Subatomic Particles

Octet Rule

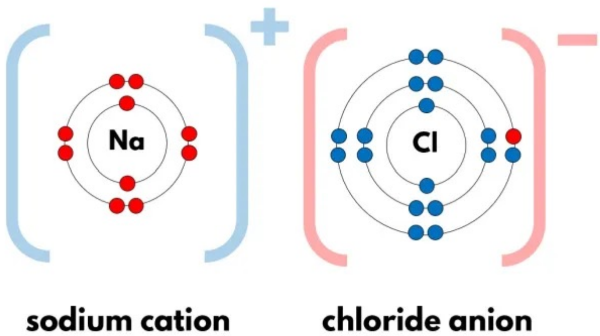
Atoms are the most stable when they have 0 or 8 valence electrons.

Ion – Atom that has lost or gained e^- to fulfil the octet rule

Sodium (Na) loses $1e^-$ to form a **cation**

1 val e^- → 0 val e^-

Cation = + Ion



Chlorine (Ca) gains $1e^-$ to form an **anion**

7 val e^- → 8 val e^-

Anion = - Ion

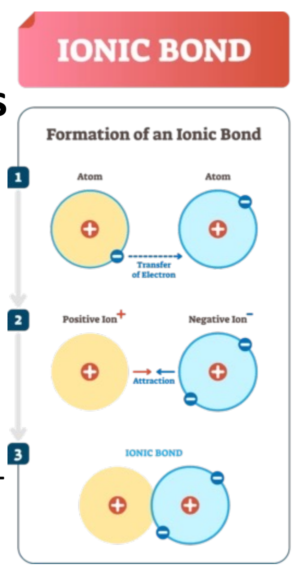
1

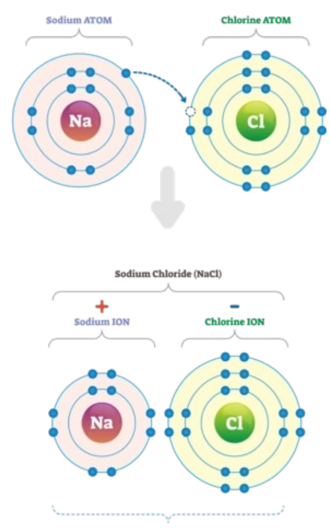
Ionic Bonding

From Ions to Ionic Bonds

Ionic Bonds are the connection between to atoms due to the transfer of electrons between a metal (*+ ion*) and a non-metal (*- ion*)

Ionization energy is the energy required to split apart two atoms into individual ions. The + and – ions give ionic bonds a very high ionization energy.



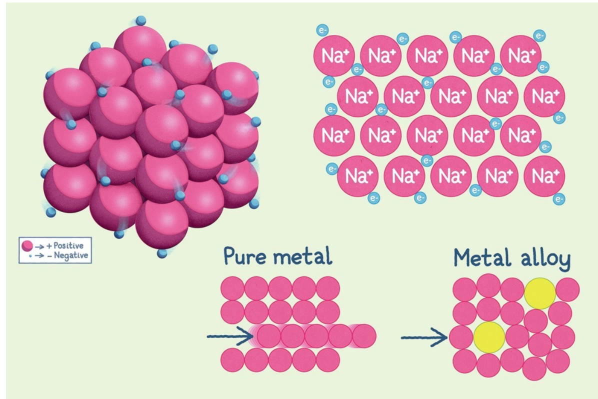


2

Metallic Bonding

Metallic Bonding is the process of positive ions (*cations, metals*) being held strongly together due to a group of *negative free electrons* (-) between atoms.

The free electrons form an **electrostatic force** (*strong connection between ions*) due to the positive ion (*+ metal ions*) and the negative electrons.



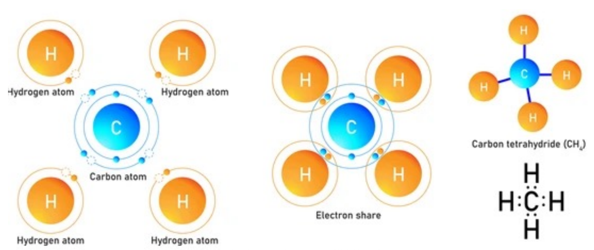
An **alloy** is a metal mixture where a different metal (*yellow in the diagram above*) that sits in the middle of other metal atoms.

3

Covalent Bonding

Covalent Bonding is the process of two non-metals **sharing electrons** to allow both atoms to obey the *octet rule* part of the time within the atomic structure

A atoms desire to obtain electrons is an atoms **electronegativity**. With non-metals high electronegativity requires atoms to share electrons to obey the *octet rule*



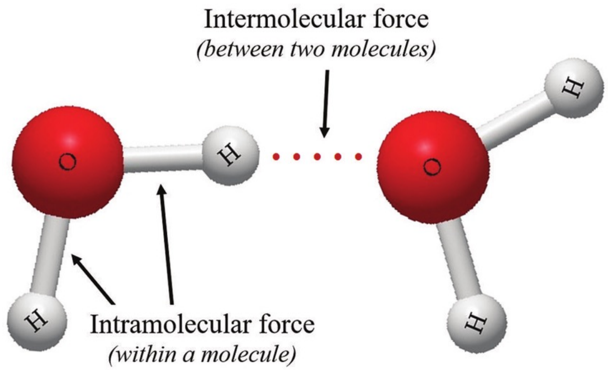
Covalent bonds can occur anytime there is a single electron available to share between electrons. In the example above, each hydrogen in CH_4 (*carbon tetrahydride*) is attached to carbon with a single covalent bond (*sharing of 2 electrons*)

4

Intramolecular and Intermolecular Forces

The attraction between atoms produces connections called forces. A force is a push or pull between objects due to a direct connection between the two objects.

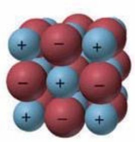
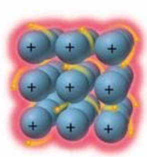
Intramolecular forces and forces that hold together the atoms in an ionic, covalent, or metallic bond. These bonds are very strong.



Intermolecular forces are forces between molecules. These bonds range from very weak to very strong.

5

Strength of Intramolecular Forces

Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Bonding				
Ionic		Cation–anion	400–4000	NaCl
Covalent		Nuclei–shared e^- pair	150–1100	H—H
Metallic		Cations–delocalized electrons	75–1000	Fe

Each intermolecular force has a specific strength. Ionic Bonds are the overall strongest, with Metallic Bonds being about 2 – 3 times weaker. Covalent Bonds are the weakest due to the sharing of electrons

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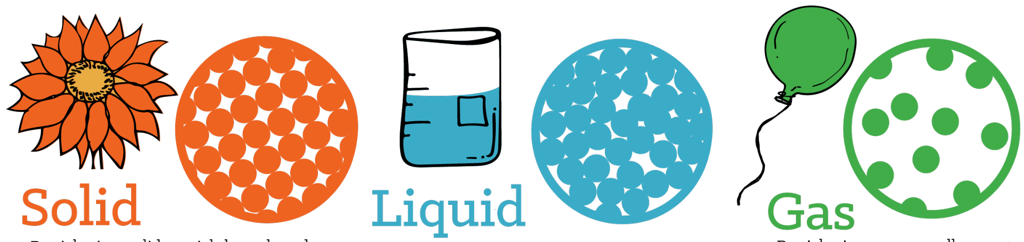
Three States of Matter



Solid
- Particles in a solid are tightly packaged and usually in a regular pattern.
- Particles in a solid will vibrate but cannot move past each other.
- Solids retain their shapes.

Liquid
- Particles in a liquid are close together with no regular pattern.
- Particles in a liquid flow and can easily move or slide past one another.
- Liquids assume the shape of their containers.

Gas
- Particles in a gas are well separated with no regular pattern.
- Particles in a gas vibrate freely at high speeds.
- Gasses assume the shapes of their containers.

A **state of matter** is the way that groups of atoms, compounds, and molecules connect together in space.

7

Intermolecular Forces & States of Matter

Solid

A **solid** is a state of matter (*the way matter exists in nature*) that has a fixed shape and volume.

In *solids* particles are held tightly together with strong intermolecular forces, normally between + and – ions from ionic compounds.

Fixed is a property that *can't change* over time. **Variable** is a property that *will change* over time in a system

8

Intermolecular Forces & States of Matter

A **liquid** is a state of matter (*the way matter exists in nature*) that has a **fixed** volume but a **variable** shape.

In *liquids* particles are held together enough intermolecular force to keep the particles together with each other but are weak enough to allow movement.

Volume is a measurement of the space a groups of particles occupy.

9

Intermolecular Forces & States of Matter

A **gas** is a state of matter (*the way matter exists in nature*) that has a **variable** shape and volume.

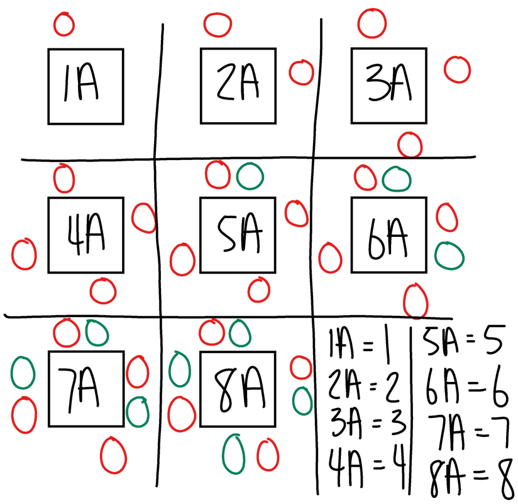
In *gases* particles are old held together with very weak intermolecular forces, only connecting together when they are close, or in contact with each other in space

---

Shape is the physical arrangement of particles in space.

10

Lewis Dot Structures Review



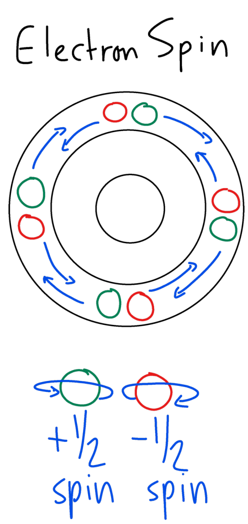
Valance Electrons are the outer *bonding* electrons.

Bonding is the connection of *two or more* atoms due to the *transfer* or *sharing* of valance electrons.

Lewis Dot Structures are visual representation of *valance electrons* using dots in a **electron filling order**

11

Pairs of Electrons



Electron Pairs
An ***electron pair*** is a set of electrons that exist in the same area (*orbital*) of the atom. An *orbital* is a location within an atom that electrons can exist (*labeled s, p, d, and f*)

Electron Spin
An ***electron spin*** is the *direction* that an individual electron travels within an atom. (*labeled as +1/2 and -1/2 spin*)

Both electrons in a pair have opposite spins

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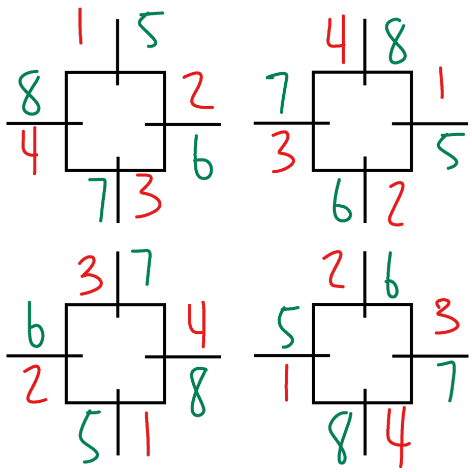
Electron Filling Order



Electron Filling Order is the method that *valence electrons* fill around the outside of the atom.

Electrons always fill in a **4 – 4 pattern** around the *Lewis Dot Structure.*

4 – 4 Pattern
One Electrons on each side of the structure, then pair in the same order

Electron filling can start on any side

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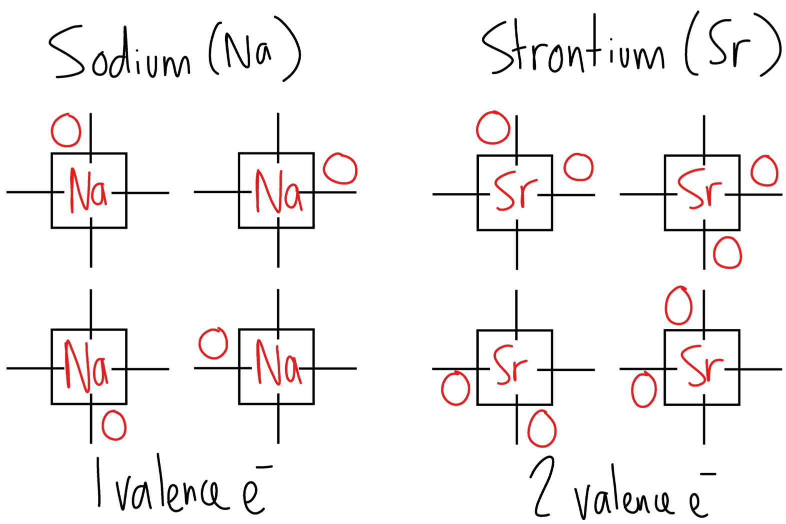
Metal Electron Filling Order



Representative Metals Atoms
(*groups 1A, 2A, Al*) only have a single valance electron on each side of the dot structure

No paired e⁻

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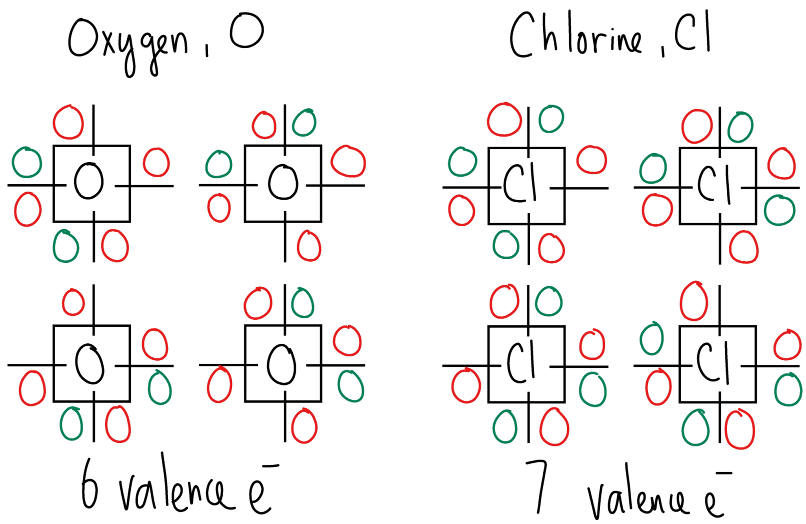
Non-Metal Electron Filling Order



Representative Non-Metal Atoms (*groups 5A – 8A*) will have a combination of paired & unpaired valance electrons

Paired e$^-$

Unpaired e$^-$

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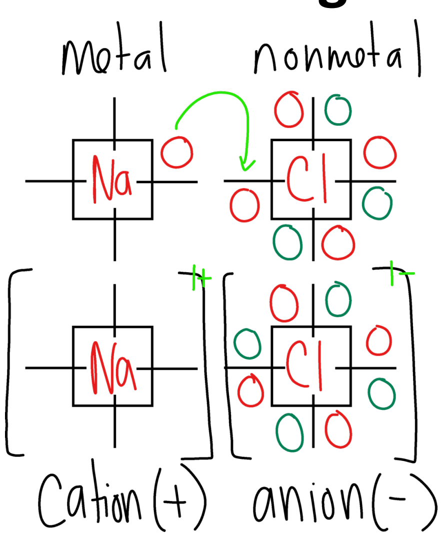
Ionic Bonding Review



Octet Rule

Atoms will *bond* with each other to obtain either 0 (*none*) or 8 (*full*) *valence electrons* around the atom.

Ionic Bonding – Transfer of Electrons

An **ionic bond** forms when electrons are transferred from one atom to another to meet the *octet rule*.

1 – 3 valence e$^-$ - Lose Electrons (*+ ion*)

5 – 7 valence e$^-$ - Gain Electrons (*- ion*)

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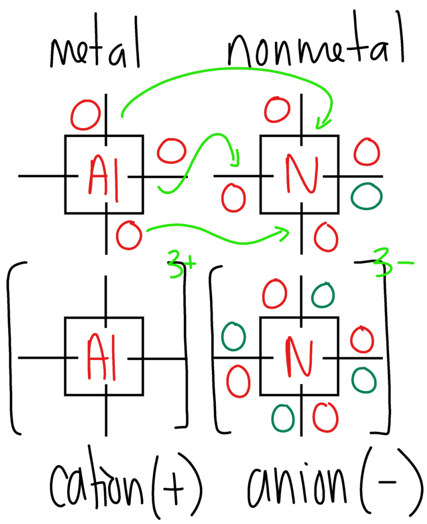
Ionic Bonding Review



Electron Transfer Procedure
Electrons transfer to an open space within an orbital, shown in the dot structures.

Energy and Electron Transfer
Multiple valance electron transfers (*metal to non-metal*) occur at the same time when the atoms *collide* and *transfer energy* between the atoms.

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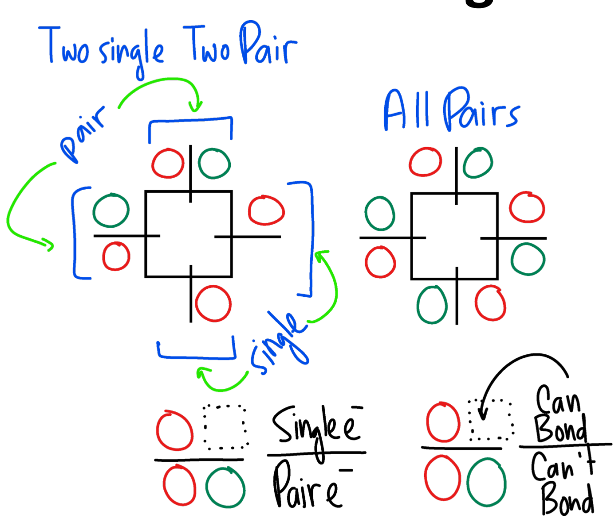
Covalent Bonding



Covalent Bonding
A **covalent bond** forms when electrons are *shared* between two atoms to satisfy the *octet rule*.

Sharing and Pairs of Electrons
If there is only a single electron on one side of a *lewis structure* bonding can occur between atoms

If the electrons are *paired* (*two electrons*) no bonds can occur

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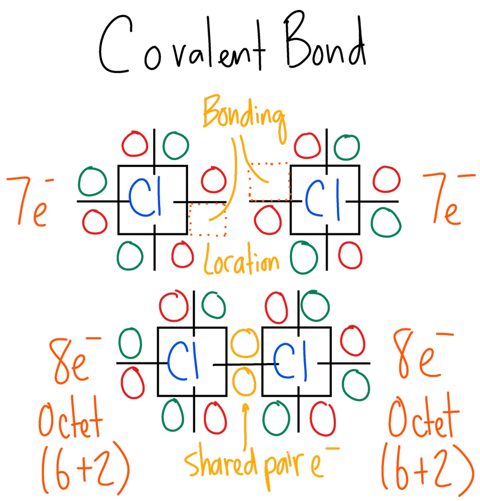
Single Covalent Bonding



Single Covalent Bond

A *single covalent bond* is a connection (*bond*) between two non-metals due to the *sharing of electrons* between the atoms.

Each atom in the bond gets the extra electron 50% (*half*) of the time

Atom 1: 6 + 2 (50%) = 8 (Octet)
Atom 2: 6 + 2 (50%) = 8 (Octet)

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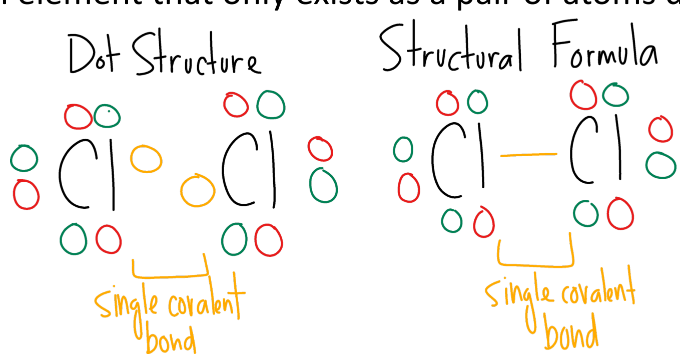
Diatomic Elements

Diatomic Element

An element that only exists as a pair of atoms due to covalent bonding



Diatomic Elements are more stable bonded together to give each a share of 8 valance e^-

H_2	**Hydrogen**
N_2	**Nitrogen**
F_2	**Fluorine**
O_2	**Oxygen**
I_2	**Iodine**
Cl_2	**Chlorine**
Br_2	**Bromine**

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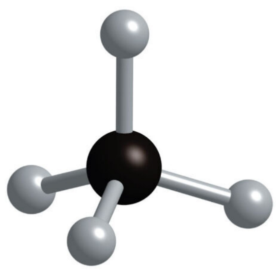
Central Atoms and Covalent Bonding

Central Atoms

Central Atoms are the atoms that sit in the middle (*center*) of the atomic structure, and have other atoms covalently bonded to its structure



Carbon (*black sphere*) is the central atom, surrounded by 4 Hydrogen Atoms in *Methane* (CH_4)

The number of bonds for each central atom is based on the group of each atom

Central Atom Group Number	Number of Single Covalent Bonds (*Steric Number*)
Group 4A (14)	4
Group 5A (15)	3
Group 6A (16)	2
Group 7A (17)	1

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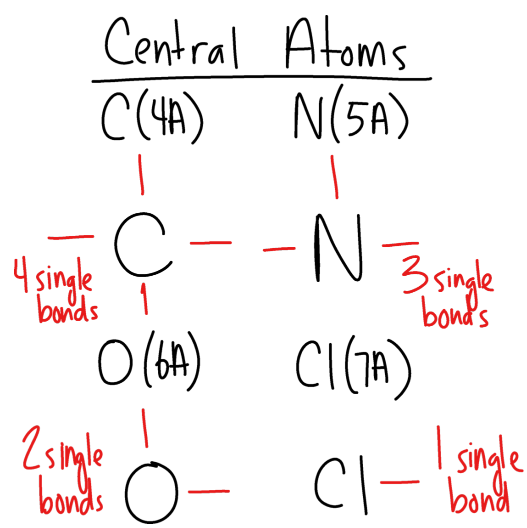
Central Atoms and Covalent Bonding

Central Atom Electron Pairs

Central Atoms that produce 1, 2, or 3 single bonds also contain pairs of electrons around atom

Central Atom Group #	Number of Electron Pairs	Number of Single Bonds
Group 4A	0	4
Group 5A	1	3
Group 6A	2	2
Group 7A	3	1



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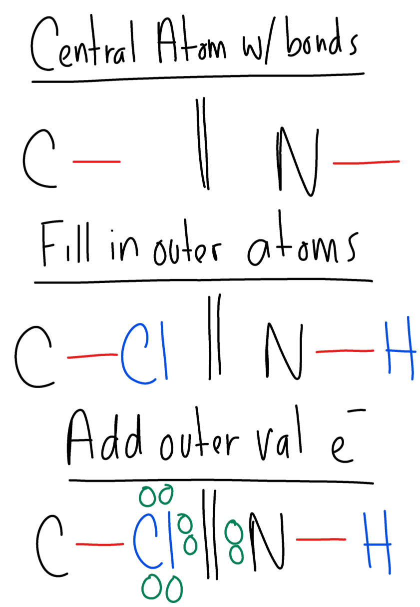
Drawing Structures

Process of Drawing Structures

1. Write the central atom in the center of the structure (*given*)
2. Write atoms and add a single bond between the central atom and each outer atom (*bonds given*)
3. Fill in the outer valence electrons.
   1. Write a pair of electrons from each location not bonded
   2. Count the total number of electrons in the structure



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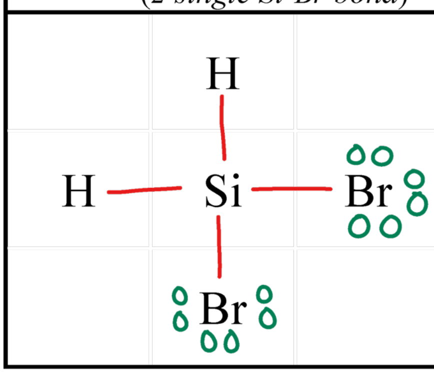
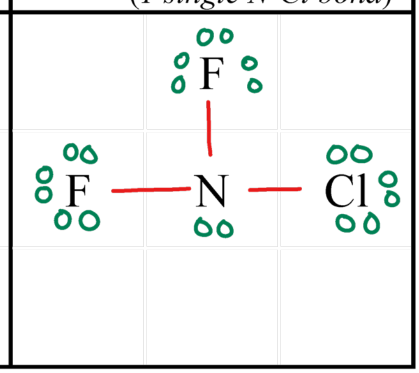
Covalent Molecule Examples

H_2SiBr_2 (*2 single Si-H bonds*) (*2 single Si-Br bond*)	NF_2Cl (*2 single N-F bonds*) (*1 single N-Cl bond*)
	

H: 1 val e⁻ Br: 7 val e⁻
Si: 4 single

F: 7 val e⁻ Cl: 7 val e⁻
N: 3 single + 1 pair

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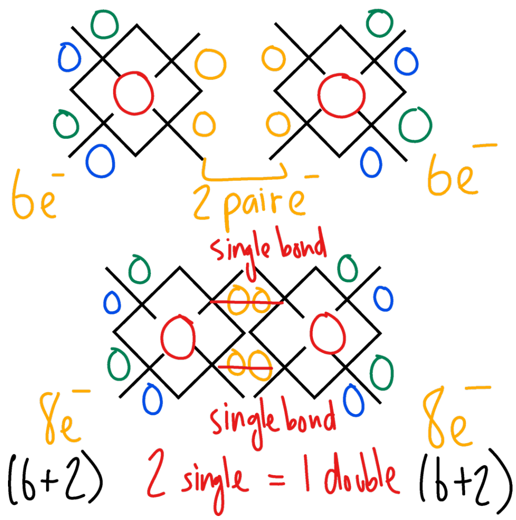
Double Covalent Bonding



Double Covalent Bond
A *double covalent bond* is a connection (*bond*) between two non-metals due to the *sharing of two pairs of electrons* between the two atoms.

Each atom in a double covalent bond shares 2 pairs of electrons equally.

Molecules can contain a combination of single and double covalent bonds

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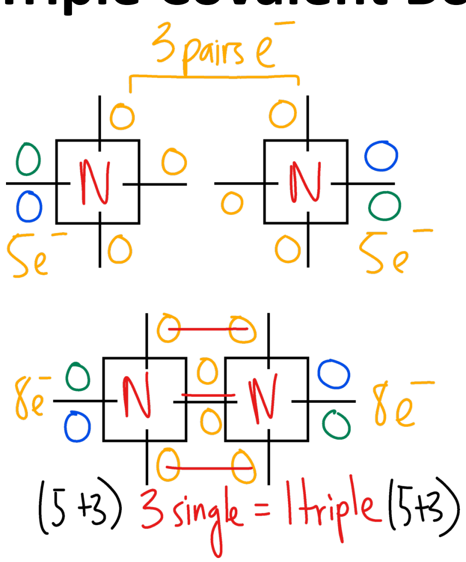
Triple Covalent Bonding



Triple Covalent Bond
A *triple covalent bond* is a connection (*bond*) between two non-metals due to the *sharing of three pairs of electrons* between the two atoms.

Each atom in a double covalent bond shares 2 pairs of electrons equally.

Molecules can contain a combination of single and double covalent bonds

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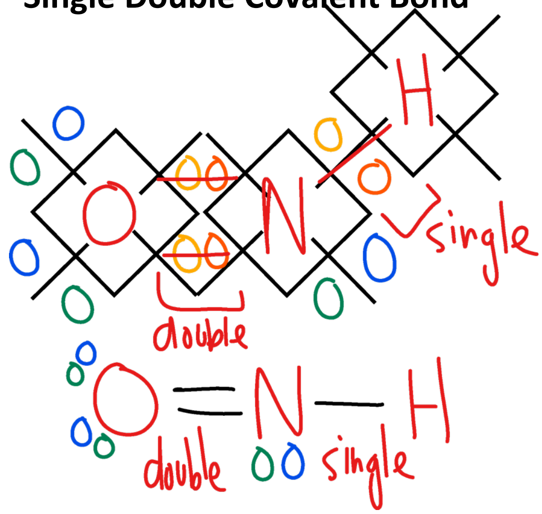
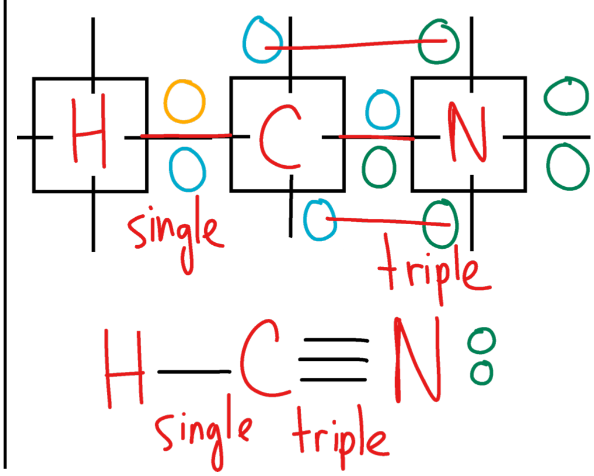
Combining Bond Types

Single Double Covalent Bond



Single Triple Covalent Bond



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Determining Covalent Bonds

Steps to Determine Type of Covalent Bonds

Covalent bonds are determined with the following steps:

1. Draw a single line between atoms
2. Add unpaired electrons to all atoms
3. Count electrons around all atoms
   - Bonds act as two electrons
4. Add additional lines when an atom does not have enough electrons
   - 6 electrons: 1 extra line
   - 4 electrons: 2 extra lines

Element Group Number of bonds and pairs

Central Atom Group #	Number of Electron Pairs	Number of Single Bonds
Group 1A	0	1
Group 4A	0	4
Group 5A	1	3
Group 6A	2	2
Group 7A	3	1
Group 8A	4	0

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Covalent Nomenclature

Every Covalent Molecule is named one of two ways:

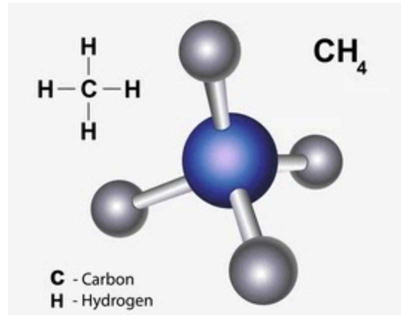
1. Prefixes: Binary Covalent
2. Organic: Larger Molecules (*by type*)

Covalent Molecules can come in many formula variations based on structure

Formulas for Example Carbon Molecules

CH_4	C_2H_6	C_2H_4	C_2H_2
C_3H_8	C_3H_6	C_3H_4	C_4H_{10}
C_4H_8	C_4H_6	C_4H_4	C_5H_{12}

Some molecules have multiple names



Binary: Carbon Tetrahydride
Organic: Methane

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Binary Covalent Nomenclature

Covalent Molecules are named based on the *prefix* model. The *prefix* is a number before each atom in a binary covalent molecule

Prefix (*no mono-*) First Element
Prefix Second Element (-ide)

Binary Molecule Examples

CH_4: Carbon Tetrahydride
S_2O_2: Disulfur Dioxide
NCl_3: Nitrogen Trichloride

Covalent Molecule Prefixes

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

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Binary Covalent Nomenclature

CH_4 — 1 C and 4 H

___ carbon tetra hydride
(no mono)

S_2O_3 — 2 S and 3 O

di sulfur tri oxide

1 : mono
2 : di
3 : tri
4 : tetra
5 : penta
6 : hexa
7 : septa
8 : octa

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Binary Covalent Nomenclature

tri (3) carbon (C) hexa (6) chloride (Cl)

C _ Cl _ → C_3Cl_6

___ (1) nitrogen (N) di (2) oxide (O)

N _ O _ → N_xO_2 → NO_2

1 : mono
2 : di
3 : tri
4 : tetra
5 : penta
6 : hexa
7 : septa
8 : octa

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Electronegativity

An atoms ability to attract electrons towards itself in a chemical bond

Rep and Trans Metals

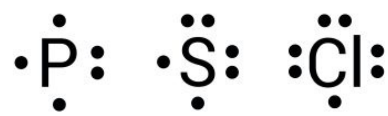
1 – 4 Valence Electrons (e^-)

- Wants to Lose Electrons
- Forms Ions vs Sharing e^-
- Low Electronegativity (*0.7-1.9*)

Metalloids and Non-Metals

4-7 Valence Electrons (e^-)

- Wants to Gain Electrons
- Shares Electrons (*Covalent*)
- High Electronegativity (*2.0 – 4.0*)



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Electronegativity Difference

The difference in electronegativity values between atoms in a bond

Ionic Bonds

Transfer of Electrons

(*Metal to Non-Metal*)

Metal – Low Electroneg.

Non-Metal – High Electroneg.

Electroneg Difference Range

1.91 – 3.50 with a Metal

Polar Covalent Bonds

Unequal Sharing of Electrons

(*Non-Metal to Non-Metal*)

Non-Metal – Mid Electroneg.

Non-Metal – High Electroneg.

Electroneg Difference Range

0.51 – 1.90 (No Metals)

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Electronegativity Difference

The difference in electronegativity values between atoms in a bond

Non-Polar Covalent Bonds
Equal Sharing of electrons
(*Non-Metal to Non-Metal*)

Non-Metal – High Electroneg.
Non-Metal – High Electroneg.

Electroneg Difference Range
0.00 – 0.50 (No Metals)

Diatomic Element
Equal Sharing of electrons
(*Same Elements*)

Element A – Electroneg. A
Element A – Electroneg. A

Electroneg Difference Range
0.00 (Same Element)

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Bonding Comparison

Ionic
Transfer
A → B (e^-)
ED > 1.9
and/or w/ metal

Polar
Unequal share
A ⇄ B (e^-)
ED = 1.9 – 0.51
no metal

NonPolar
Equal Share
A ⇄ B (e^-)
ED = 0 – 0.5

Diatomic
Same Atoms
A ⇄ A (e^-)
ED = 0.0

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Calculating Electronegativity Difference

Electroneg. Difference (ED) = High Electro. Value – Low Electro. Value

Ionic Bond

Na (*E = 0.9*) and Cl (*E = 3.0*)

ED = Electro. Cl – Electro. Na
ED = 3.0 – 0.9 = 2.1 (*Ionic*)

Polar Covalent Bond

H (*E = 2.1*) and Cl (*E = 3.0*)

ED = Electro. Cl – Electro. H
ED = 3.0 – 1.9 = 0.9 (*Polar*)

Non-Polar Covalent Bond

Cl (*E = 3.0*) and S (*E = 2.5*)

ED = Electro. Cl – Electro. S
ED = 3.0 – 2.5 = 0.5 (*Non-Polar*)

Diatomic Element Bond

Cl (*E = 3.0*) and Cl (*E = 3.0*)

ED = Electro. Cl – Electro. Cl
ED = 3.0 – 3.0 = 0.0 (*Non-Polar*)

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Electronegativity Examples

Al and F
Al: 1.5 F: 4.0
ED = 4.0 – 1.5 = 2.5
Ionic (>1.9 w/metal)
Al is a metal: Ionic Bond

H and Cl
H: 2.1 Cl: 3.0
ED = 3.0 – 2.1
= 0.9
Polar Covalent (0.51–1.9) no metal
Polar Covalent

Ionic
> 1.9 w/metal
Polar
0.51 – 1.9 no metal
Non-Polar
0.1 – 0.5
Diatomic
0.0

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Electronegativity Examples

Pb and P

Pb: 1.8 P: 2.2

ED = 2.2 − 1.8 = 0.4

Ionic: metal (>1.9)

Covalent: no metal (1.9 − 0.51)

Pb is a metal = Ionic

Steps

1. metal? Ionic (IA-2A, Al, IB-8B, In, Sn, Tl, Pb, Bi)
2. Calculate ED
3. Use chart to determ. bond type

Ionic
> 1.9 w/metal
Polar
0.51 − 1.9 no metal
Non-Polar
0.1 − 0.5
Diatomic
0.0

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Bond Dipoles

Dipole is an uneven sharing of electrons in bond

Polar = Dipole

Dipole → H — Ö: (dipole arrow ┼——→ from H to O)

H: 2.1, +end; O: 3.5, − end

ED = 1.4 Polar

Dipole: ┼——→

+ end	− end
Lower Electro.	Higher Electro.
gets e⁻ less	gets e⁻ more
attracts − end	attracts + end

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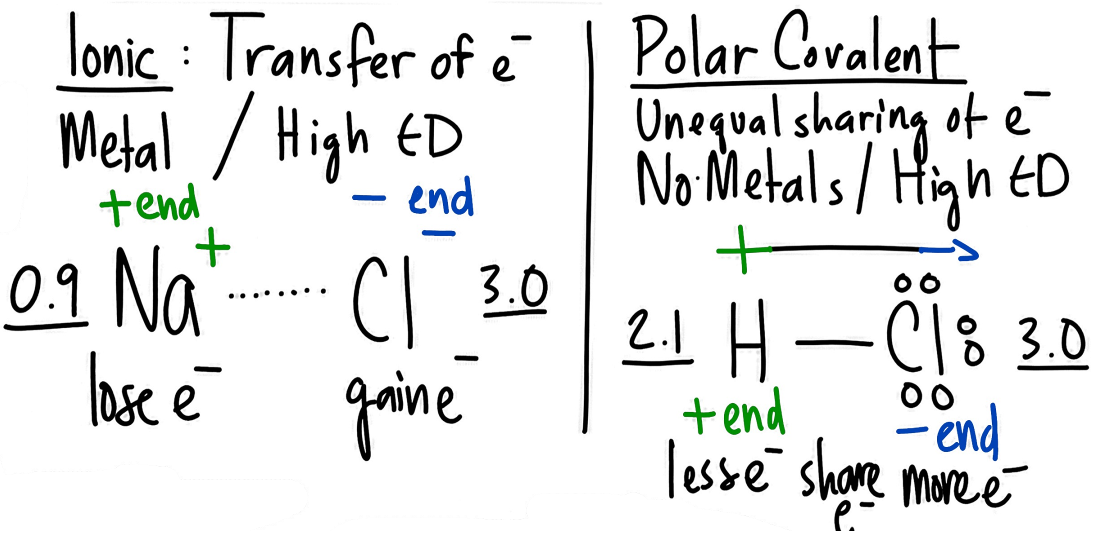
Ionic vs Covalent Bonds



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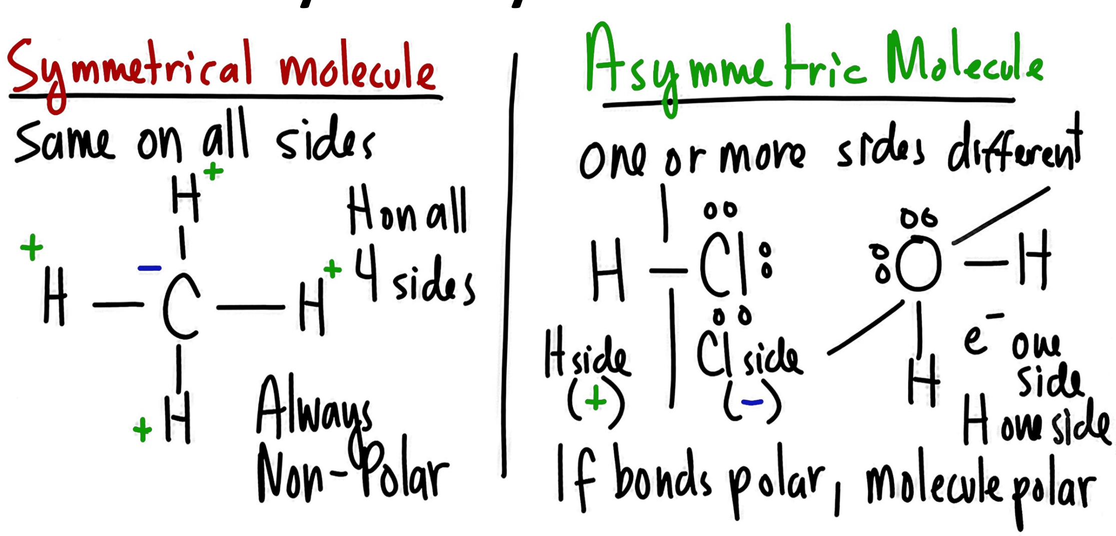
Molecule Symmetry



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Determining Polarity of Molecules

Molecule Polarity

1. Find bond ΔEN
2. NonPolar Bond = NonPolar Molecule
3. For Polar Bond
   - symmetrical molecule : NonPolar Molecule
   - asymmetrical molecule : Polar Molecule

2.1 H – Cl 3.0

ΔEN = 3.0 – 2.1 = 0.9

Bond is Polar Covalent Bond

H ⟶ Cl

Asymmetrical

Polar Molecule

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States of Matter

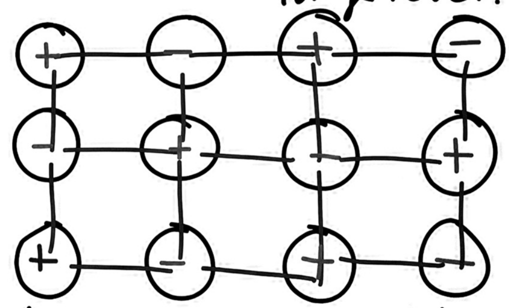
State: How particles interact with each other at a macro large level.

Solid

Fixed Shape

Fixed Volume

Strong Interactions



Ionic Compound as a solid

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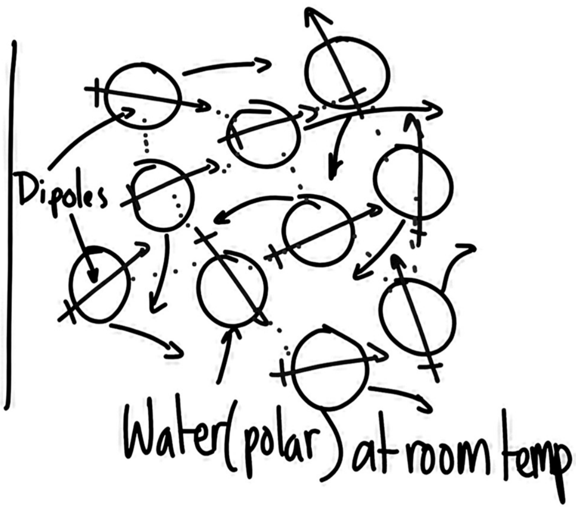
States of Matter

Liquid

Variable Shape

Fixed Volume

Medium Interactions



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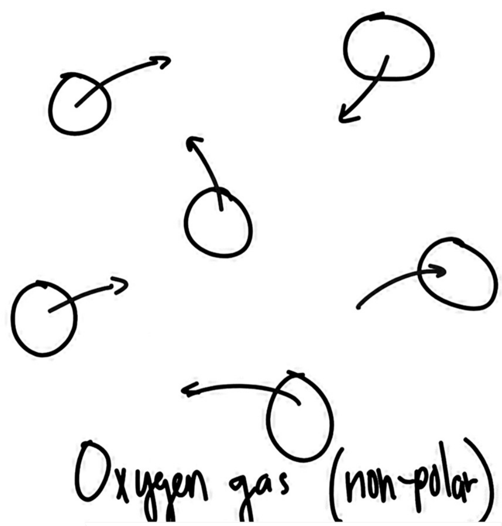
States of Matter

Gas

Variable Shape

Variable Volume

Weak Interactions



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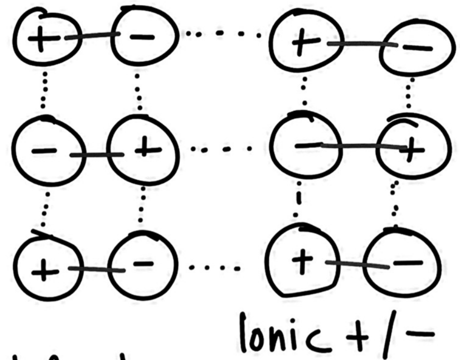
Molecular Interactions

Ionic – Ionic (Bond)

Type +/– crystal structure

State Solid

Strength Strong



Particle Connection

+/– charge attraction

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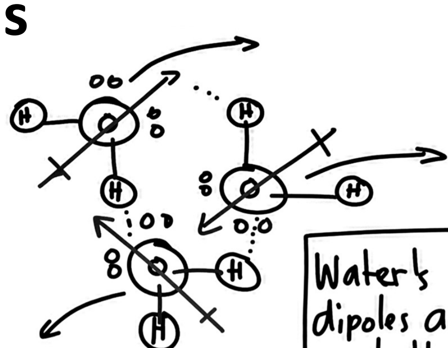
Molecular Interactions

Polar – Polar (Bond)

Type Dipole-Dipole

State Liquid

Strength Medium



Water's dipoles are constantly interacting

Particle Connection

Positive dipole end w/ negative dipole end (and – to +)

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Molecular Interactions

Non Polar - Non Polar (Bond)

Type Induced Dipole / Induced Dipole

State gas

Strength Weak

Induced dipole: temporary (short) interaction due to movement of val. e^-

Particle Connection

Quick connection when particles collide

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