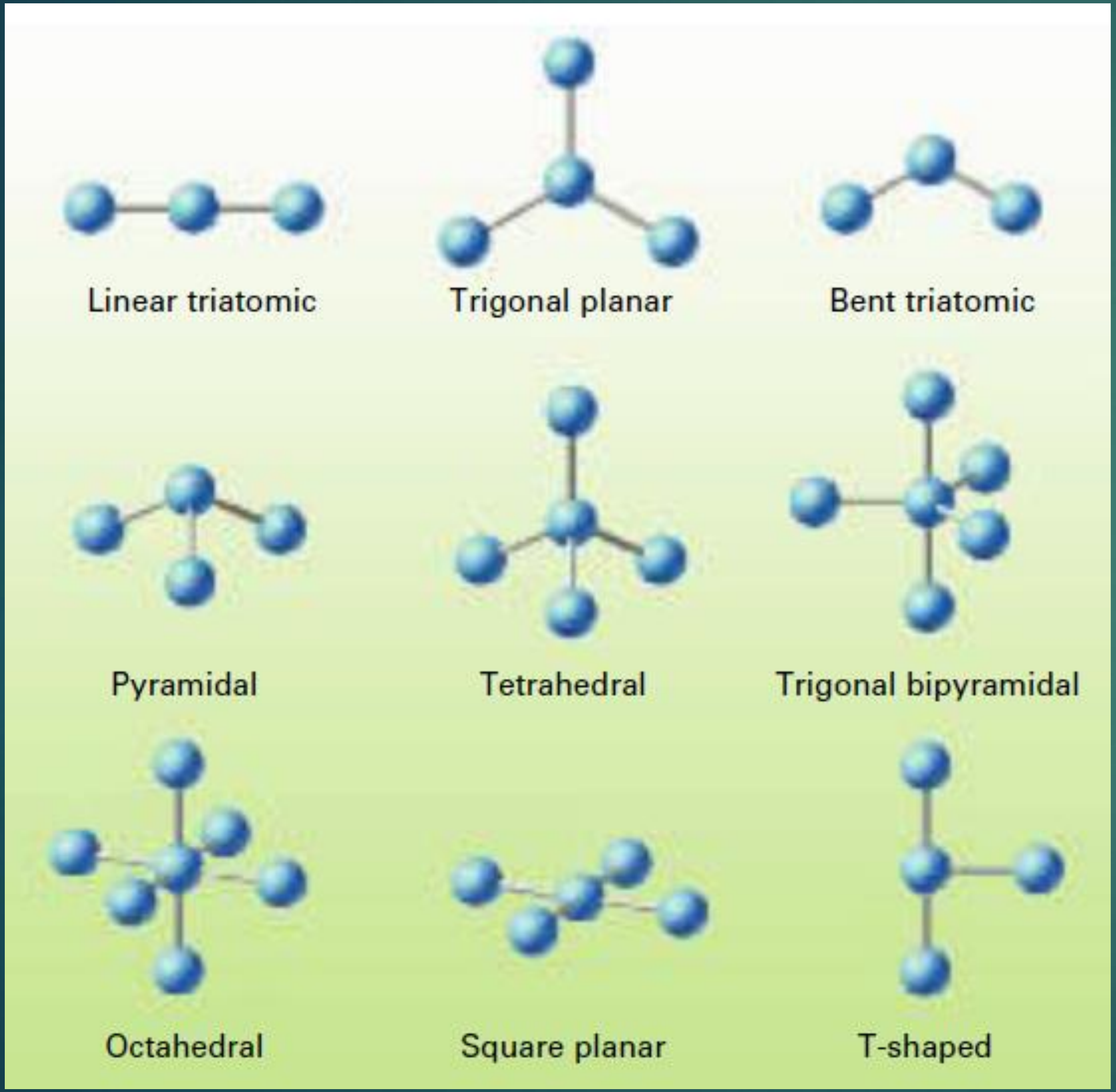
VSEPR Theory

- **V**alence **S**hell **E**lectron-**P**air **R**epulsion
  - Theory governing the shape a molecule takes.
  - Their shape plays a major role in physical properties such as melting and boiling points.
- **VSEPR Theory**: The repulsion between lone electron pairs causes molecular shapes to adjust so that the valence-electron pairs stay as far apart as possible.

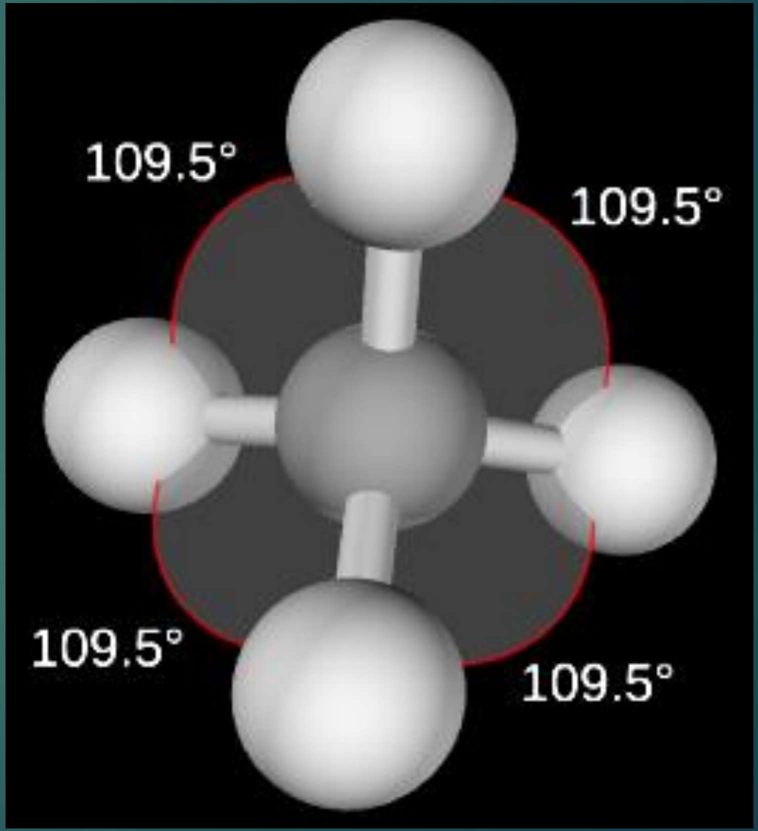
Visualizing e^- Pairs: Molecular Shapes



Types of Shapes



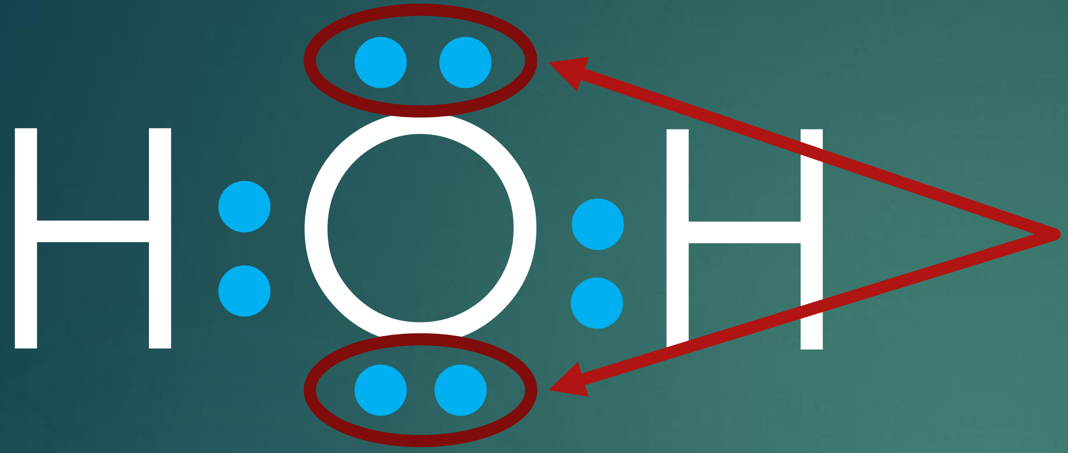
Tetrahedral



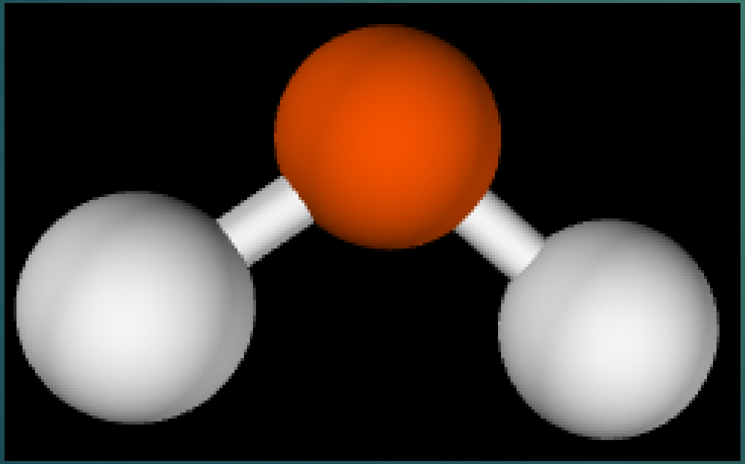
Determining Molecular Shapes: Gr. 11

- For a given molecule, draw its structural diagram and locate the electron pairs that are not part of a covalent bond.
  - If lone pairs of electrons exist on the central atom, the molecule will bend, otherwise it will not.
  - Count how many atoms in the molecule and determine the shape.
- **Note**, we will use this only for linear triatomic, bent triatomic, trigonal planar, or pyramidal.
  - Many molecules do not obey the octet rule for all atoms. So how they bond together is more complicated and hence we will stick to octet obeying molecules.

Bent vs Linear: H_2O vs CO_2

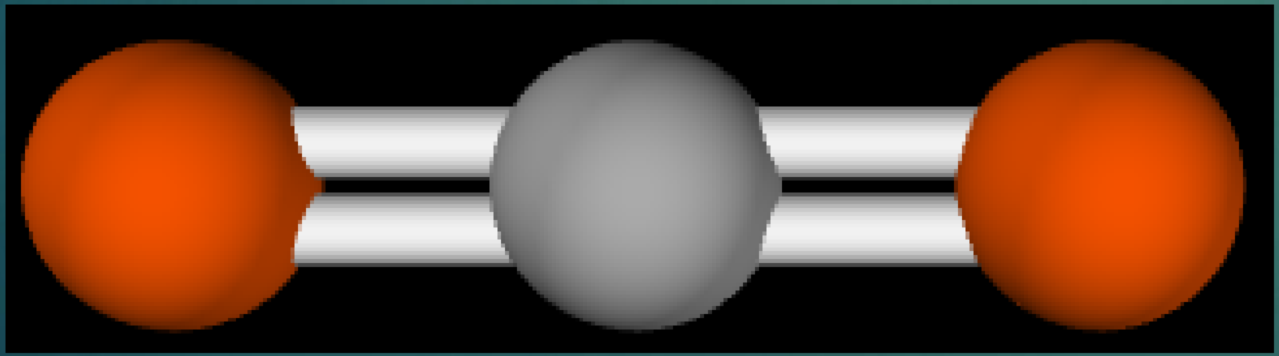


Lone Pairs:
Molecule will bend.

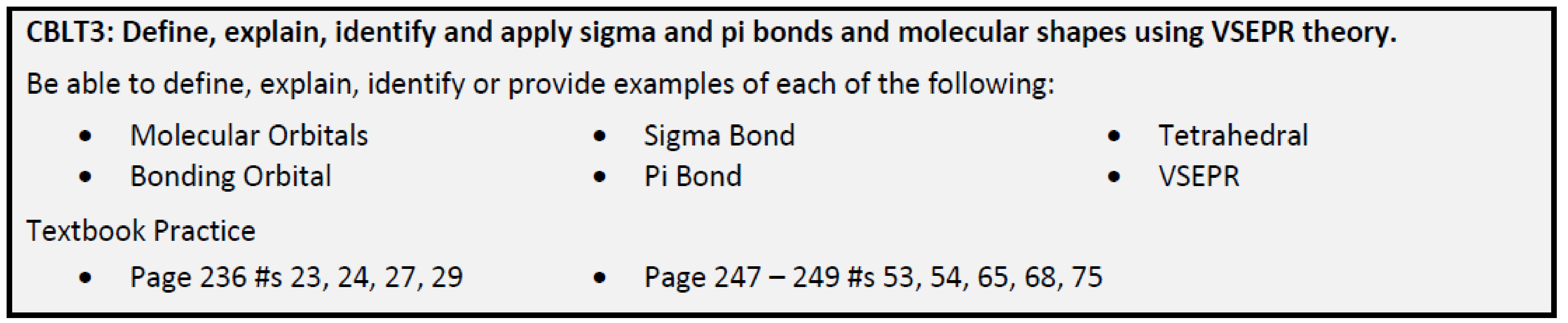


Bent vs Linear: H_2O vs CO_2

No Lone Pairs on C:
Molecule will not bend.



CBLT3 Review



CBLT3: Define, explain, identify and apply sigma and pi bonds and molecular shapes using VSEPR theory.

Be able to define, explain, identify or provide examples of each of the following:

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- Sigma Bond
- Pi Bond
- Tetrahedral
- VSEPR

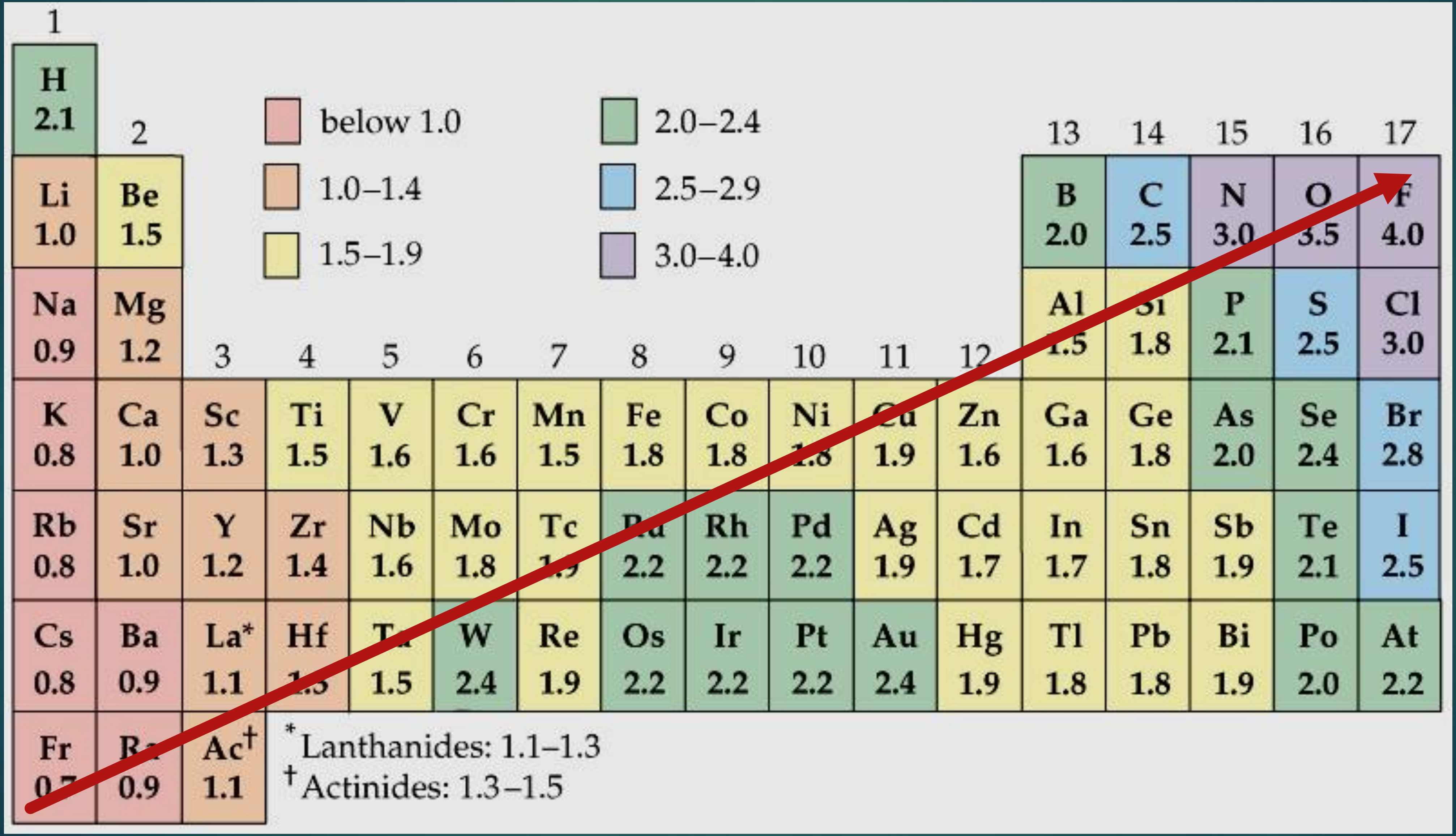
Textbook Practice

- Page 236 #s 23, 24, 27, 29
- Page 247 – 249 #s 53, 54, 65, 68, 75

Molecular Compounds:
Polar Bonds and Their Applications

- When two different atoms bond covalently, there could be an ***unequal sharing*** of the electrons.
- The more ***electronegative*** atom will have a stronger pull on the electrons and will acquire a slight negative charge.
  - Electronegativity is the ability of an atom to attract shared electrons to itself. It is a number from 0 to 4.
- Such a bond is called a ***polar bond***.

Periodic Table: Electronegativity



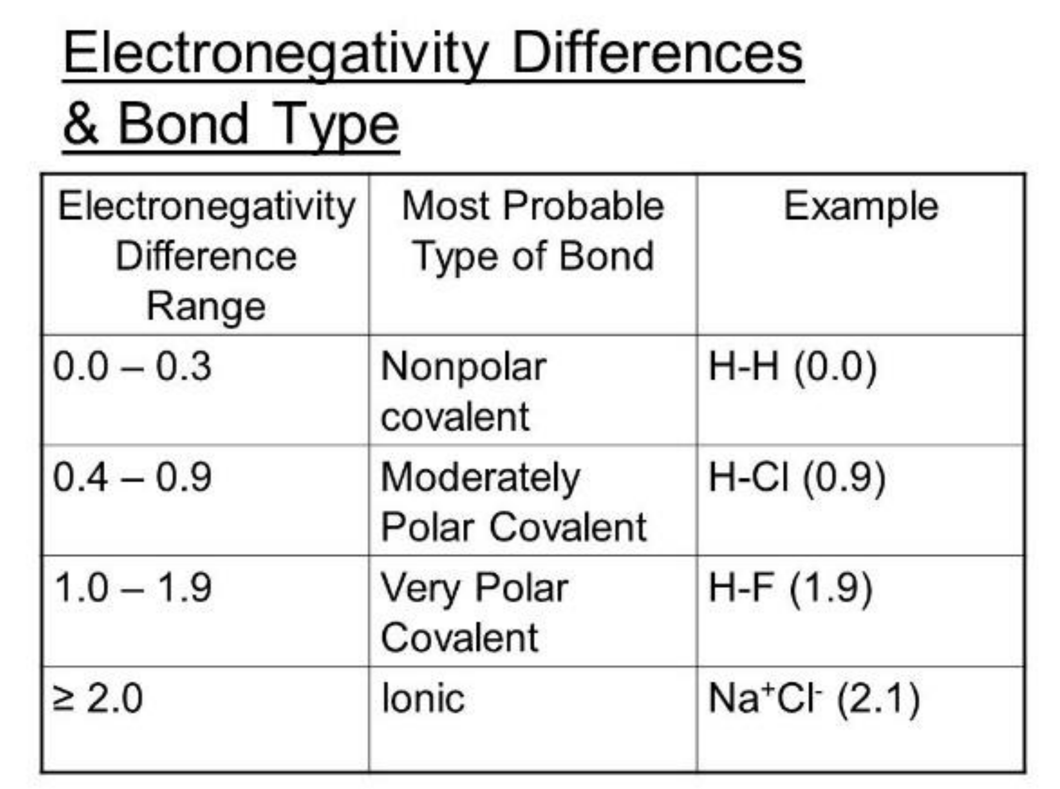
🟥 below 1.0
🟧 1.0–1.4
🟨 1.5–1.9
🟩 2.0–2.4
🟦 2.5–2.9
🟪 3.0–4.0

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

Determining Bond Type

- Use the periodic table of electronegativity to find the electronegativity of each atom and subtract lowest from the highest (have a positive answers).
- Use the chart to determine the bond type.
- Molecular polarity effects the physical state of molecules.



Electronegativity Differences & Bond Type

Electronegativity Difference Range	Most Probable Type of Bond	Example
0.0 – 0.3	Nonpolar covalent	H-H (0.0)
0.4 – 0.9	Moderately Polar Covalent	H-Cl (0.9)
1.0 – 1.9	Very Polar Covalent	H-F (1.9)
≥ 2.0	Ionic	Na^+Cl^- (2.1)

Labeling Bond Polarity

- Atoms obtain partial charges, much less than a ±1 charge like in an ionic bond.
- Partial charges are represented with the lowercase Greek letter ***delta, δ***. δ+ if positive, δ- if negative.
- For example, HCl:

$$\overset{\delta+}{\mathrm{H}}\text{——}\overset{\delta-}{\mathrm{Cl}}$$

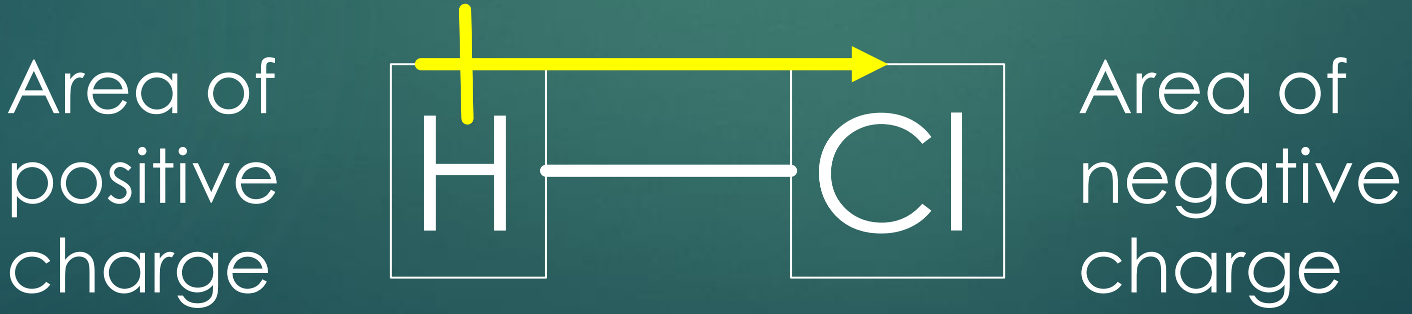
Labeling Bond Polarity

- Can also be shown using an arrow above the structural diagram:

$H-Cl$

Polar Molecules

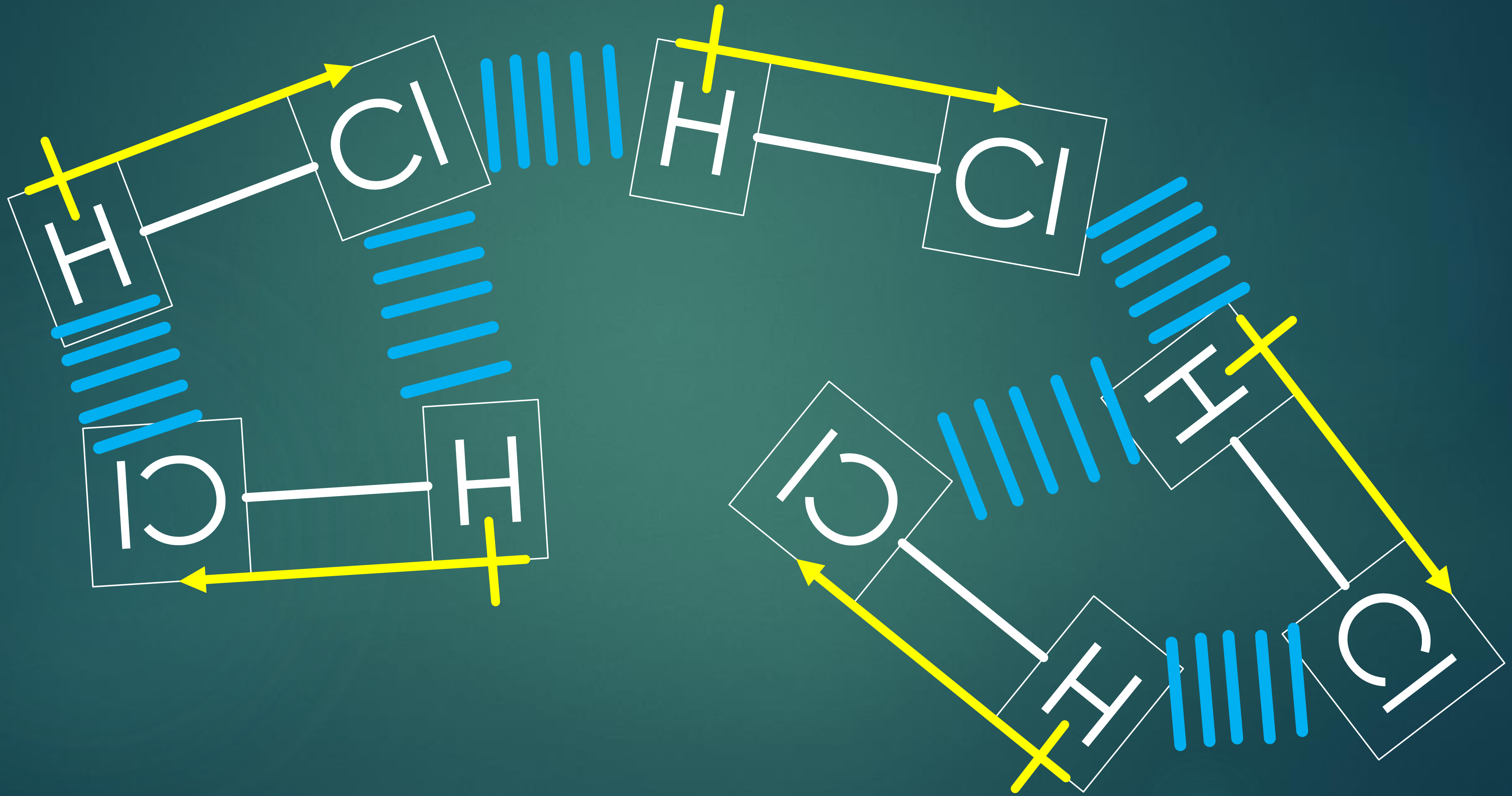
- Polar bonds may make the ***entire molecule polar***.
  - If areas of charge are created through the existence of polar bonds.
- A molecule that has two poles is called a ***dipole*** (like small magnets).
- For example, HCl is a polar molecule.



Area of positive charge

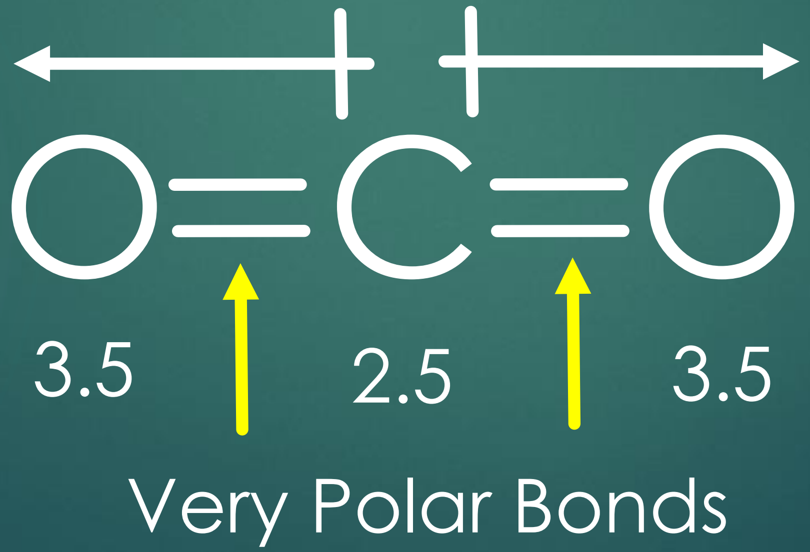
Area of negative charge

Polar Molecule Interactions: HCl



Molecular Polarity

- Polar bonds *don't necessarily* result in a polar molecule. The *shape* of the molecular plays a role in determining areas of charge difference.
- For example, CO_2



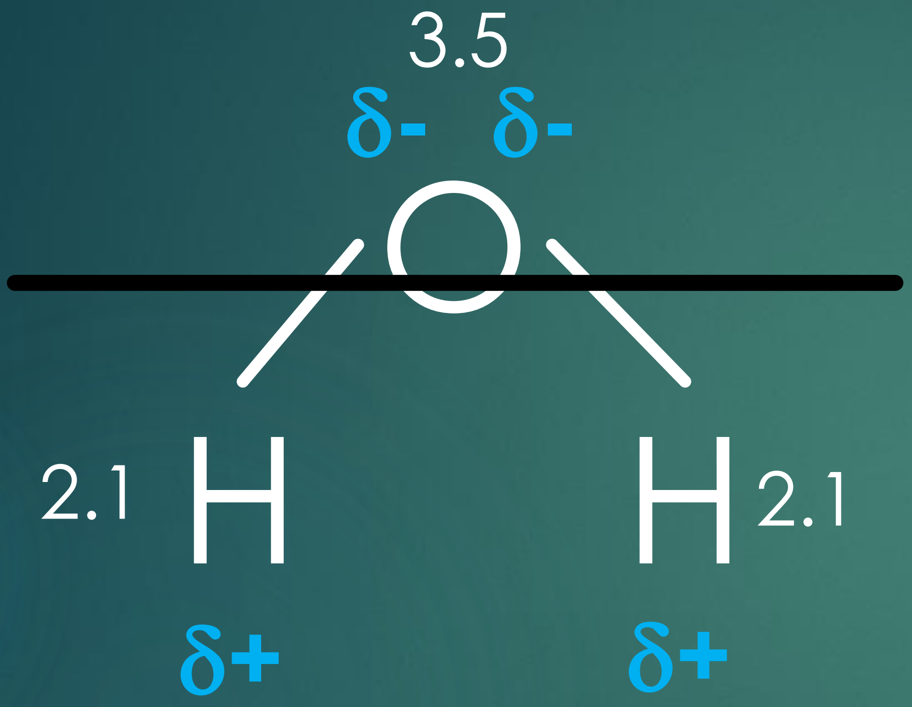
$O=C=O$

3.5 2.5 3.5

Very Polar Bonds

- No distinct positive and negative regions.
- Try to draw a line through the molecule so the areas of positive and negative are on either side.

Molecular Polarity: Water, H_2O

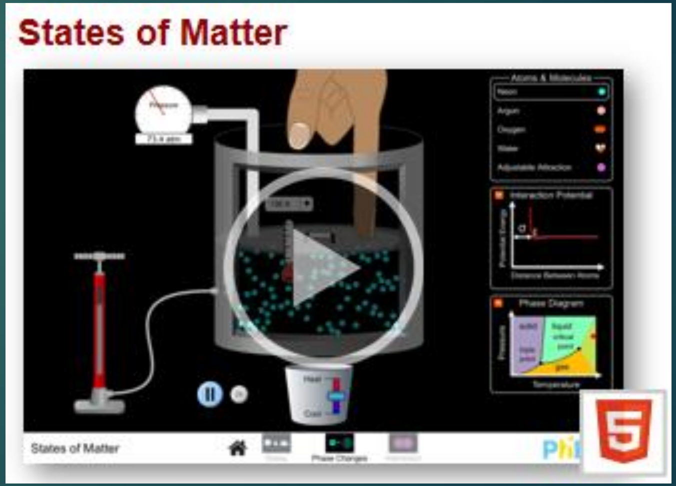


- Due to the bent shape, water is a polar molecule.

Attractions Between Molecules

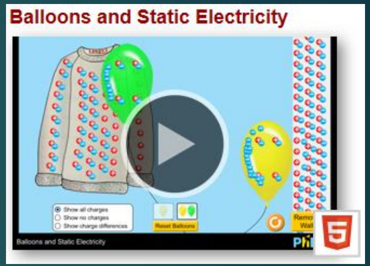
- Forces exist between molecules of a compound, called intermolecular forces.
- Weaker than both covalent and ionic bonds.
- Responsible for the state (solid, liquid or gas) of a molecule at a particular temperature.
- Three types of intermolecular forces:
  - Dispersion Forces
  - Dipole Interactions
  - Hydrogen Bonds

Called Van der Waals forces



Dispersion Forces

- Caused by the motion of electrons.
- Force between molecules increases as the number of electrons increase.

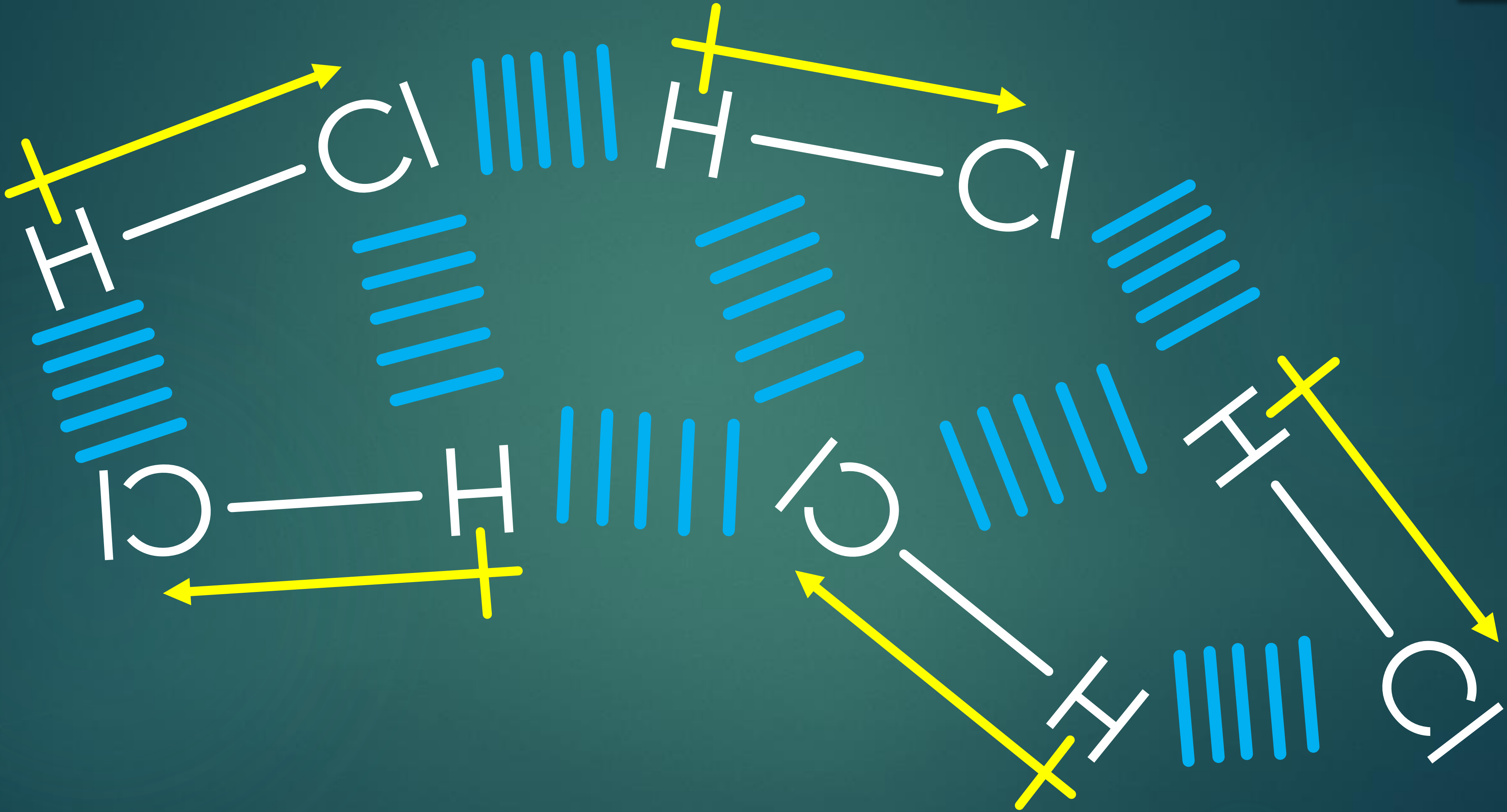


- Weakest of all intermolecular forces.
- Most prominent effect is that of the state of the halogens at room temperature and normal pressure:
  - Fluorine and chlorine are gases.
  - Bromine is a liquid.
  - Iodine and astatine is a solid.

Dipole Interactions

- Polar molecules are attracted to one-another.
- Positive area of one molecule attracts the negative area of another, and vice-versa.

Dipole Interactions



Hydrogen Bonding

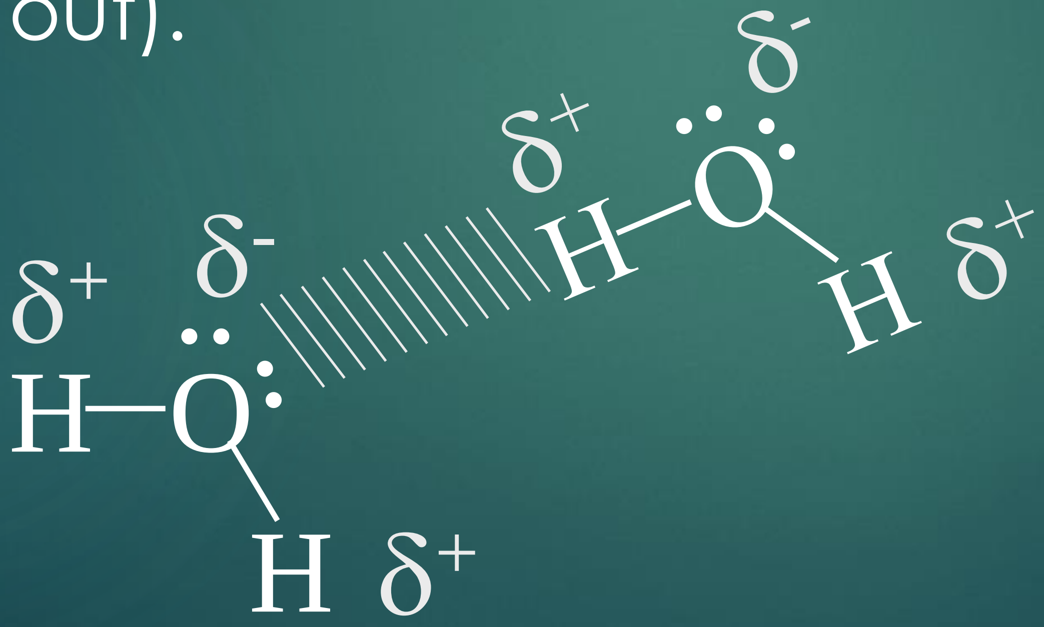
- Two conditions for hydrogen bonding to occur:

1. A hydrogen is covalently bonded to a very electronegative atom (creates a very polar bond).
    - With no other electrons to shield it, the positive nucleus is exposed.
2. The electronegative atom has at least one lone pair of electrons.

- The hydrogen bond is between the hydrogen in one molecule and the lone pair of electrons in another molecule.

Hydrogen Bonding

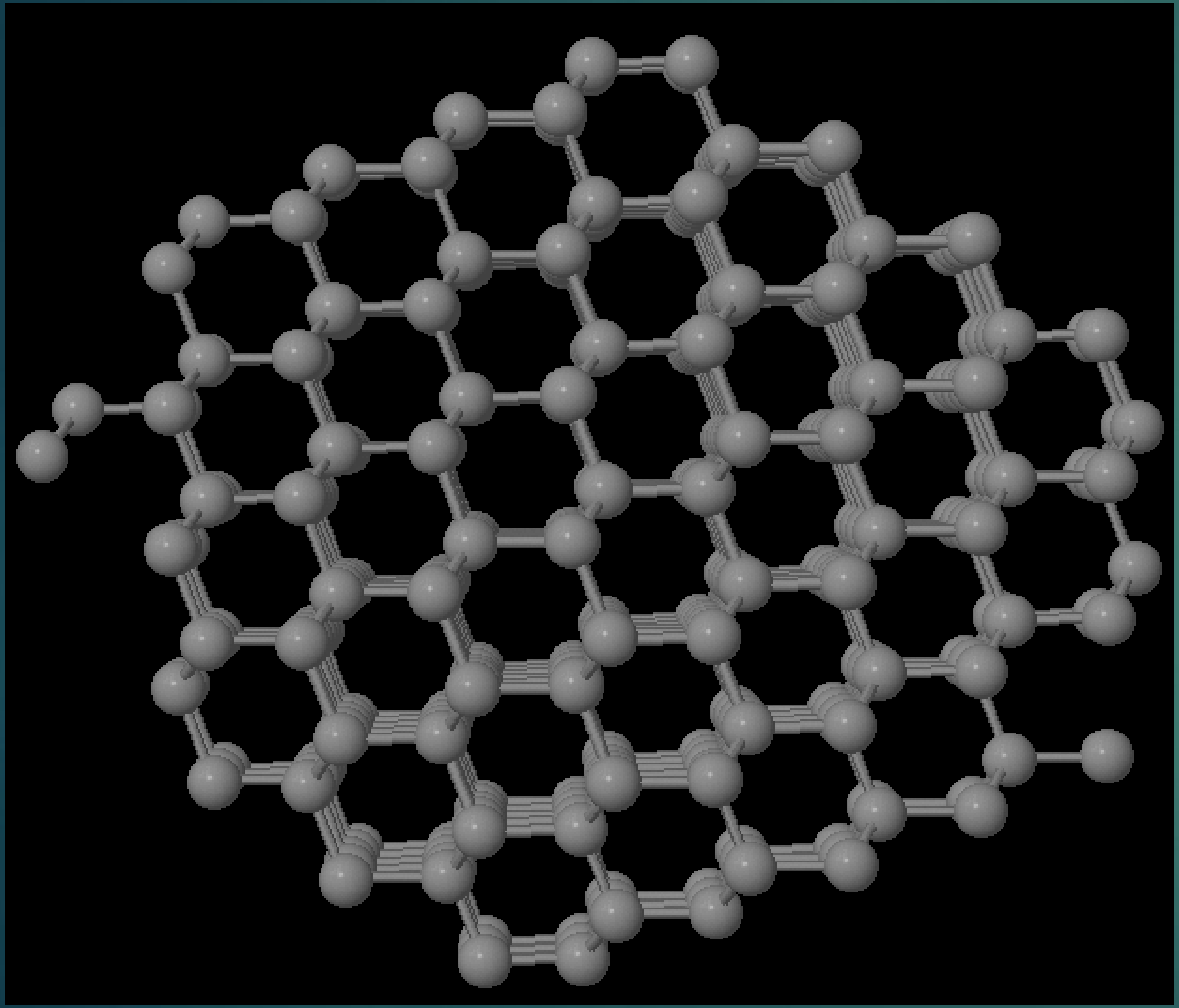
- Extremely important in determining the properties of water and biological molecules like proteins.
  - Relatively high melting and boiling points of water.
  - The surface tension of water (how it beads and does not spread out).

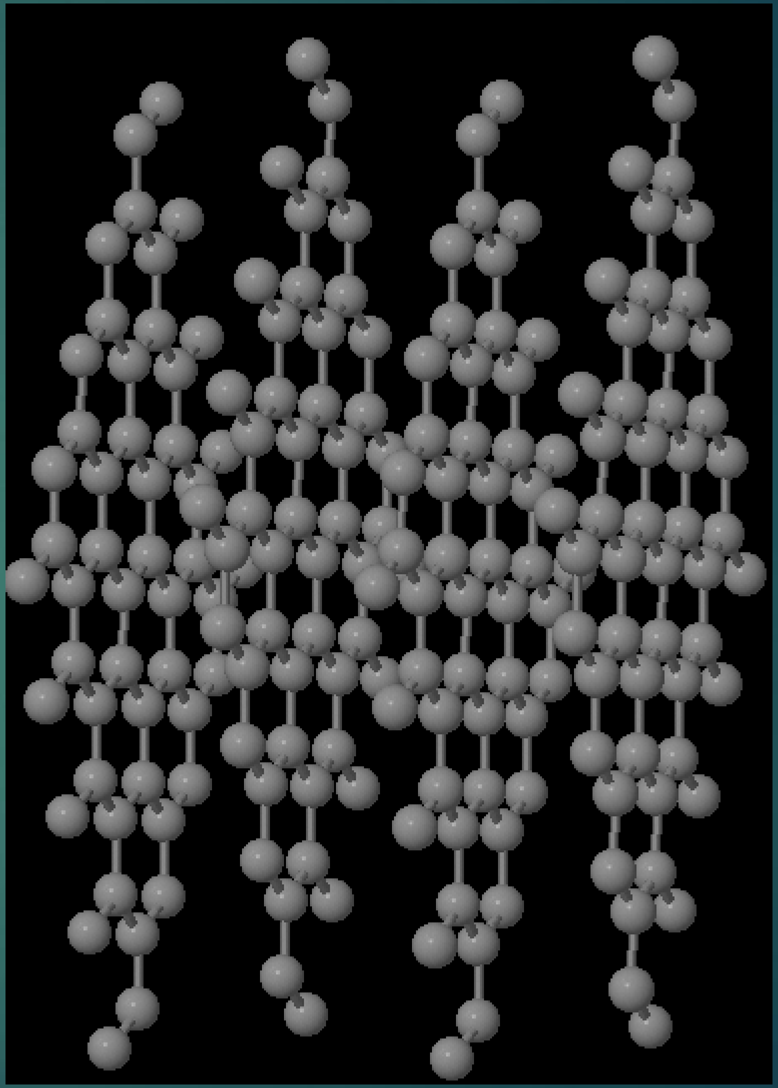


This hydrogen is bonded covalently to:

1) the highly negative oxygen, and
2) a nearby unshared pair.

Network Solid





Network Solid

- All the atoms are covalently bonded together.
  - There are no individual molecules.
  - Think of it as one big molecule.
- Very high melting points as all the covalent bonds throughout the solid must be broken (as opposed to only the intermolecular bonds of other compounds)
  - Diamond, graphite (carbon atoms)
  - Silicon Carbide

Summary of Ionic and Covalent Bonds

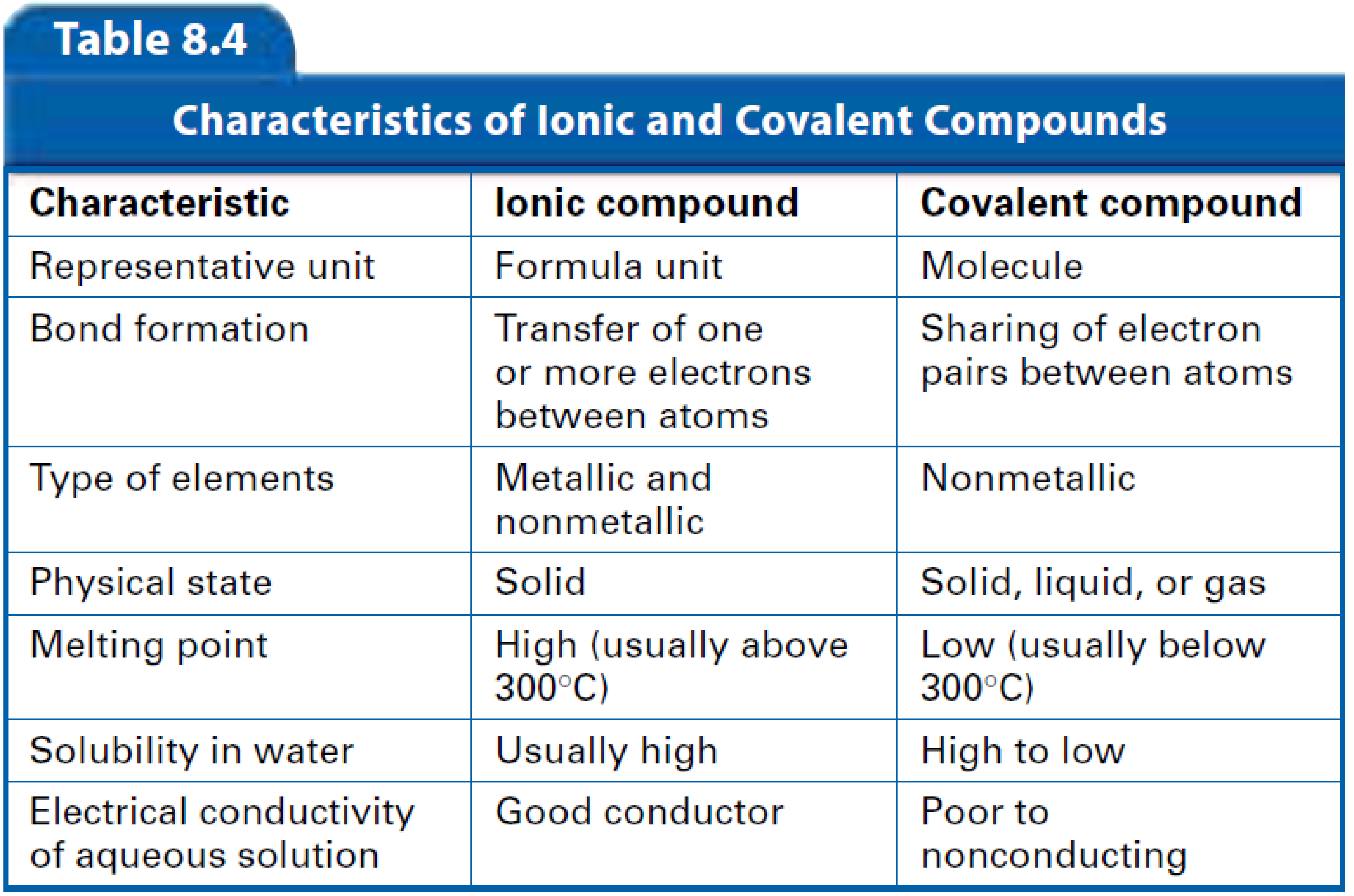


Table 8.4

Characteristics of Ionic and Covalent Compounds

Characteristic	Ionic compound	Covalent compound
Representative unit	Formula unit	Molecule
Bond formation	Transfer of one or more electrons between atoms	Sharing of electron pairs between atoms
Type of elements	Metallic and nonmetallic	Nonmetallic
Physical state	Solid	Solid, liquid, or gas
Melting point	High (usually above 300°C)	Low (usually below 300°C)
Solubility in water	Usually high	High to low
Electrical conductivity of aqueous solution	Good conductor	Poor to nonconducting

Learning Target Complete