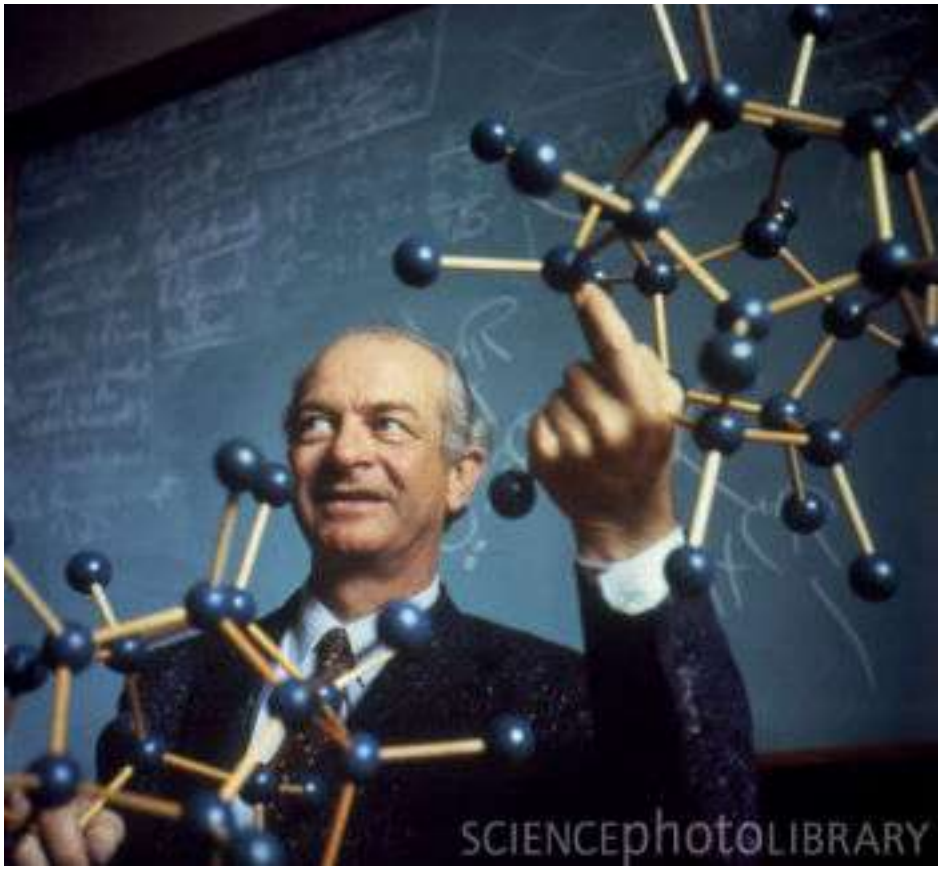
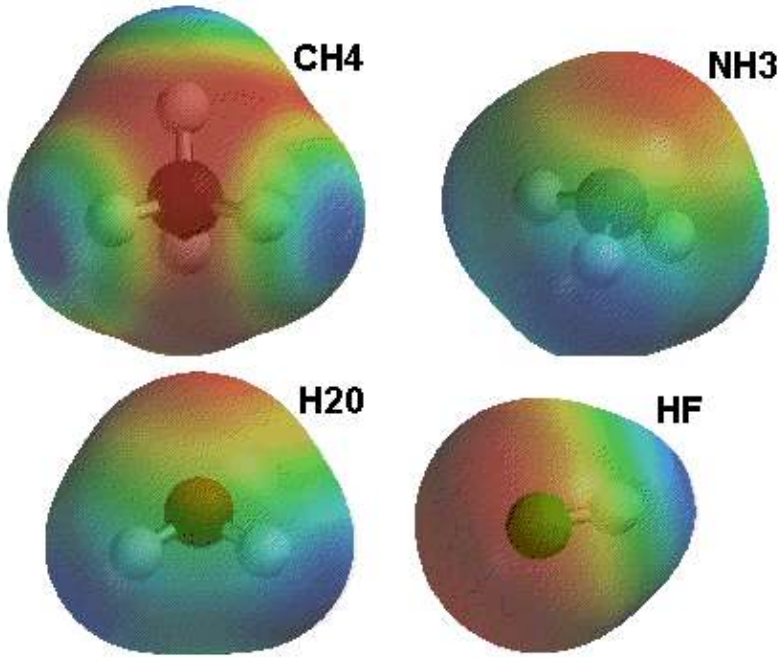
Pre-AP Chemistry

September 4, 2011

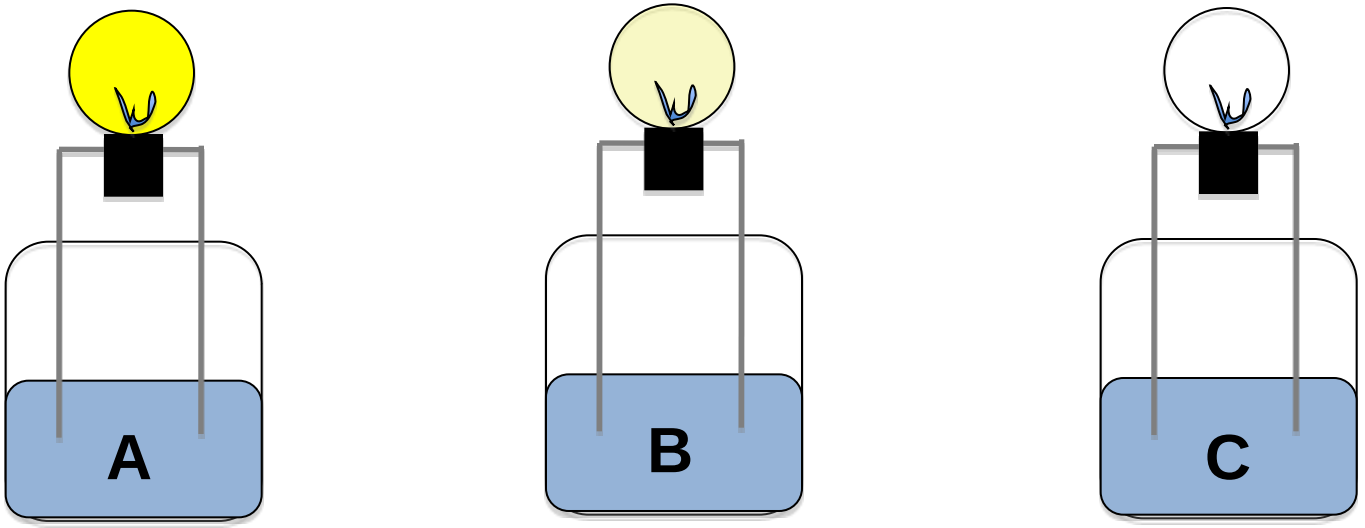
14. Covalent Bonding I





Review

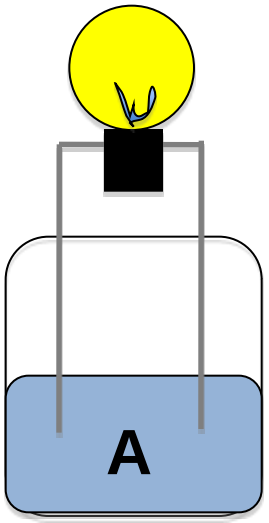
A researcher places electrodes in solutions A, B, and C, and finds the following results:



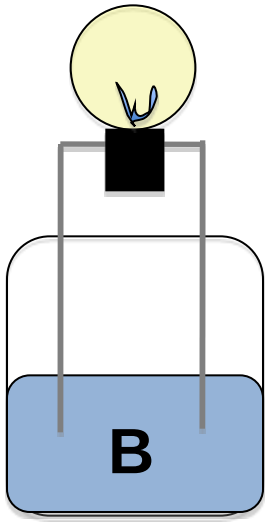
Which solution(s) are electrolytes? Which electrolyte is stronger?

Review

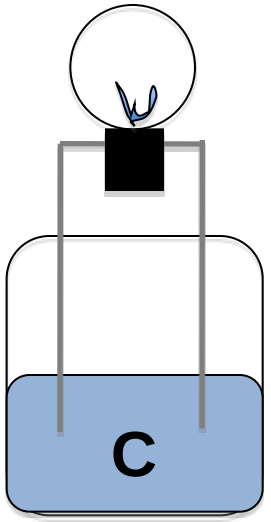
A researcher places electrodes in solutions A, B, and C, and finds the following results:



Stronger electrolyte



Electrolyte



Non-Electrolyte

Most compounds are NOT electrolytes.

Today, we'll learn about one type of bond in non-electrolytes.

Outline

- Covalent Bonds
- Molecular Lewis Structures
- Electronegativity

- Covalent Bonds
  - Ionic vs Covalent
  - Origin of Covalent Bonds
- Molecular Lewis Structures
  - Representing Covalent Bonds
  - Bond Order
- Electronegativity
  - Homonuclear vs Heteronuclear Bonds
  - Periodic Trends
  - Bond Polarity

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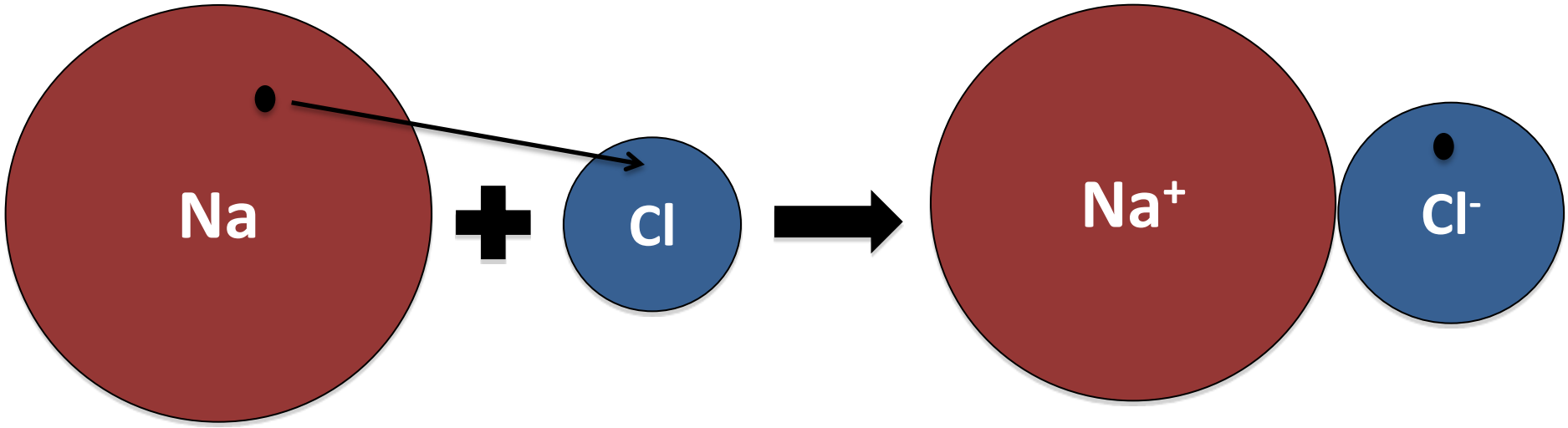
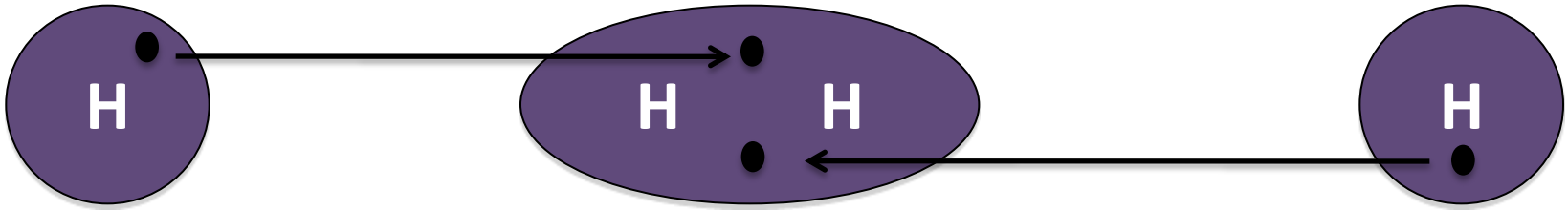
Covalent Bonds

- In **ionic bonds**, electrons are **transferred**
- In **covalent bonds**, electrons are **shared**



Ionic Bonds = Electron Transfer

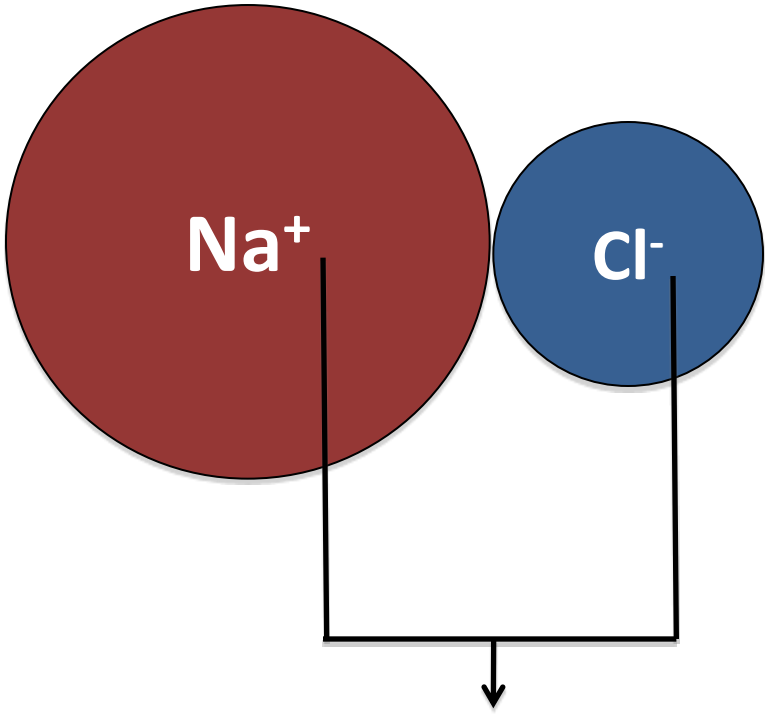
Covalent Bonds = Electron Sharing

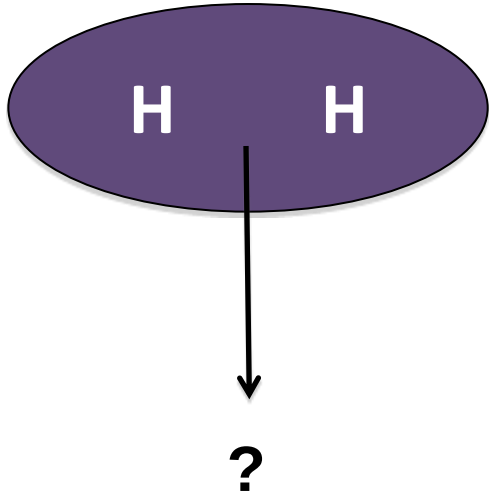


Covalent Bonds

- **Ionic bonds** form because oppositely charged anions and cations attract

- Why does a **covalent bond** form between two neutral atoms?





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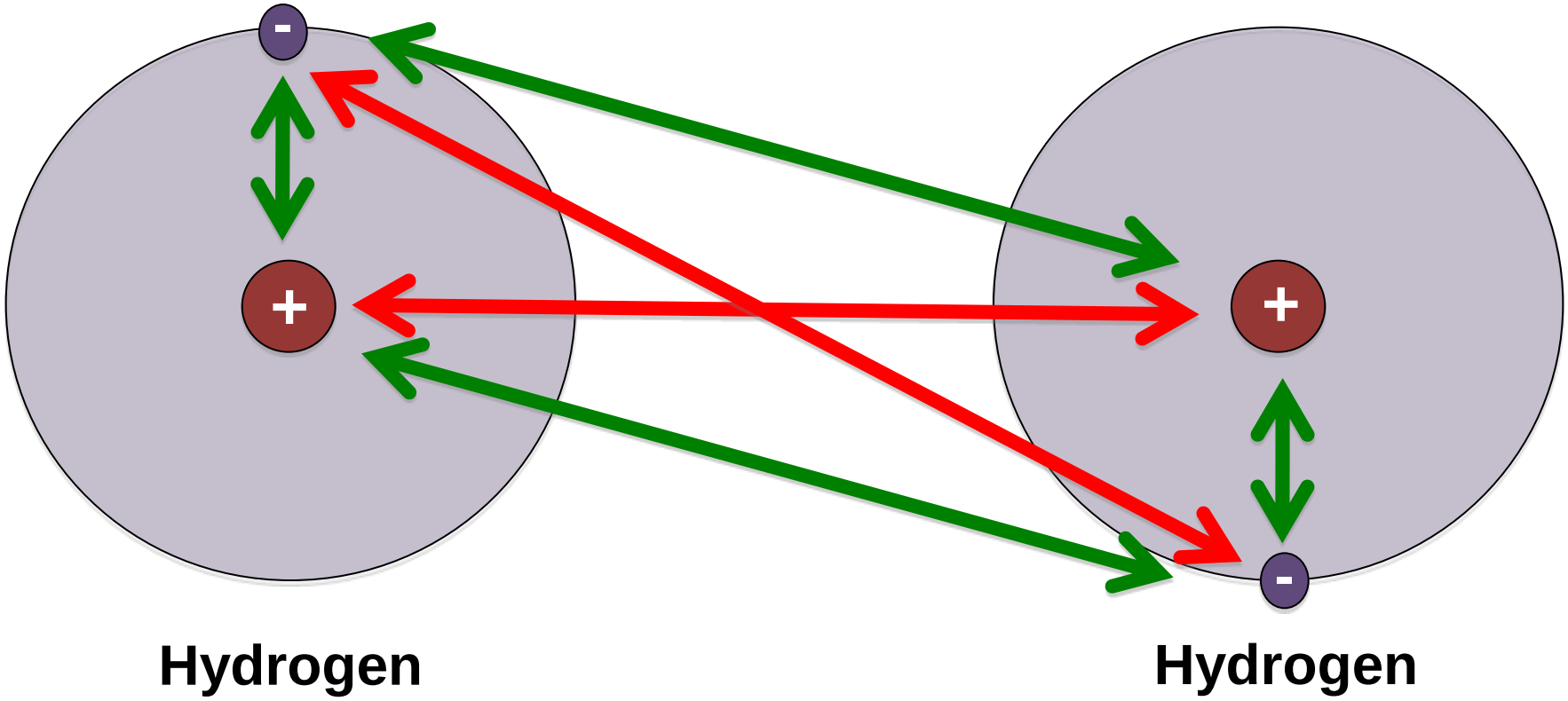
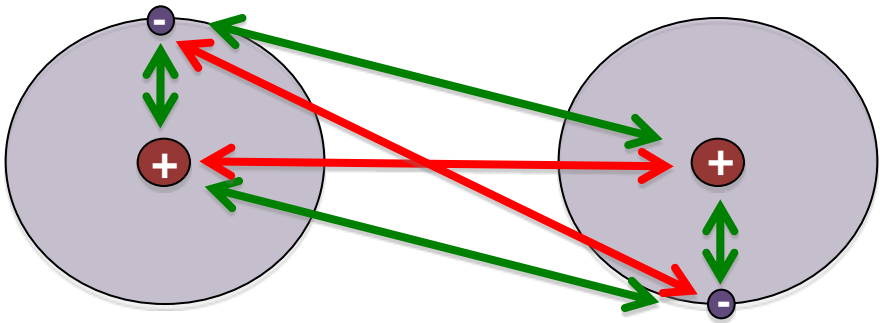
Covalent Bonds

- Consider two hydrogen atoms
  - Electrons repel each other
  - Nuclei repel each other
  - Electrons and nuclei attract each other

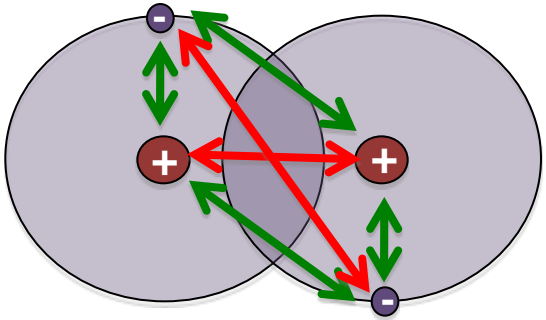


Covalent Bonds

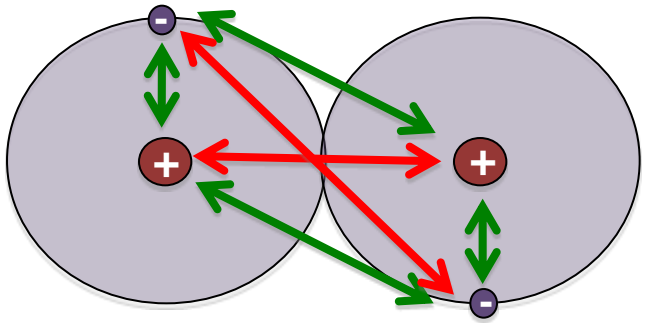
- At bond distance, repulsive forces = attractive forces



Too Far Apart = Too Attractive



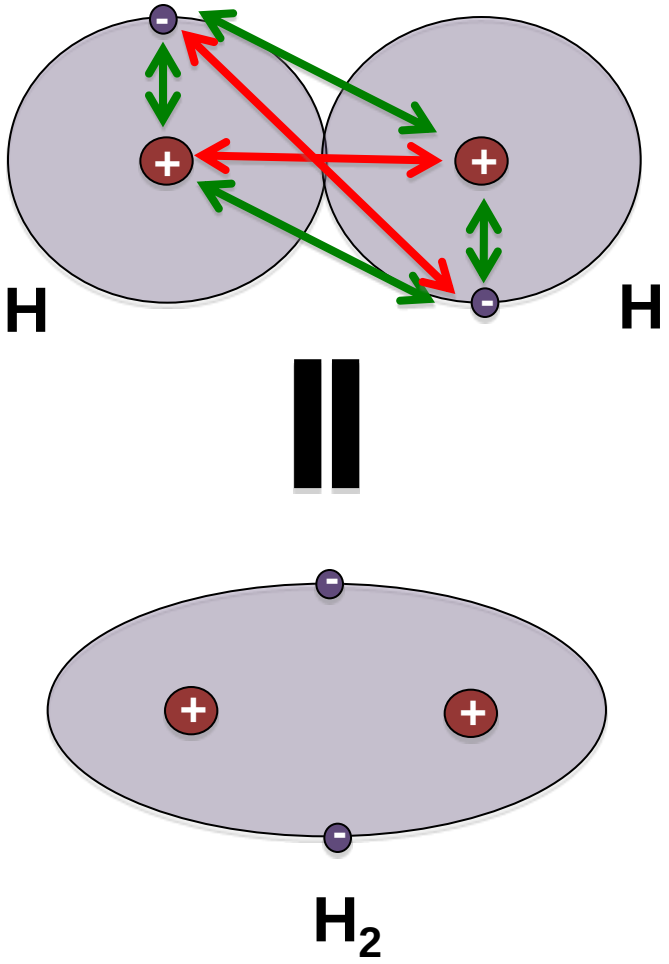
Too Close = Too Repulsive



Covalent Bond Distance = Forces Equal

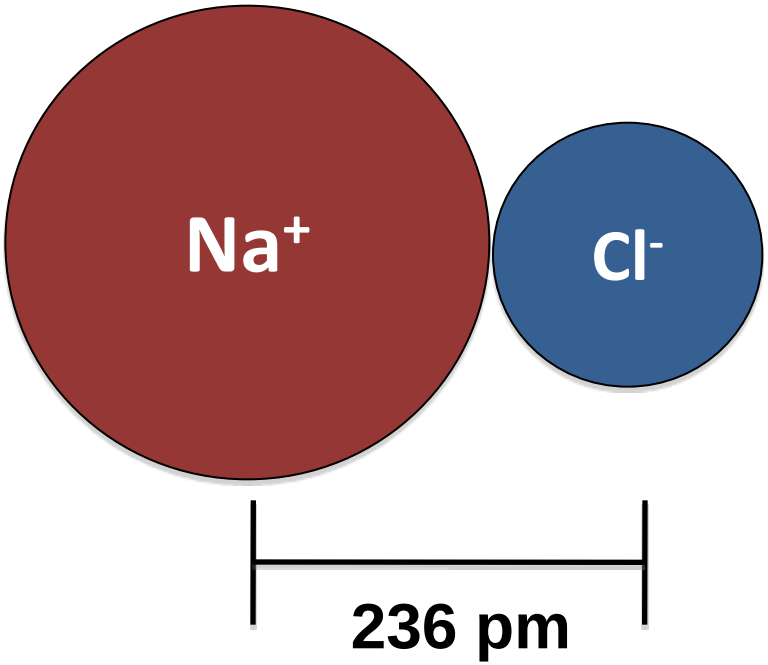
Covalent Bonds

- **Electrons** from both atoms are evenly **shared** between identical hydrogen atoms in **covalent bond**



Mini Quiz

- NaCl has a bond length of 236 pm. Why isn't the bond length 0pm?



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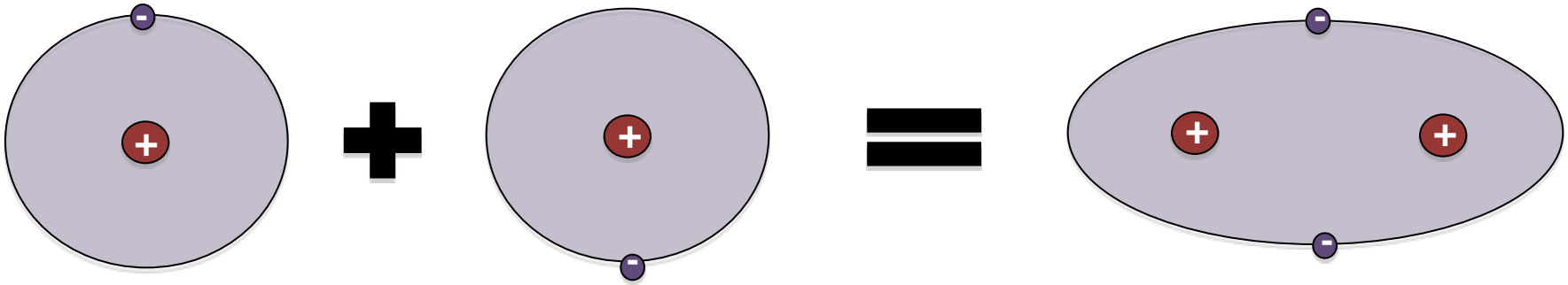
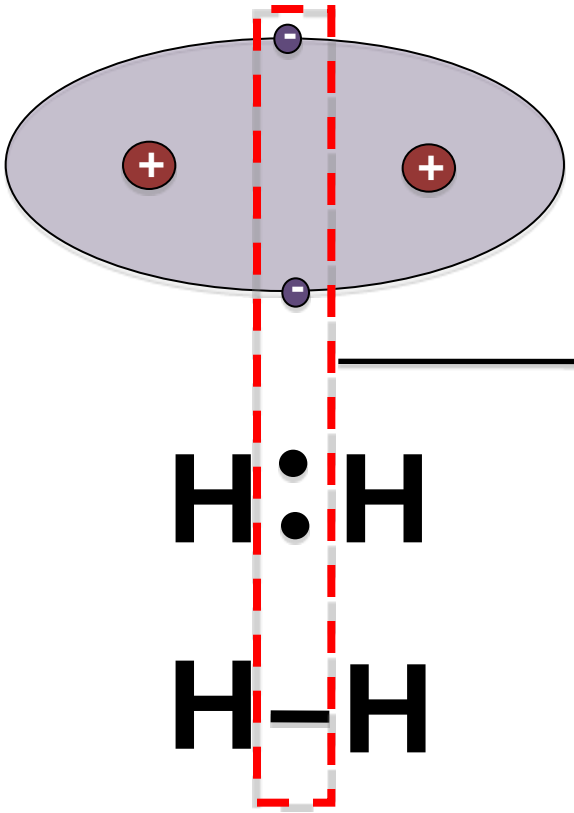
Molecular Lewis Structures

- Like atoms, molecules can be represented with Lewis structures



Molecular Lewis Structures

- Shared electrons in a covalent bond are placed between the two atoms
- A line between two atoms represents a covalent bond of two electrons



Covalently bonded (shared) electrons

H : H

H – H

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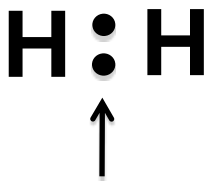
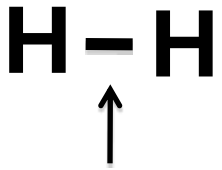
Molecular Lewis Structures

Two electrons shared between two atoms form a **single bond**



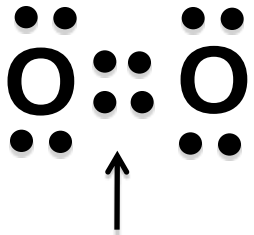
Single Bond



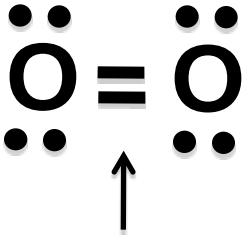
Single Bond

Four electrons shared form a **double bond**

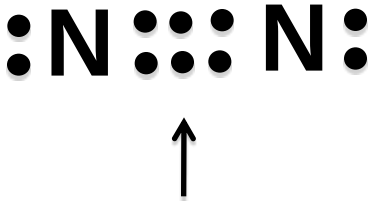
Six electrons shared form a **triple bond**



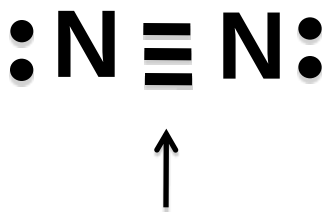
Double Bond



Double Bond



Triple Bond



Triple Bond

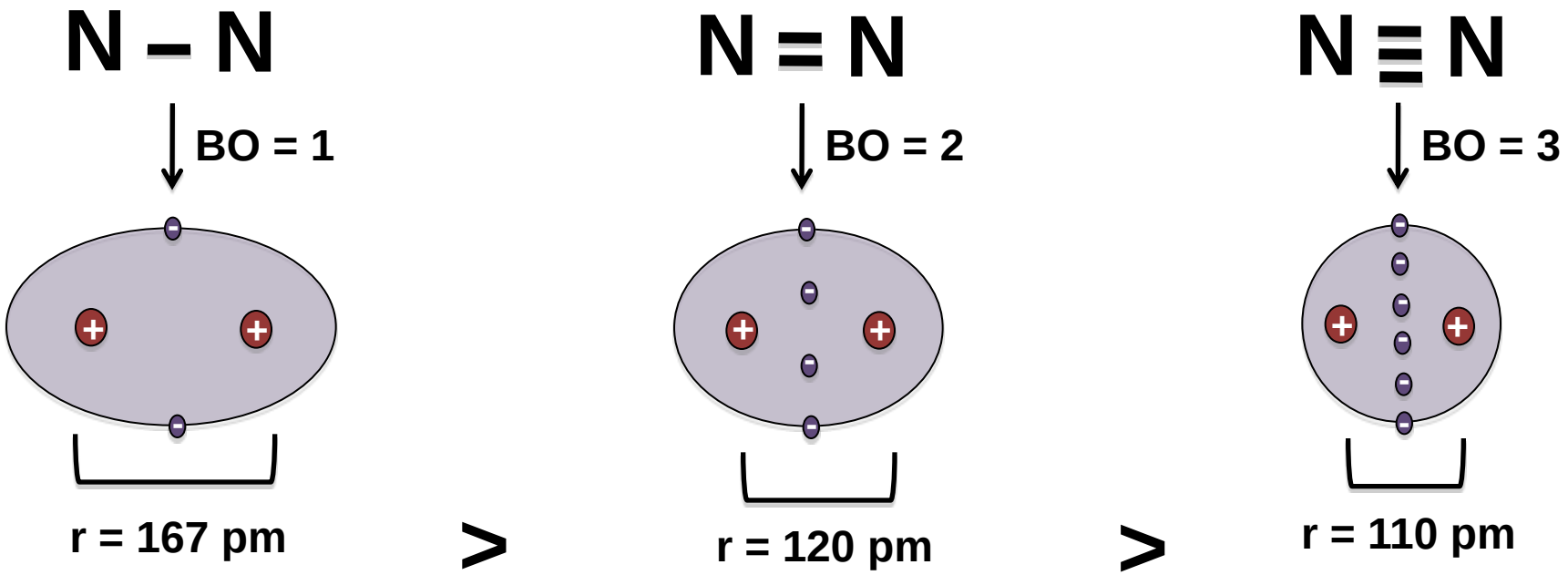
Bond order refers to the number of bonds between two atoms

What is the bond order of each of the molecules above?

Molecular Lewis Structures

Higher bond order = shorter bond length (why?)

1. Higher bond order = more electrons shared
2. More electrons shared = more negative charge between nuclei = more attraction of nuclei to middle of molecule
3. More nuclear attraction to middle = shorter bond length



Mini Quiz

Match the following bonds and bond lengths:

C – C	**C = C**	**C ≡ C**
120 pm	**154 pm**	**134 pm**

How do we draw Lewis structures with more than two electrons?

To get started, we need one more concept: **Electronegativity**

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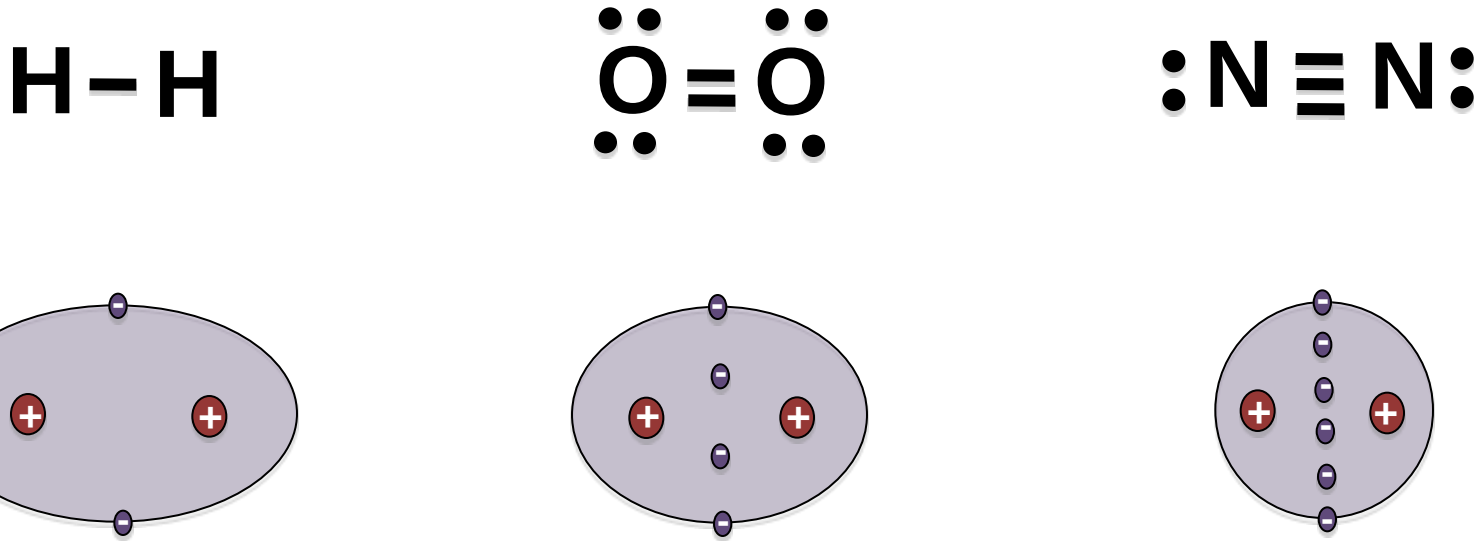
Electronegativity

- We have looked at **homonuclear diatomic** molecules so far
- **Homonuclear** = same element; **diatomic** = two atoms
- In **homonuclear diatomic** molecules**, electrons** are **shared equally** between the two atoms



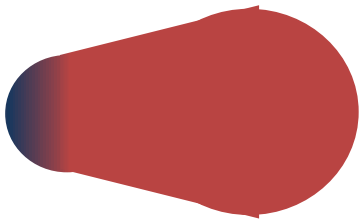
Electronegativity

- In **heteronuclear diatomic** molecules, **electrons** may **not** be **shared equally**
- **Heteronuclear** = different elements

H – H

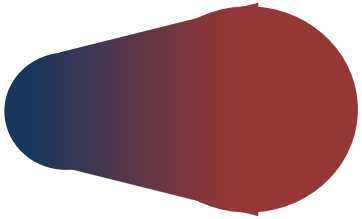
e- shared equally

H – C



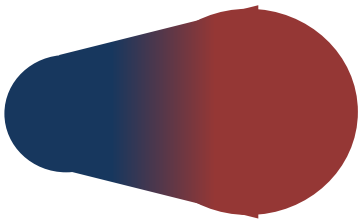
e- shared almost equally

H – O



most e- around O

H – Cl



almost all e- around Cl

Positively Charged

Negatively Charged

Electronegativity

- **Electronegativity** describes the **tendency** of an **element to 'steal' electrons** from another element in a bond
- Higher electronegativity = higher tendency to steal electrons

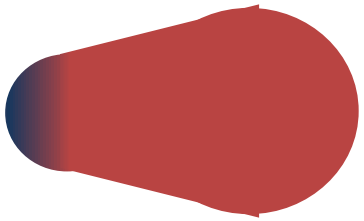
EN(H) = 2.2

H – H

e- shared equally

EN(C) = 2.55

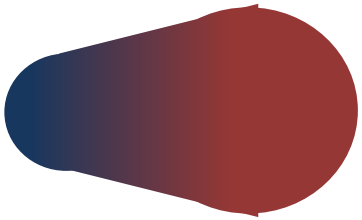
H – C



e- shared almost equally

EN(Cl) = 3.16

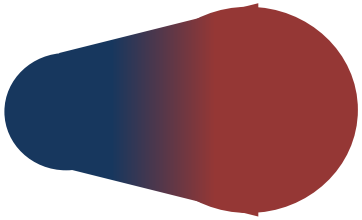
H – Cl



most e- around Cl

EN(O) = 3.44

H – O



almost all e- around O

Positively Charged

Negatively Charged

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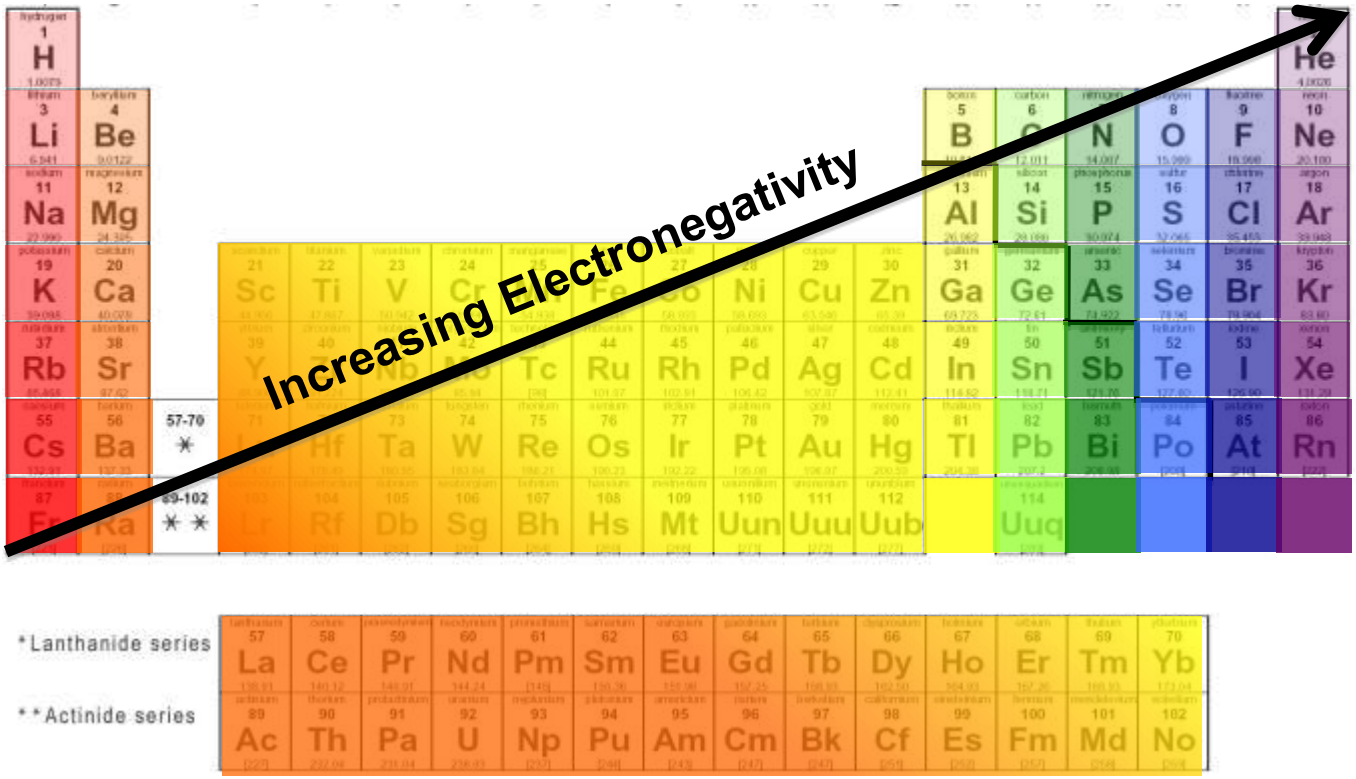
Electronegativity

- **Electronegativity increases across a period** due to increasing effective nuclear charge (Z_{eff}), which pulls electrons closer
- **Electronegativity decreases down a column** as distance between nucleus and valence electrons increases



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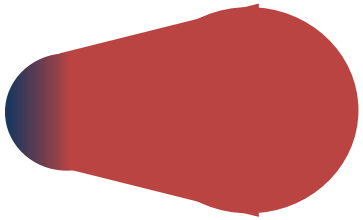
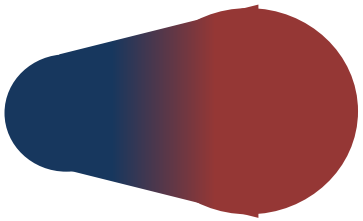
Electronegativity

- **Polar covalent bonds** = bonds between atoms with **large differences in electronegativity**
- **Non-polar colvalent bonds** = bonds between atoms with **small differences in electronegativity**

Bond	Electron Density	ΔEN	Polar vs Non-polar
H – C		EN(C)-EN(H) = 2.55-2.2= 0.35	Non-polar
H – O		EN(O)-EN(H) = 3.44-2.2= 1.22	Polar

Electronegativity

Chemical bonds fall on a continuum from non-polar to ionic

Positively Charged

Negatively Charged

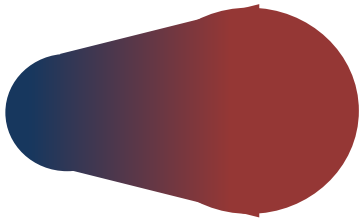
δ+ δ-

H – H

ΔEN(H) = 0

δ+ δ-

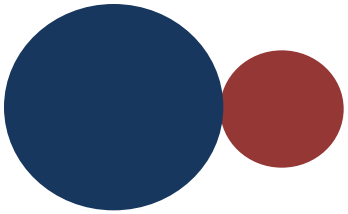
H – Cl



ΔEN(H) = 0.94

\+ -

Na – Cl



ΔEN(H) = 2.1

0.0 Mostly Covalent 0.4 Polar Covalent 1.7 Ionic 3.3

Mini Quiz

List the following elements
in order of increasing electronegativity:

F **Li** **N** **Fr** **Si**

Predict whether each bond below is
non-polar, polar covalent, or ionic:

O–O **Cs–O** **C–O**

Summary

- Electrons are **shared** in covalent bonds
- We can Lewis structures to indicate covalent bonds
- Differences in electronegativity explain unequal sharing of electrons in a bond

Homework

- TBD