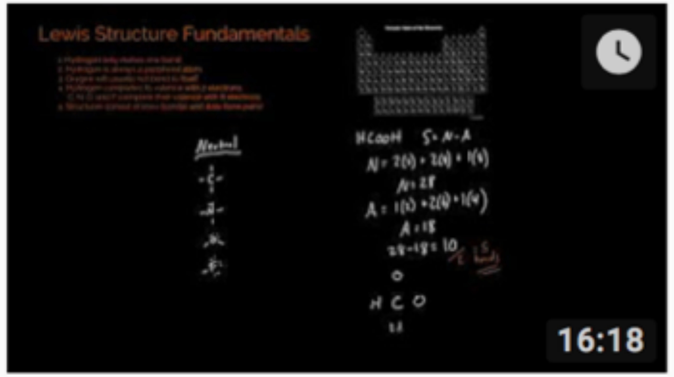
CH301 Unit 2

REVIEW THREE: PERIODIC TRENDS, INTRO TO BONDING

Foundations for these topics:

Just one YouTube video this week:



Lesson Three: Lewis Structures (Bonding)

Jimmy Wadman • 43 views • 3 months ago

In this video, we will discuss a couple different intuitive approaches to drawing basic **Lewis Structures**. Later on, we will discuss ...

Periodic Trends and Bonding

Periodic Table of the Elements

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.20	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (281)	111 **Rg** (281)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (289)	116 **Lv** (293)	117 **Ts** (293)	118 **Og** (294)

58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

- In chemistry, we are primarily concerned with the **valence electrons**
- The periodic trends in ionization energy, electron affinity, Z-effective, atomic radius, and electronegativity help to explain the behavior of valence electrons
- An element with a **very low ionization energy** is more likely to form a **cation** in our atmosphere
- An element with a **very high electron affinity** is more likely to form an **anion** in our atmosphere

* Imperfect Trends! *

Basic Periodic Table Trends

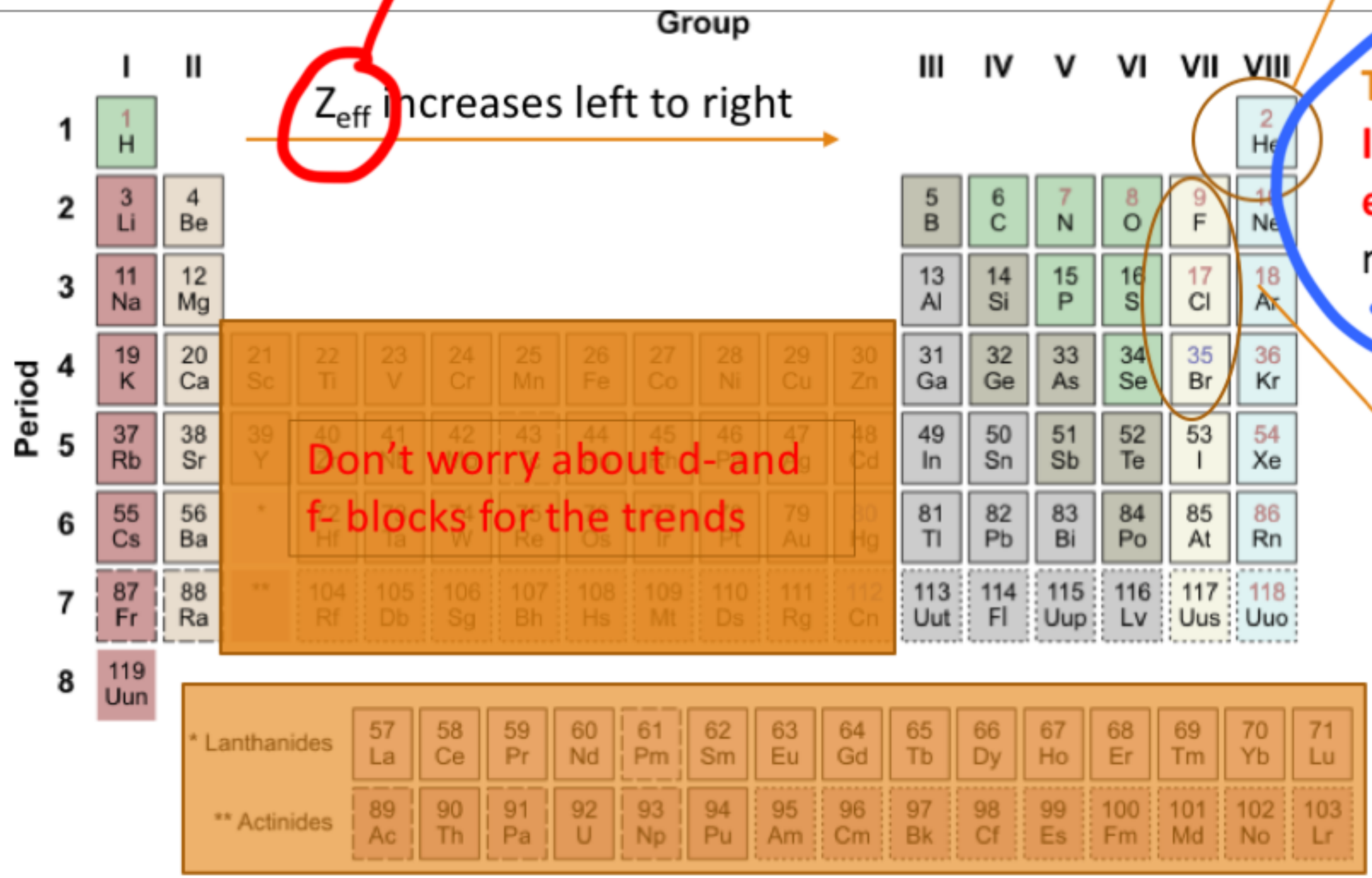
- Ionization Energy: the energy required to **remove an electron** from an atom in the gas phase *
  - Trend: generally increases going to the top right of the periodic table **(has notable exceptions)**
  - Increases with each successive ionization (the second ionization energy is greater than the first, and so on)
- Electron Affinity: electron affinity is the energy released to **add an electron** to an atom in the gas phase *
  - Trend: Electron affinity does not follow a stable trend; however, it generally trends toward the top right of the periodic table.
- Atomic Radius: size of the nucleus and its electrons (most of the size is in the space taken up by the electron clouds)
  - Trend: increases as you go to the bottom left where shielding is maximized and z_{eff} is minimized
- Z_{eff}: a measurement of the **pull of the outer (valence) electrons by the nucleus**
  - Trend: increases as you go from left to right across a period
  - For ground state atoms, Z_{eff} is simply the group number

Periodic Trends

→ Positive pull experienced by outer e^-

Helium has the maximum ionization energy

Shielding Increases going down

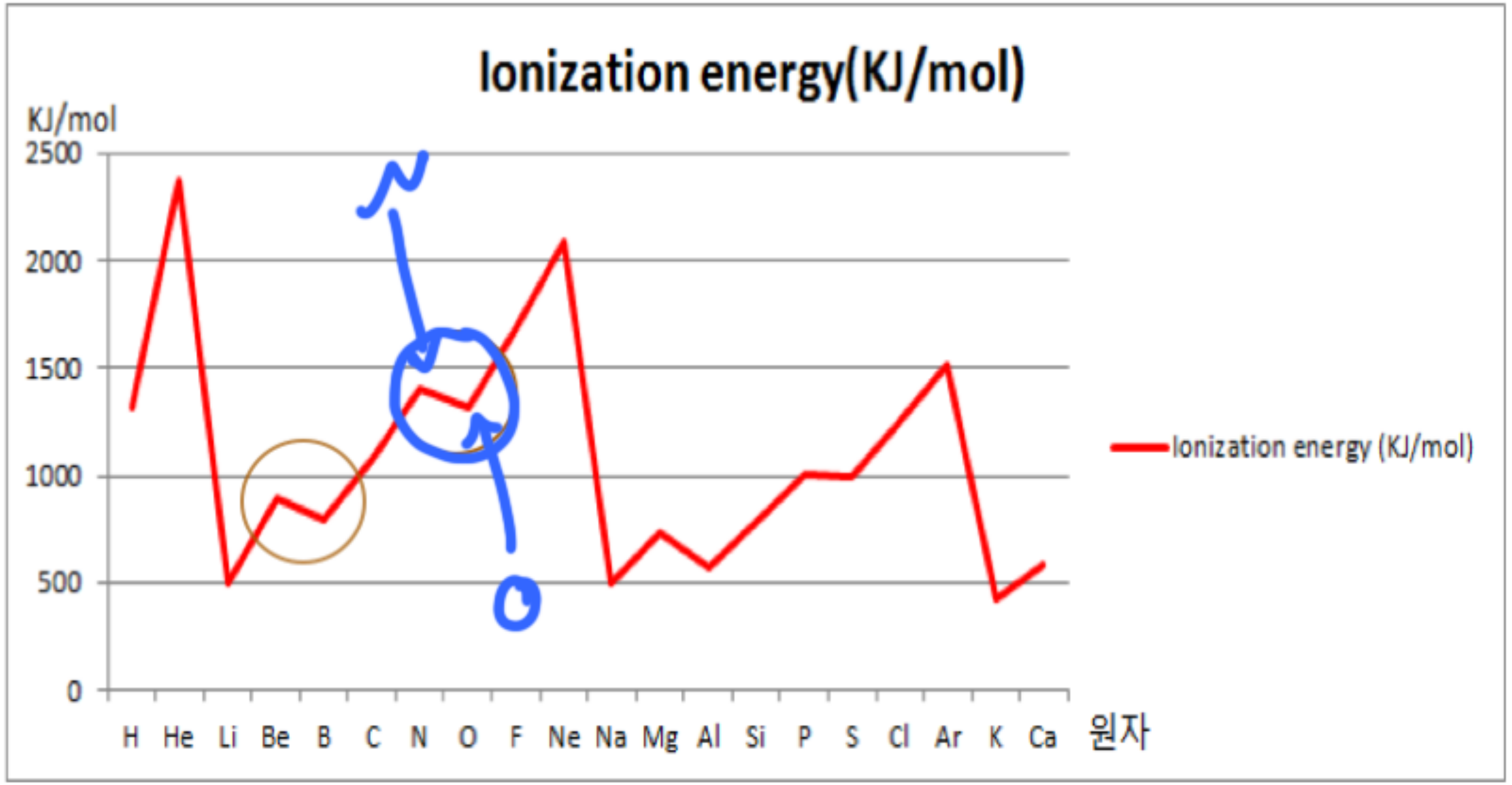


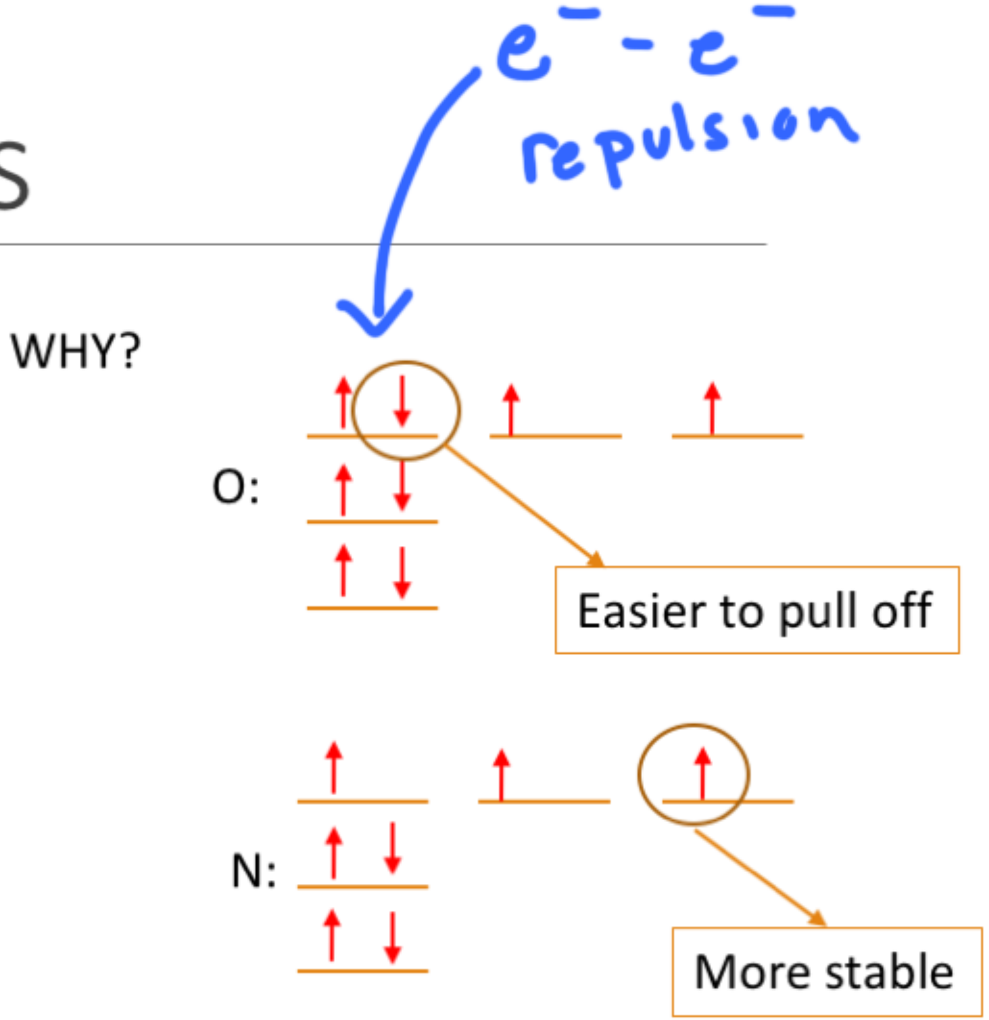
TOP RIGHT: **High Ionization energy, high electron affinity,** small radius

The halogens have the highest electron affinity

Bottom Left: Low ionization energy, low electron affinity, **large radius**

Ionization Energy Exceptions

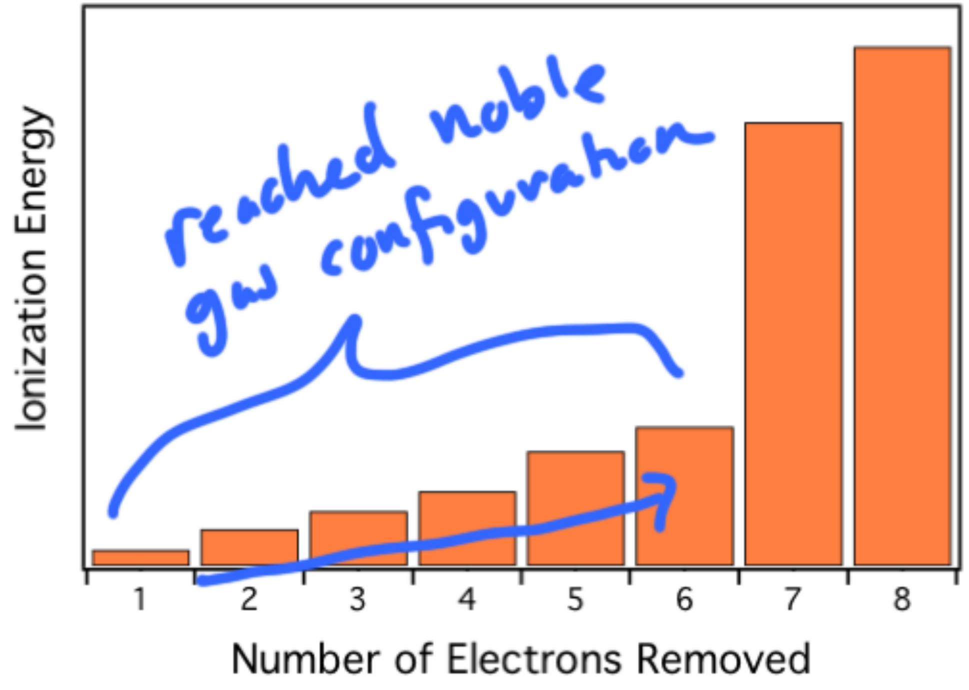




The "special" quantum mechanical stability of a fully filled s or p orbital, or a half-filled p orbital, causes exceptions to the trend of ionization energy

Note: all S atoms in these equations are in the gas state

Successive Ionization Energies



Energy increases for each successive ionization energy for the same atom

Big jump in IE means S^{6+} is most likely the noble gas configuration

1. $S \rightarrow S^{+} + e^{-}$
2. $S^{+} \rightarrow S^{2+} + e^{-}$
3. $S^{2+} \rightarrow S^{3+} + e^{-}$
4. $S^{3+} \rightarrow S^{4+} + e^{-}$
5. $S^{4+} \rightarrow S^{5+} + e^{-}$
6. $S^{5+} \rightarrow S^{6+} + e^{-}$
7. $S^{6+} \rightarrow S^{7+} + e^{-}$
8. $S^{7+} \rightarrow S^{8+} + e^{-}$

Ionic Radius

Note on this diagram:
Neutral atoms are grey
Cations are red
Anions are blue

Last year's exam:

For an isoelectronic series of ions, the ion that is the smallest is always

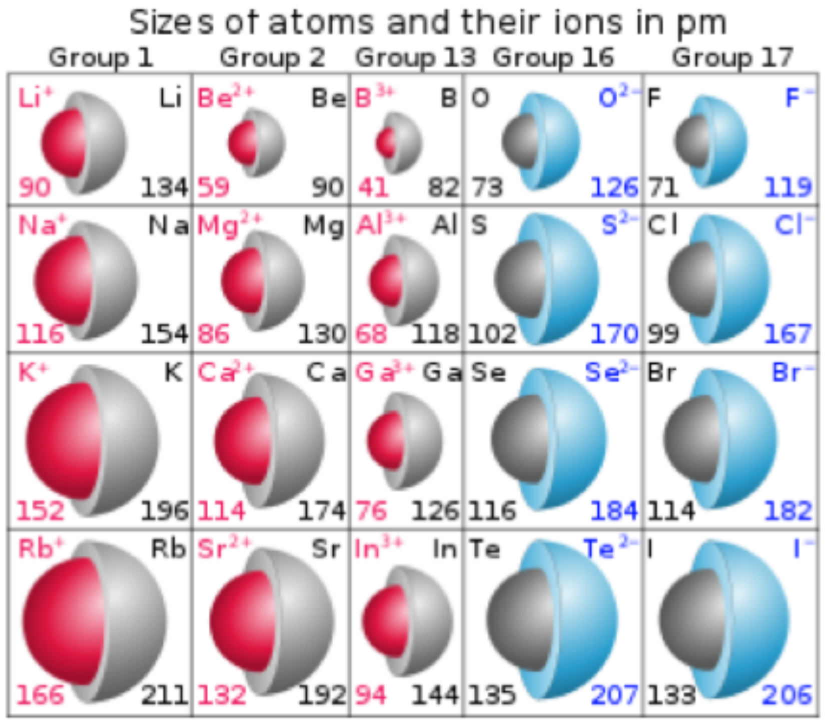
1. the ion with the highest atomic number.

2. the ion with the most electrons.

3. the ion with the most neutrons.

4. the ion with the fewest protons.

5. the least positively (or most negatively) charged ion.



For ionic radius, there are two good comparisons you can make:

- **For isoelectronic ions, the larger the Z (positive nuclear charge), the smaller the radius**
- For ions of the same element, the more negative the charge, the larger the radius

X, X^+, X^-

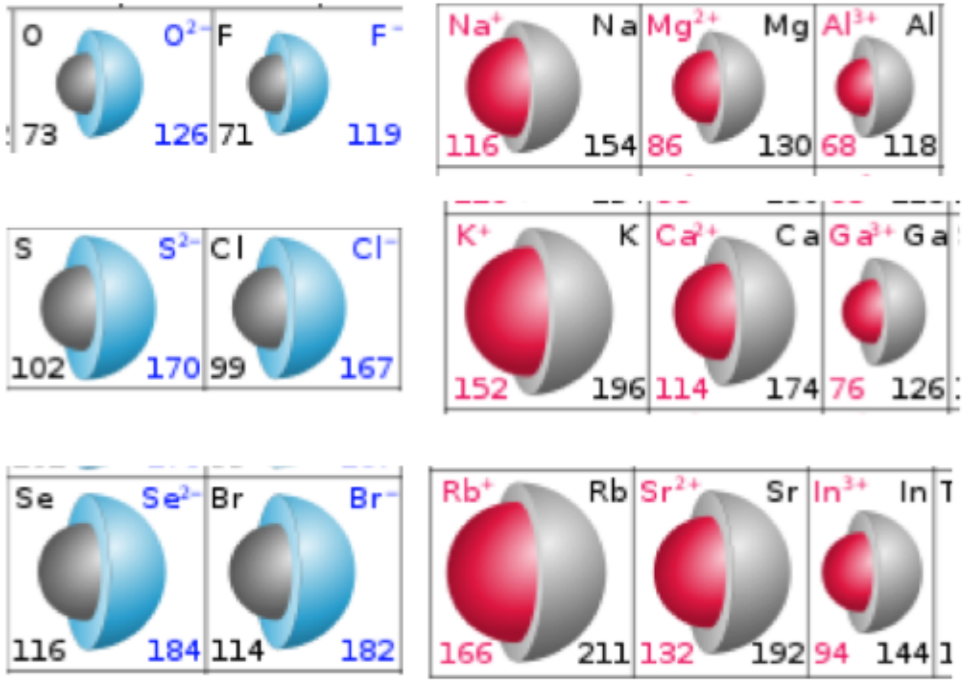
Comparing Isoelectronic Ions

Same electronic configuration; increasing Z; decreasing radius

[Ne]:

[Ar]:

[Kr]:



Note on this diagram:

Neutral atoms are grey

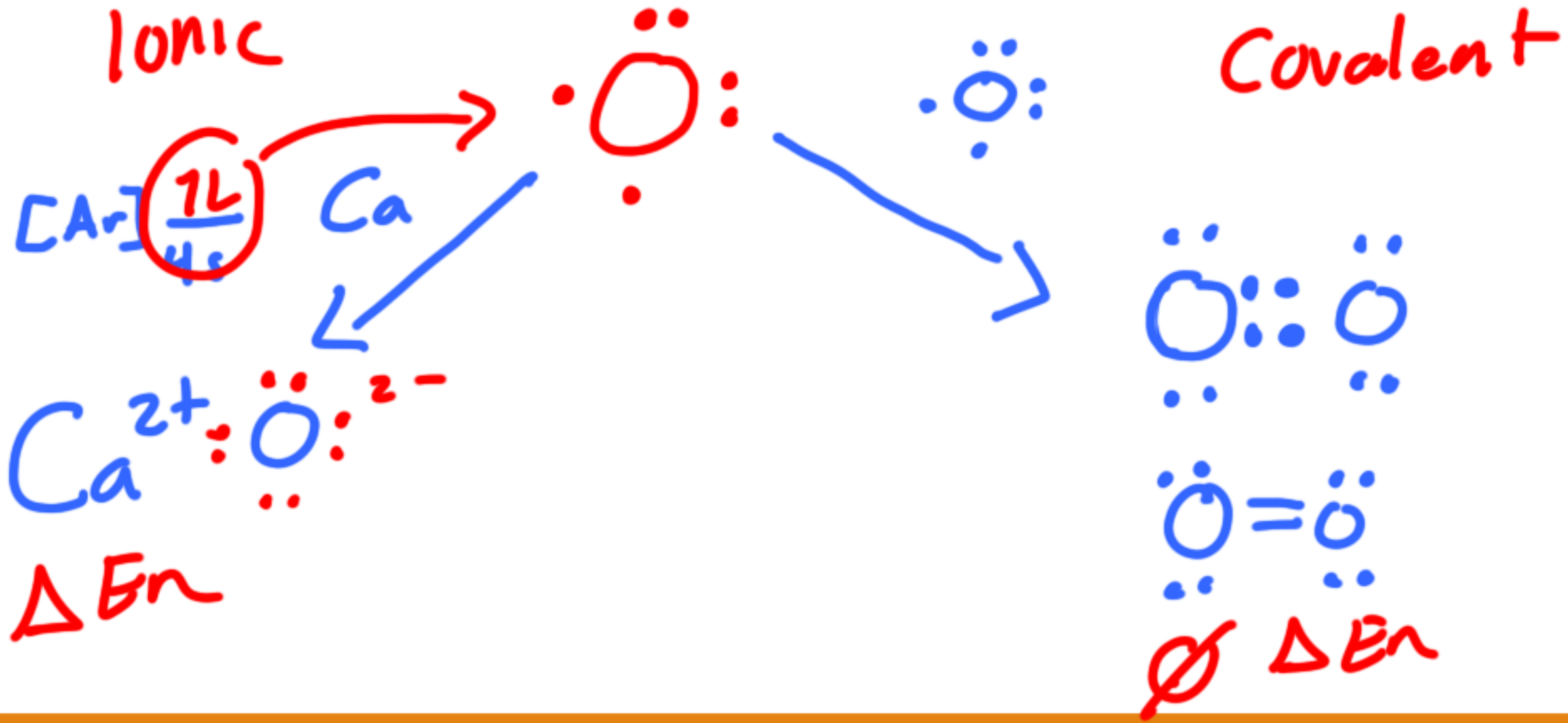
Cations are red

Anions are blue

Bonding Fundamentals

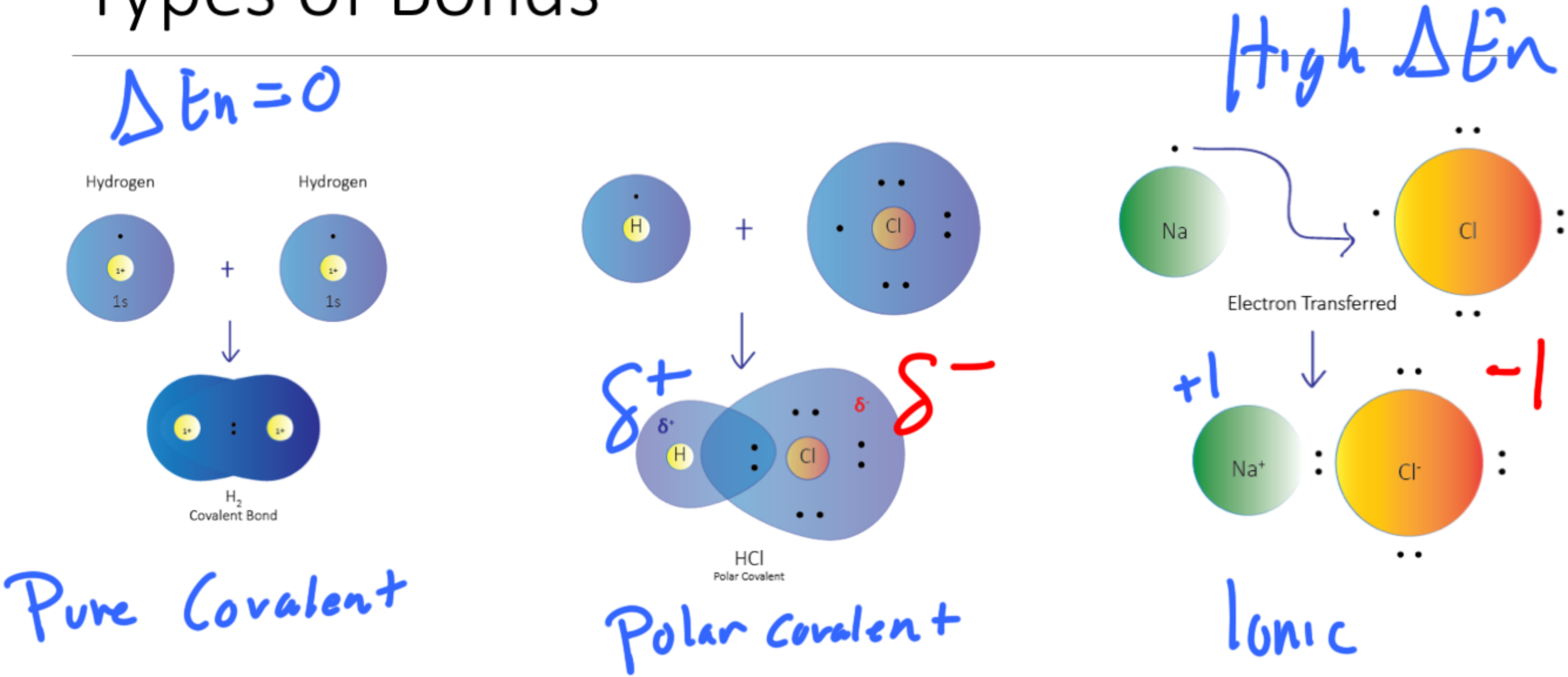
What is the point of making a bond?

All chemical species want to reach the lowest possible energy state. A complete valence is a very common "desired" final state for elements, which is accomplished by sharing or transferring electrons.



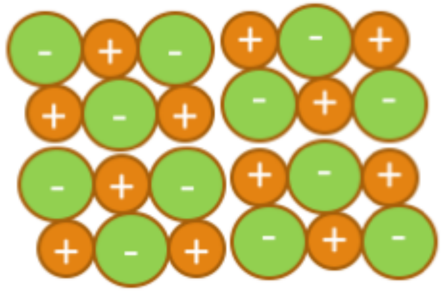
Increasing ΔEn

Types of Bonds

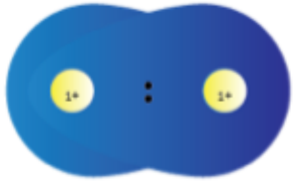


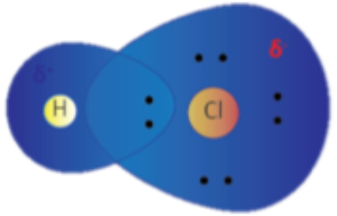
Ionic vs. Covalent Bonds

- Ionic bonds are the electrostatic interaction between discrete fully positive and fully negative charges.
  - Metal – Nonmetal
  - Ionic bonds are quantified by lattice energy
  - Ionic compounds are almost always a metal (+) and nonmetal (-)
  - The ions can be elements or polyatomic ions (which are, themselves, covalent compounds)
  - **Results in an alternating lattice of discrete positive and negative charges**



- Covalent bonds are the sharing of electrons between two species
  - Nonmetal – Nonmetal
  - Covalent bonds are quantified by bond length, bond strength, bond order, and polarity (Δen)
  - Nonpolar covalent bonds result from equal sharing of electrons between two nuclei
  - Polar covalent bonds result from the electrostatic interaction between partial charges (∂^+, ∂^-)
  - Results in molecules





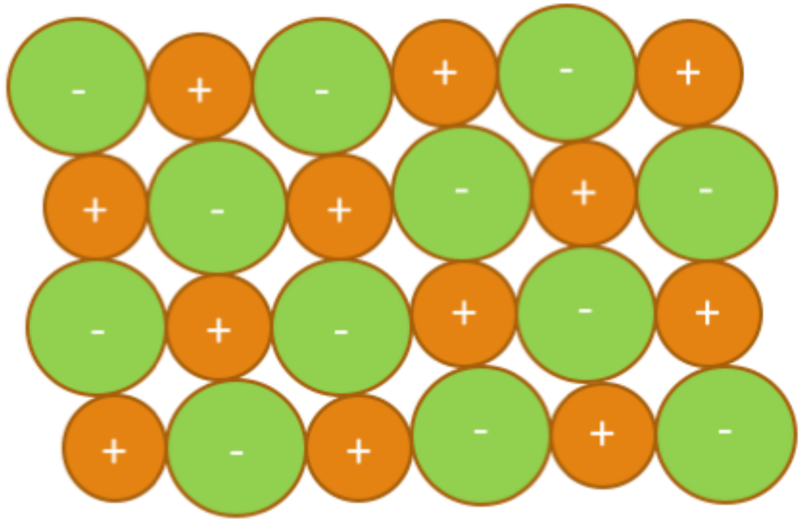
Lattice Energy

Is lattice energy positive or negative?

- Depends on how you're measuring it. But use the magnitude to answer homework and exam questions.
- **Lattice energy indicates the amount of energy necessary to overcome the negative potential energy binding the charges of an ionic compound**

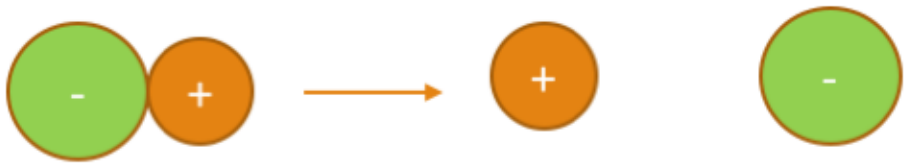
- Lattice Energy: is directly proportional to the charges of the ions (q_1 and q_2) and inversely proportional to the size of the ions (r)

$$\Delta H_{lattice} \propto \frac{q_1 q_2}{r}$$



- This describes the energy required for the following reaction:

$$NaCl \rightarrow Na^{+} + Cl^{-}$$



When Ranking Lattice Energies:

1. Prioritize **charge first**
2. If two ionic compounds have the same charge, the **smaller one will have a greater lattice energy**
3. Polyatomic ions are **big**

Lattice Energy: Conceptual Questions

- Which has the greater lattice energy?

+1 -1
$NaCl$

vs.

+2 -1
$CaBr_2$

Lattice Energy: Conceptual Questions

- Which has the greater lattice energy?

+2 -1
$MgBr_2$

vs.

+2 -2
$CaCO_3$

Lattice Energy: Conceptual Questions

- Which has the greater lattice energy?

+2 −2
MgO vs. $CaSO_4$
+2 −2

big

r↑ E↓

Lattice Energy: Challenge Question

List the following in terms of theoretical decreasing lattice energy (greatest to smallest):

$MgO, MgSO_4, Na_2S, LiF, LiBr, NH_4Br$

+1,-1   +2,-2   +1,-2   +1,-1   +2,-2   +1,-1

LiF, MgO, Na_2S, LiBr, $MgSO_4$, NH_4Br

+1,-1

LiF

~NH_4Br

LiBr

+1,-2

Na_2S

+2,-2

MgO

$MgSO_4$

least

$NH_4Br < LiBr < LiF < Na_2S < MgSO_4 < MgO$

greatest

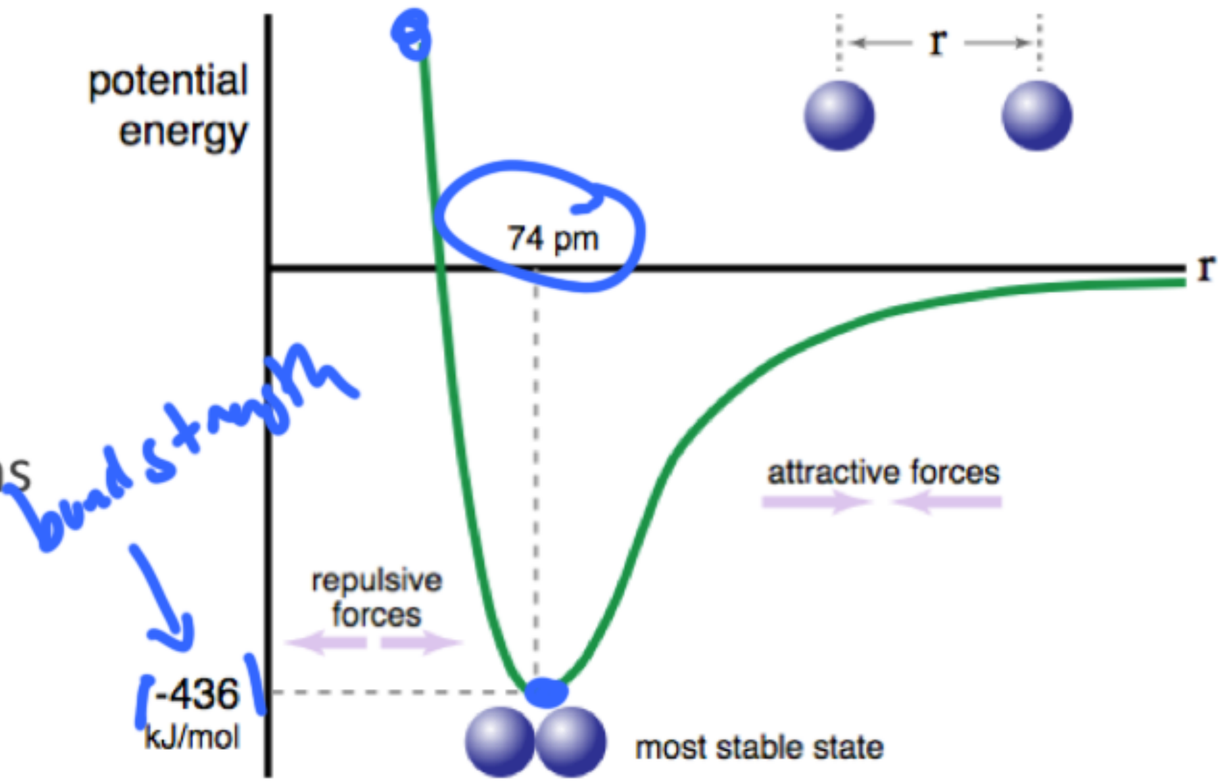
Covalent Bond Definitions

- Bond strength: the energy required to break a covalent bond (also known as Bond Dissociation Enthalpy)
- Bond order: a measurement of the average number of chemical bonds between two atoms in a species
  - This gets more detailed as time goes on; for now, think of single, double, and triple bonds
- Bond length: the distance between atoms in a bond
- Electronegativity: a measurement of the electron-withdrawing nature of an element in a bond. Leads to unequal sharing of electrons in a covalent bond, or polarity.
- Resonance: multiple acceptable Lewis Structures are available for a given compound, meaning that the compound actually exists as the average of all acceptable structures
- Delocalized electrons: in the case of resonance, electrons are not confined to a single bond. Instead they “resonate” over multiple bonds. We call this delocalization.

Note: bond strength and order are both indicators of stability. Higher bond order means a higher bond strength, which also correlates inversely with bond length.

Covalent Bond Strength and Stability

- **Bond strength**: the energy required to break a covalent bond (also known as Bond Dissociation Enthalpy)
- **Bond order**: a measurement of the average number of chemical bonds between two atoms in a species
- **Bond length**: the distance between atoms in a bond
- **KEY: stronger bonds are shorter; higher order bonds are stronger (and thus shorter)**



Notice how when a covalent bond gets just right, your potential energy is **most negative** just like lattice energy!

Lewis Structures

- Lewis structures are the main form of drawing molecules. There are two important equations to know when drawing Lewis Structures:

$$S = N - A$$

- The number of shared electrons is equal to the total electrons **needed** to fill the valence (2 for hydrogen, 8 for other non-exception elements) minus the electrons **available** (valence electrons). Account for positive charge by subtracting from available; account for negative charge by adding to available.
- **You can then calculate the number of bonds by dividing the shared electrons by two (bonds = S/2)**

$$FC = \text{Valence} - (\text{lone electrons} + \text{bonds})$$

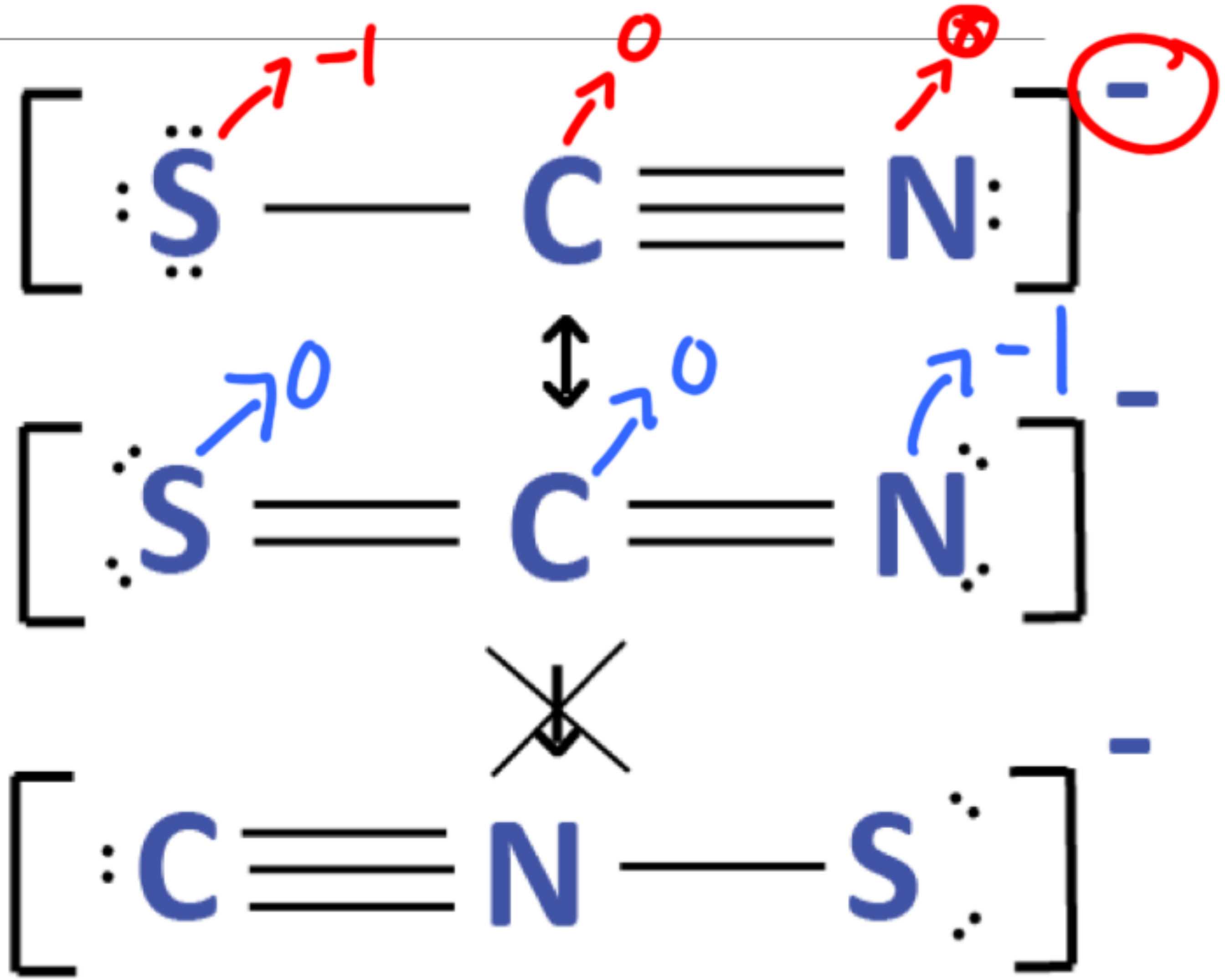
- **The formal charge of any atom in a molecule is equal to the the valence electrons of the atom (as seen on the periodic table) minus the "things it's touching" in the Lewis structure (the lone electrons and the number of bonds)**

$$FC = \text{Valence} - (\text{"things it's touching"})$$

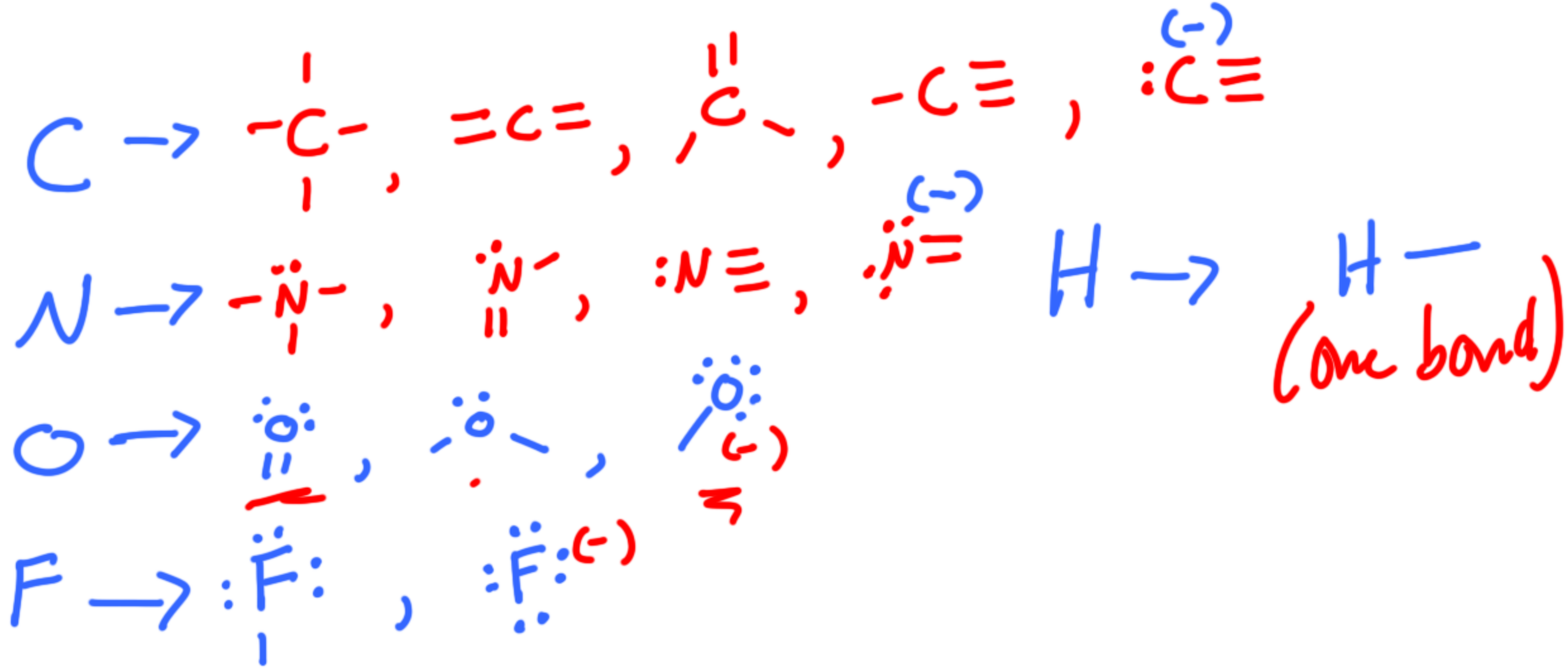
Multiple Structures: Resonance

- What constitutes an "acceptable" Lewis Structure?
  - Try to place the negative charge on only the electronegative atoms before your central atom
  - Any individual charge cannot exceed +/- 1
  - Put your charges on as few atoms as possible
  - **Your structure accurately reflects the number of available electrons in your atoms**

Sometimes there can be multiple structures that follow your basic rules. These rules are an extra test to show that you are drawing the best structure.



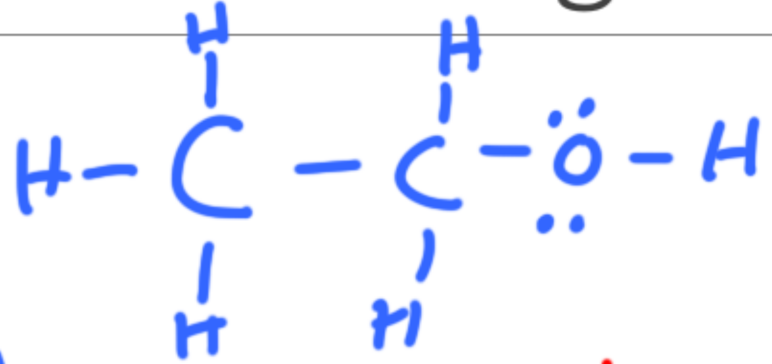
Lewis Structure Intuition



Lewis Structure Drilling

See last page for solutions

1. CH_3CH_2OH



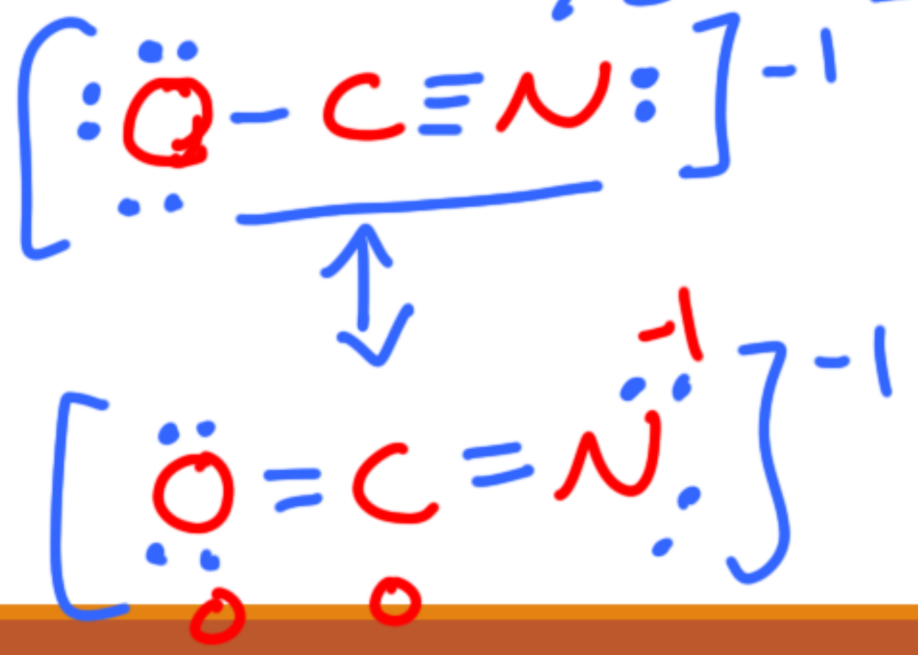
2. OCN^-

S = N - A

→ 6 + 4 + 5 + 1 = 16 e⁻

→ 8 + 8 + 8 = 24 e⁻

24 - 16 = 8/2 = 4 bonds



3. $C_2O_4^{2-}$

4. NH_4^+

Lewis Structure Formal Charge Foundations

Element	-1 Charge	Neutral	+1 Charge
Hydrogen	x	1 bond	x
Carbon*	3 bonds, 1 LP (rare)	4 bonds	x
Nitrogen	2 bonds, 2 LP (rare)	3 bonds, 1 LP	4 bonds
Oxygen	1 bond, 3 LP	2 bonds, 2 LP	3 bonds, 1 LP
Halogens**	x	1 bond, 3 LP	x

This chart represents a "best guess" to draw out simple Lewis Structures. The formal charges are calculated based on a full octet and no exceptions.

*Carbon almost always has four bonds. Some notable exceptions are CO, CN^-, CNO^-

**Represents halogens are a peripheral atom, not a central atom. In other words, this applies to the fluorine in ClF_3 but not the chlorine.

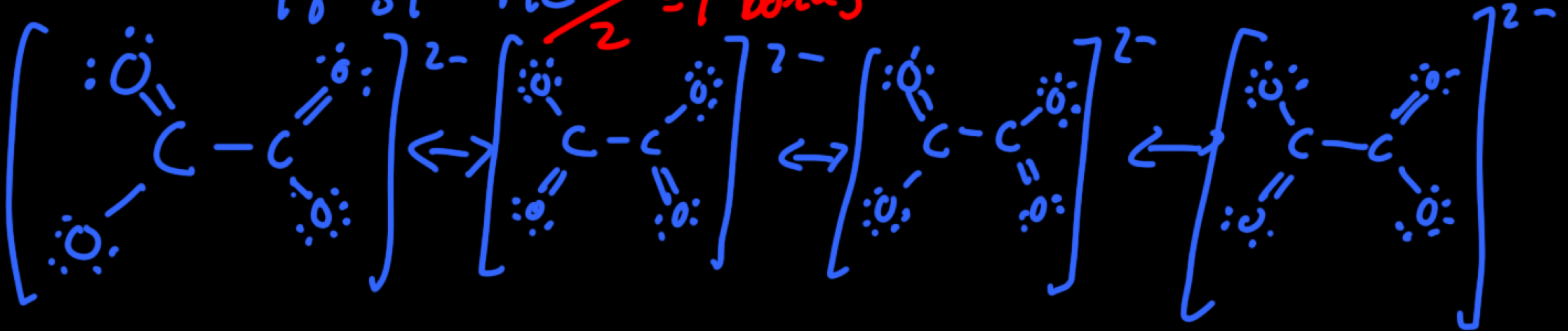
3. $C_2O_4^{2-}$

$S = N - A$

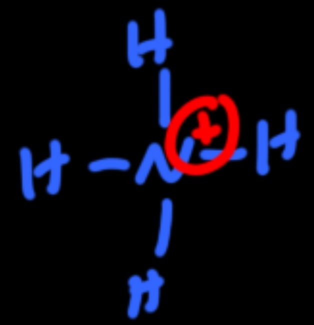
$\hookrightarrow 2(4) + 4(6) + 2 = 34\ e^-$

$\hookrightarrow 2(8) + 4(8) = 48\ e^-$

$48 - 34 = 14e^- / 2 = 7$ bonds



4. NH_4^+



$S = N - A$

$\hookrightarrow 5 + 4(1) - 1 = 8$

$\hookrightarrow 8 + 4(2) = 16\ e^-$

$16 - 8 = 8e^- / 2 = 4$ bonds