N20 – Bonding

Energy of Bonding

Link to YouTube Presentation: https://youtu.be/bJm5LGLngfo

N20 – Bonding

Energy of Bonding

Target: I can perform calculations related to lattice energy and bond energy, and can explain the connection to Coulomb's Law

Covalent

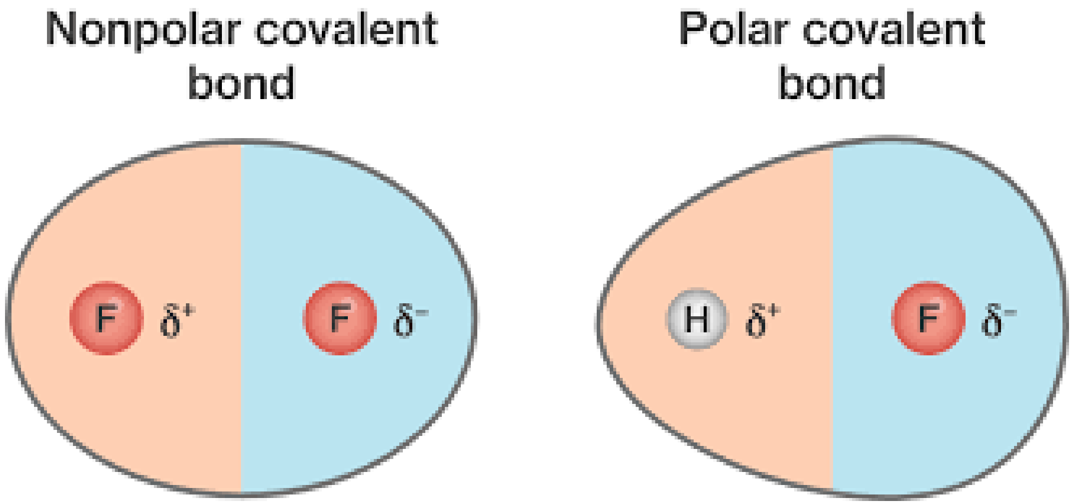
Covalent Bonds

Polar-Covalent bonds

- Electrons are *unequally* shared
- Electronegativity difference between 0.3 and 1.7

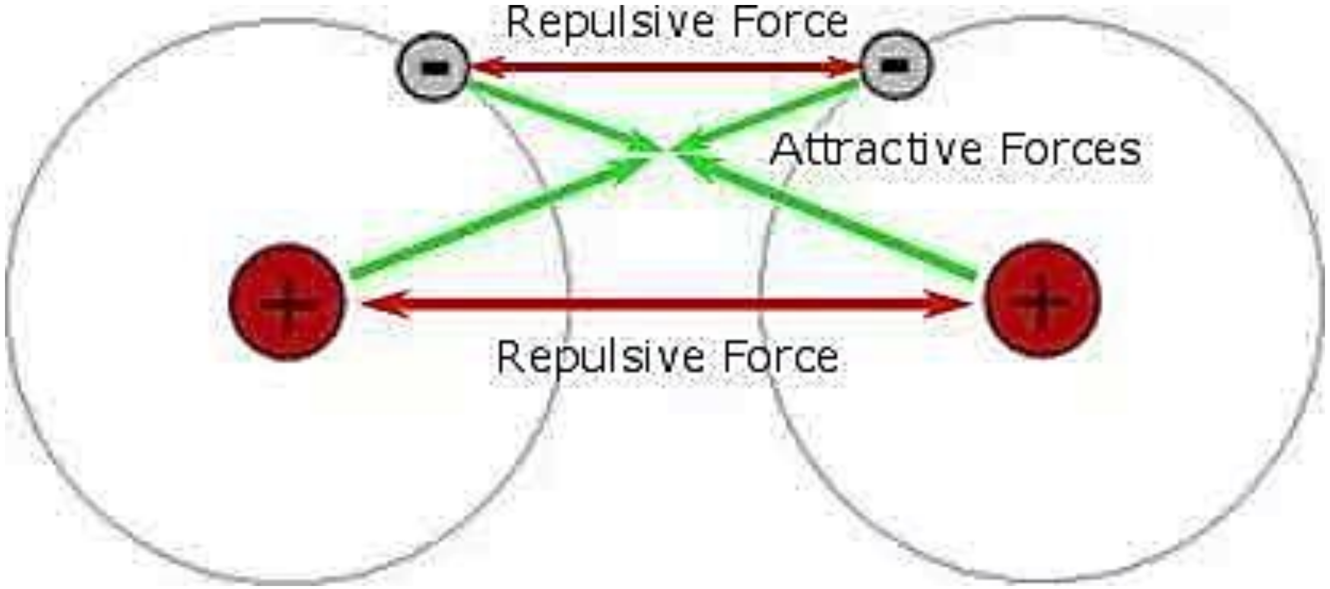
Nonpolar-Covalent bonds

- Electrons are *equally* shared
- Electronegativity difference between 0 to 0.3



Covalent Bonding Forces

- **Electron – electron repulsive forces = Bad**
- **Proton – proton repulsive forces = Bad**
- **Electron – proton attractive forces = Good**



How Close Together Before “Bonded” ?

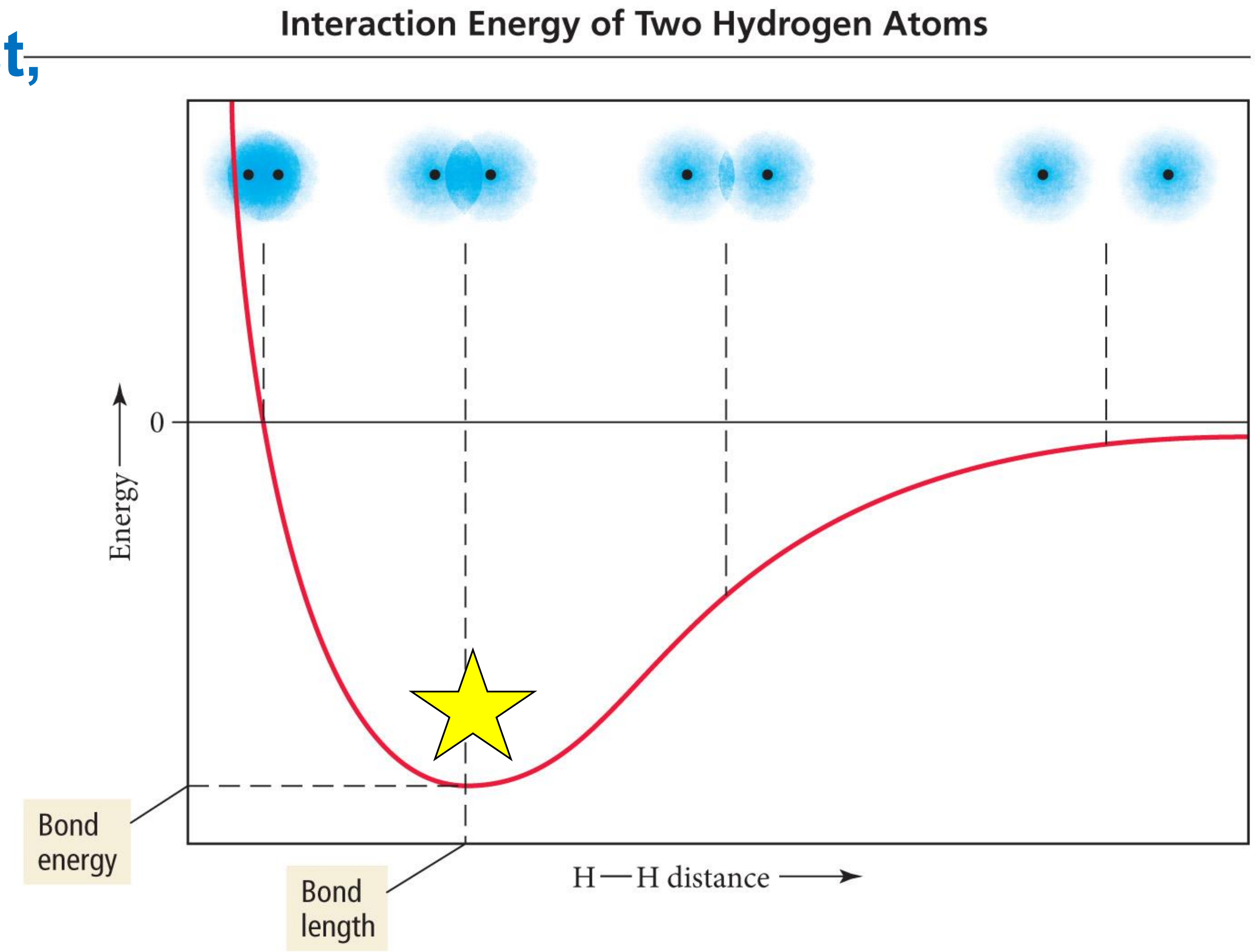
“Bonded” when at lowest, most stable energy.

Goldie Locks…

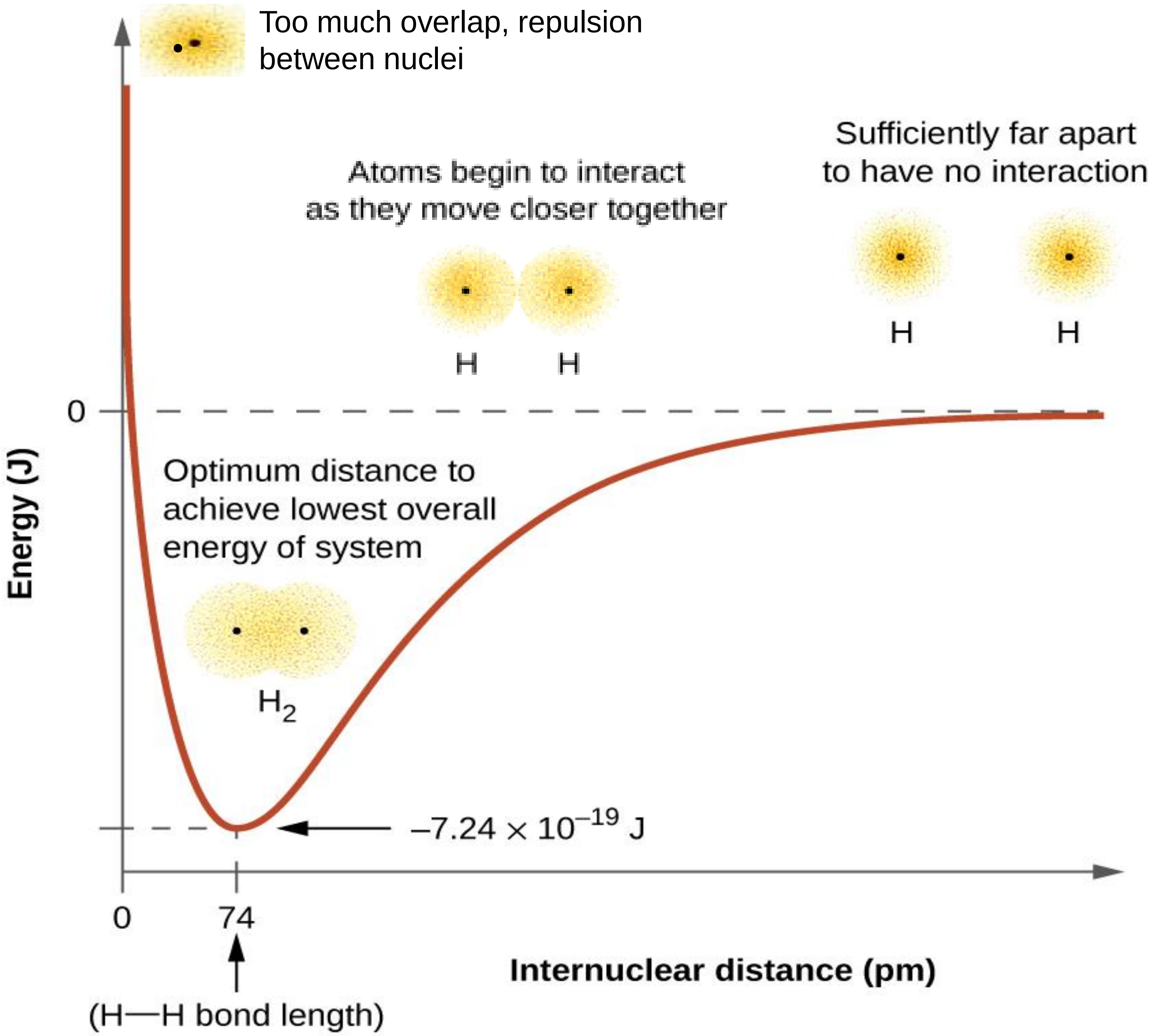
Too far = bad

Too close = bad

You want it juuust right



Bond Length Diagram



Bond Length and Energy

Bonds between elements become *shorter* and *stronger* as **multiplicity** increases.

Bond	Bond type	Bond length (pm)	Bond Energy (kJ/mol)
C - C	Single	154	347
C = C	Double	134	614
C ≡ C	Triple	120	839
C - O	Single	143	358
C = O	Double	123	745
C - N	Single	143	305
C = N	Double	138	615
C ≡ N	Triple	116	891

Bond Energy and Enthalpy

$$\Delta H = \sum D_{bonds broken} - \sum D_{bonds\ formed}$$

D = Bond energy **per mole** of bonds

Breaking bonds always requires energy

Breaking = endothermic = (+)

Forming bonds always releases energy

Forming = exothermic = (-)

Bond Energy and Enthalpy

You will see numbers vary a decent amount from chart to chart. Use what is in the problem, otherwise look them up an don't stress about slight differences.

"Takes to Break" = + endo
"Frees to Form" = - exo

How much energy does it take to break $2H_2O$ into $2H_2$ and O_2?
Bond energies: O-H 463 kJ/mol, H-H 436 kJ/mol, O=O 498 kJ/mol

- **Breaking:** 4 O-H bonds → + values, absorbed, endo, +
- **Making:** 2 H-H bonds, and 1 O=O bond → - values, released, exo, -

$$\Delta H = [\underbrace{4(463)}_{4\ O\text{-}H}] + [\underbrace{2(-436)}_{2\ H\text{-}H}+\underbrace{1(-498)}_{1\ O=O}] = 482\ \text{kJ/mol}$$

Ionic

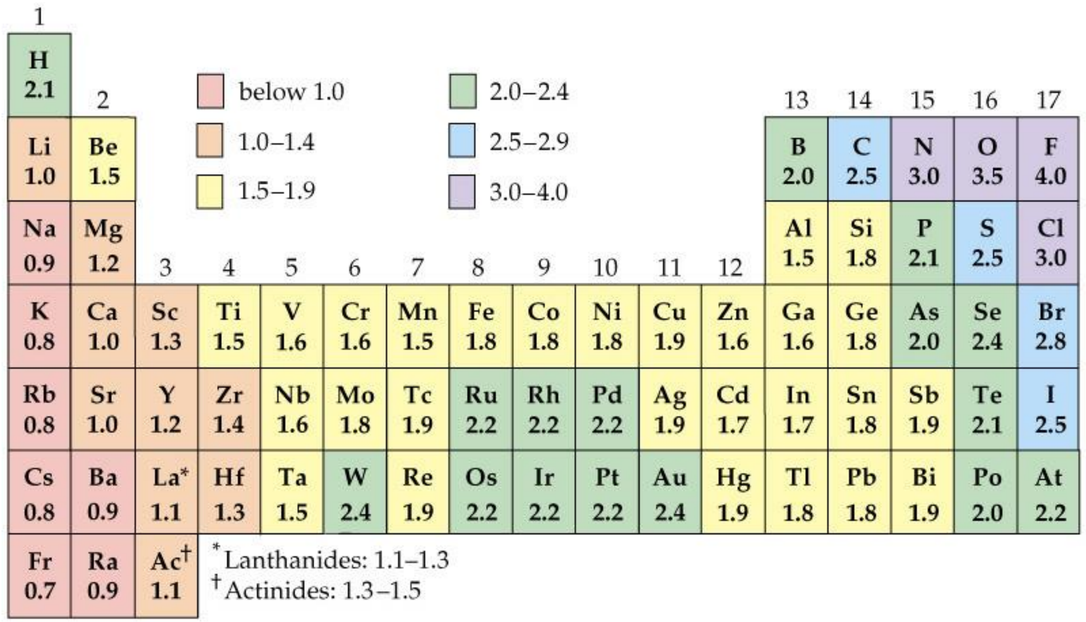
Remember these things?

Electronegativity, Atomic Radius, Ion Charge…

Why are these things important here?

Energy of bonding comes back to…

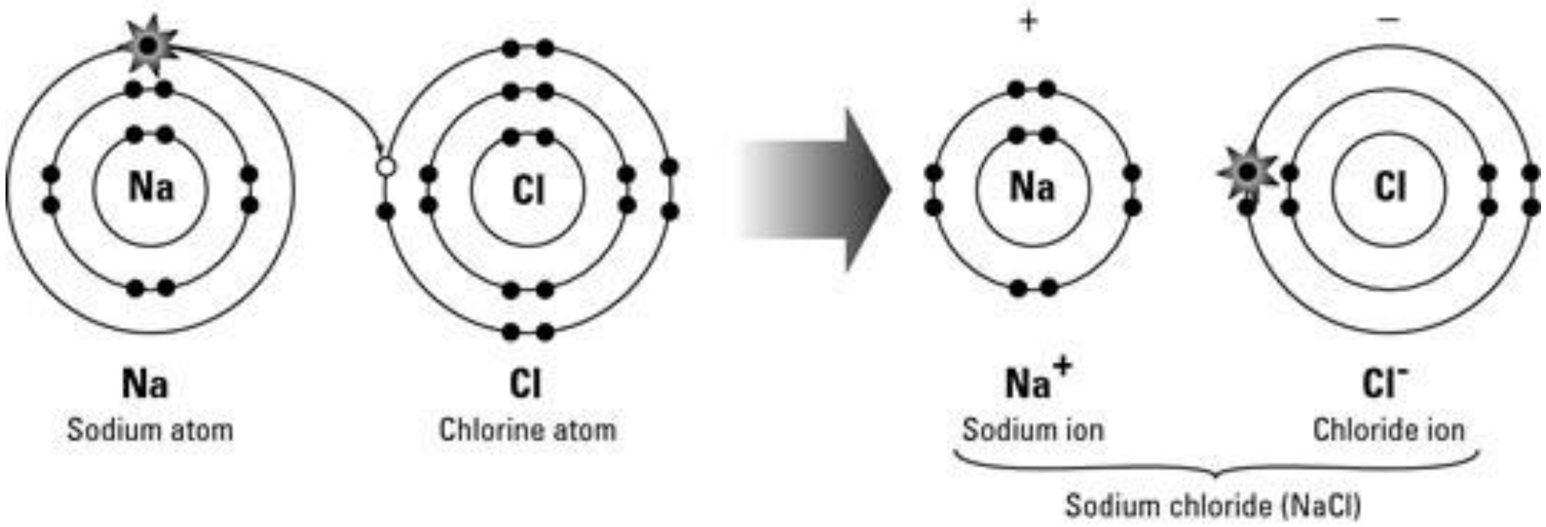
Attractions and Repulsions!



Ionic Bonds

1. e^{-}'s transferred → makes ions
2. THEN ions are "electrostatically attracted" to each other

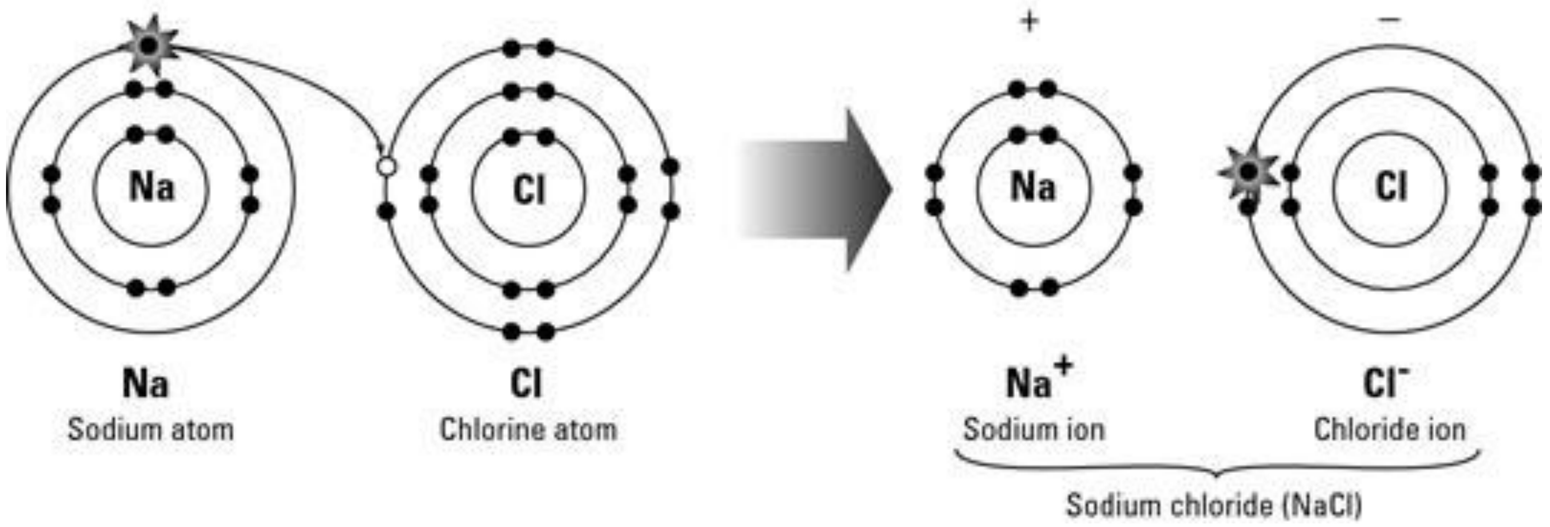
***That attraction is* the "bond"**



Ionic Bonds

- **Electronegativity differences are generally greater than 1.7 – large difference**

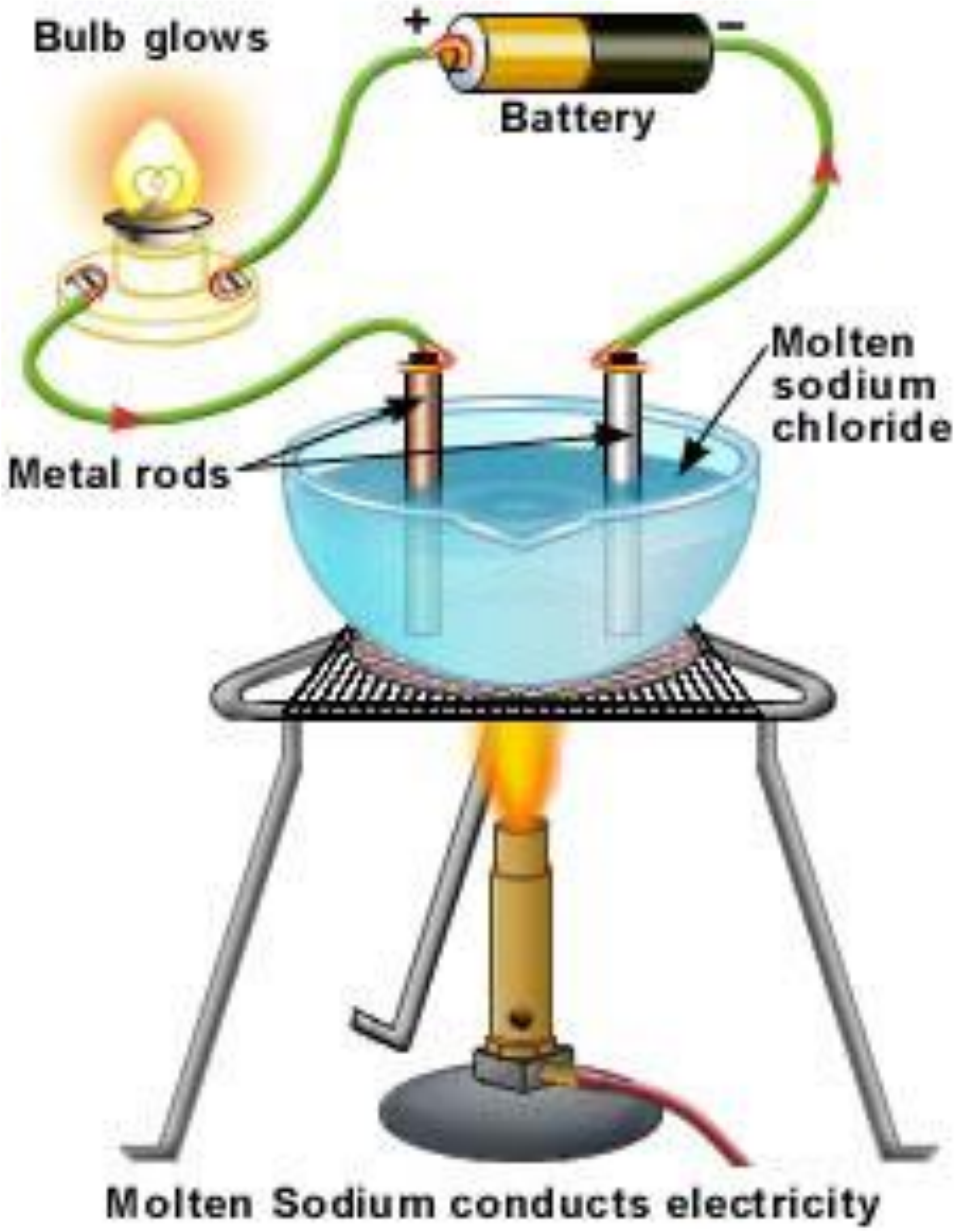
- **The last step in the formation of ionic bonds is the ions coming together - always exothermic!**



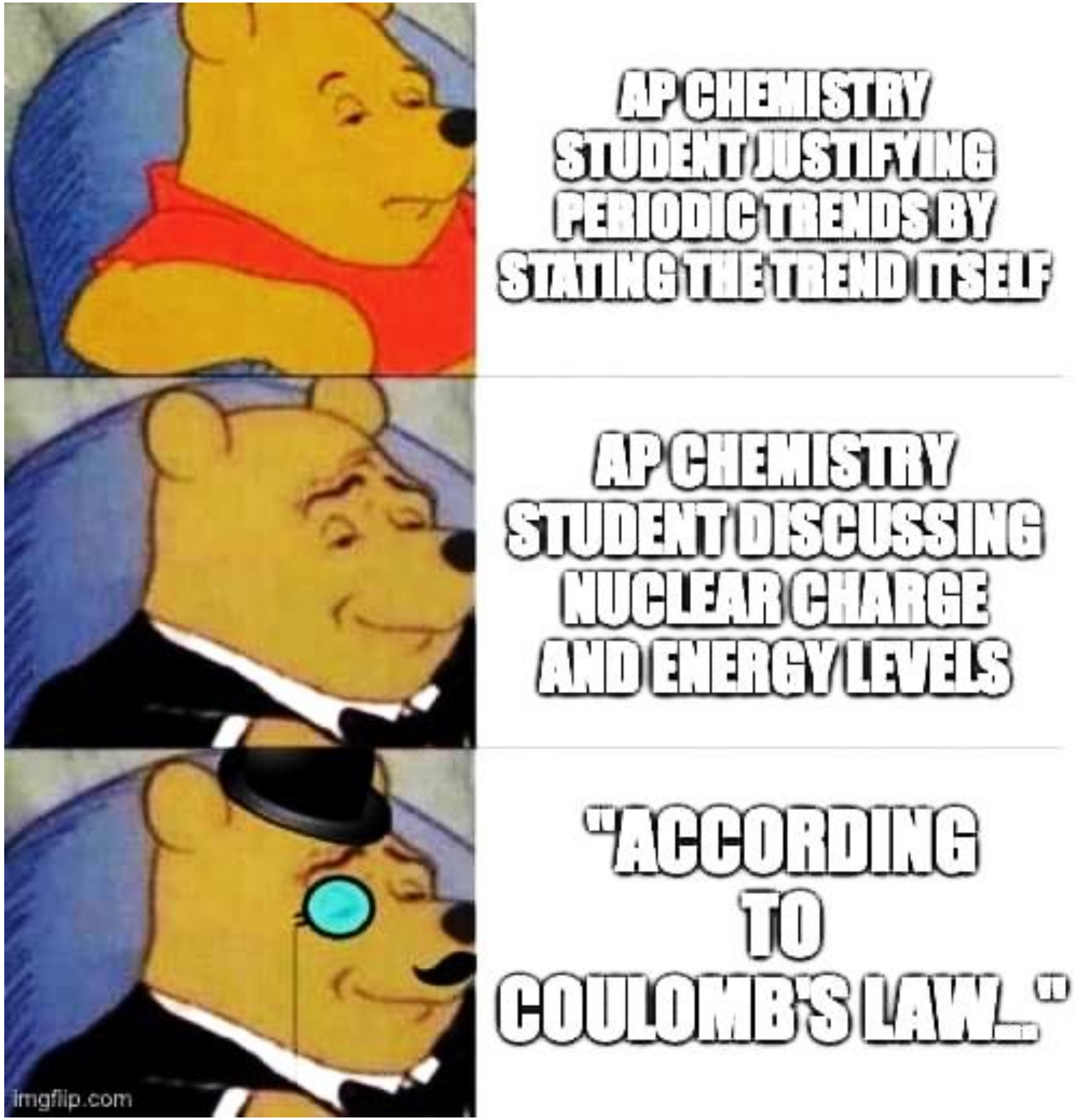
<u>Determination of Ionic Character</u>

Electronegativity difference is <u>not</u> the final determination of ionic character

Compounds are ionic if they conduct electricity in their molten state



Coulomb's Law and Periodic Trends



Coulomb's Law

Describes the attractions and repulsions between charged particles.

– Seen represented in various ways, no big deal!

$$F \propto \frac{q_1 q_2}{r^2}$$

q = absolute value of charge on particles
r = distance btwn particles

$$F = k\frac{q_1 q_2}{r^2}$$

$$E = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r^2}$$

k and the $\frac{1}{4\pi\varepsilon_0}$ are Coulomb's constant which varies based on what substance the objects are in. **k is NOT the rate constant**

Effect of Distance Between Particles (r)

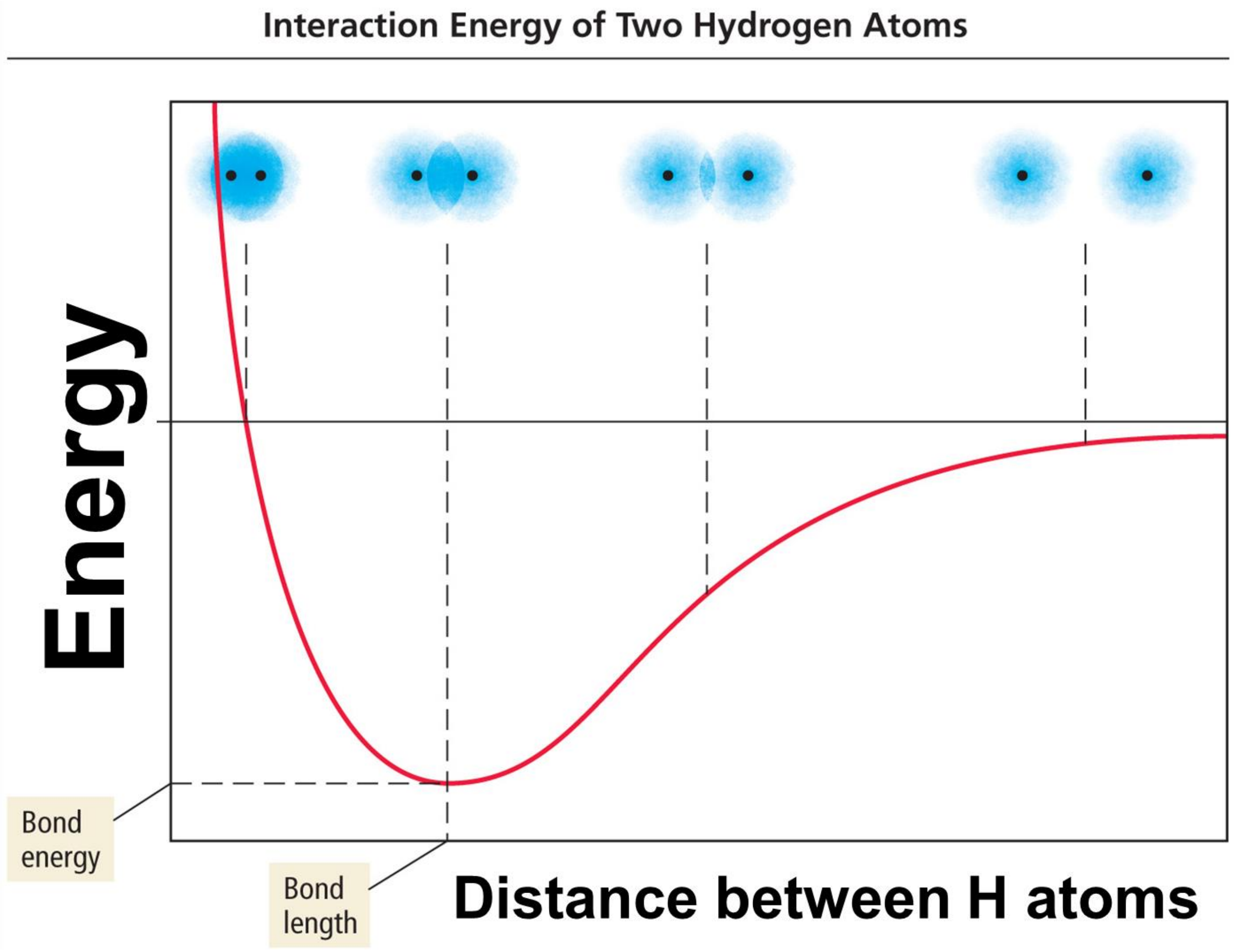
For like charges, (+ and +, or – and –)

- Like charges repel. Takes Energy to push them close.
- Potential energy (E) is positive.
- E decreases as the particles get farther apart as *r* increases.

For opposite charges, (+ and –)

- Opposite charges attract. More stable closer together.
- Potential energy is negative. (Negative is good!)
- E becomes more negative as the particles get closer together.

Effect of Distance Between Particles (r)



Effect of Charge Magnitude (q)

- **Strength of the interaction ↑ as size of the charges ↑**
  - Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge.

Therefore…

- **Strongest ionic bond would be:**
  - **Large charge magnitude** *(example: +2 versus +1, or -3 versus -2)*

    AND
  - **Small ionic radius** *(example: Li^+ versus Cs^+, or Cl^- versus I^-)*

Which factor matters more?

- ***Usually,*** **the charge magnitude is a bigger impact than the radius. Check it first!**
  - If things have the SAME charge magnitude, then check radius

How Strong is the Bond?

The more energy required to separate an ion pair (from a lattice) into ions the stronger the bond.

Usually simplified into "a modified form of Coulomb's Law" with r instead of r^2

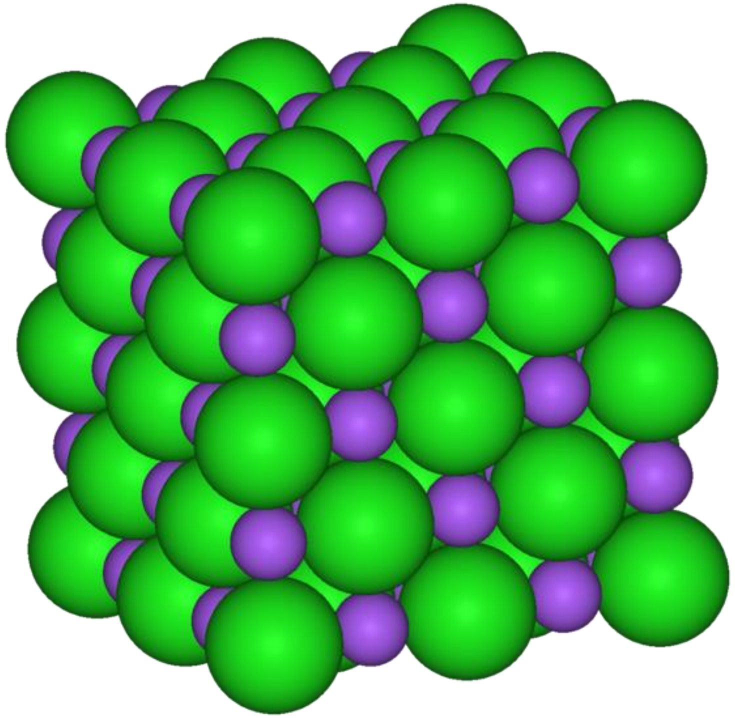
$$Lattice\ Energy = k\left(\frac{Q_1 Q_2}{r}\right)$$

$$E = \Delta H_{dissociation} \propto \left(\frac{Q_1 Q_2}{r}\right)$$

Sodium Chloride Crystal Lattice

- Ionic compounds form solids at ordinary temperatures.

- Organized in a characteristic crystal lattice of alternating positive and negative ions.

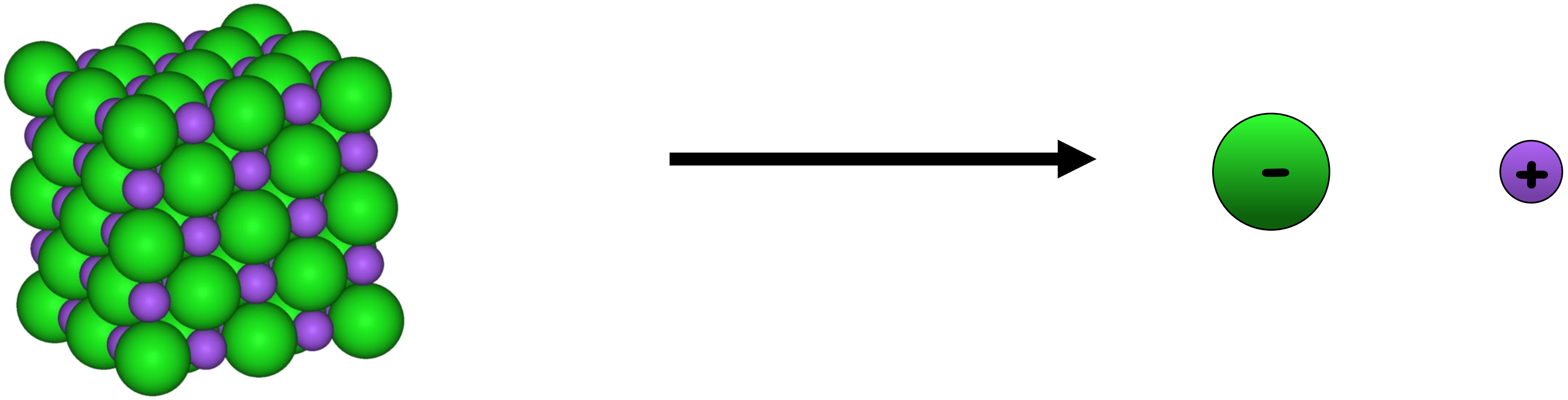


Lattice Dissociation Energy

The amount of energy absorbed to separate a *mole* of solid ionic compound into its gaseous ions

- Often just called "The Lattice Energy"

ENDOTHERMIC PROCESS (+)



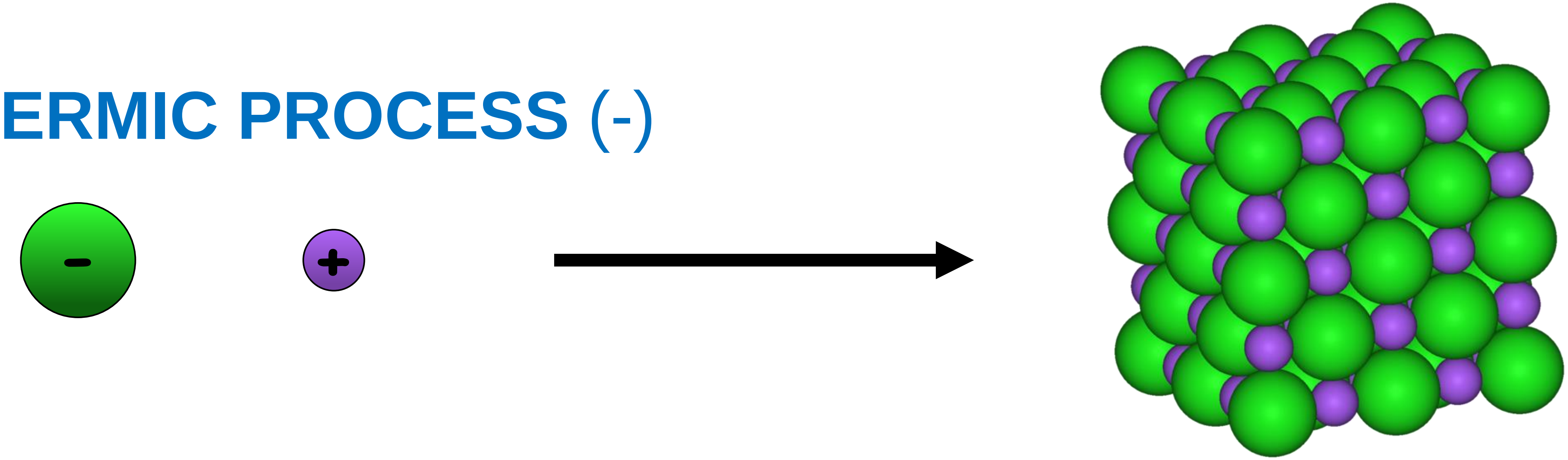
Pretend there is a mole of the solid here ☺

<u>Lattice Formation Energy</u>

The amount of energy released to form a *<u>mole</u>* of solid ionic compound from its gaseous ions

- Often just called “The Lattice Energy”

EXOTHERMIC PROCESS (-)



Pretend there is a mole of the solid here ☺

Lattice Dissociation Energy

(b) The energy required to separate the ions in the $Mg(OH)_2$ crystal lattice into individual $Mg^{2+}(g)$ and $OH^{-}(g)$ ions, as represented in the table below, is known as the lattice energy of $Mg(OH)_2(s)$. As shown in the table, the lattice energy of $Sr(OH)_2(s)$ is less than the lattice energy of $Mg(OH)_2(s)$. Explain why in terms of periodic properties and Coulomb's law.

Reaction	Lattice Energy (kJ/mol)
$Mg(OH)_2(s) \rightarrow Mg^{2+}(g) + 2\ OH^{-}(g)$	2900
$Sr(OH)_2(s) \rightarrow Sr^{2+}(g) + 2\ OH^{-}(g)$	2300

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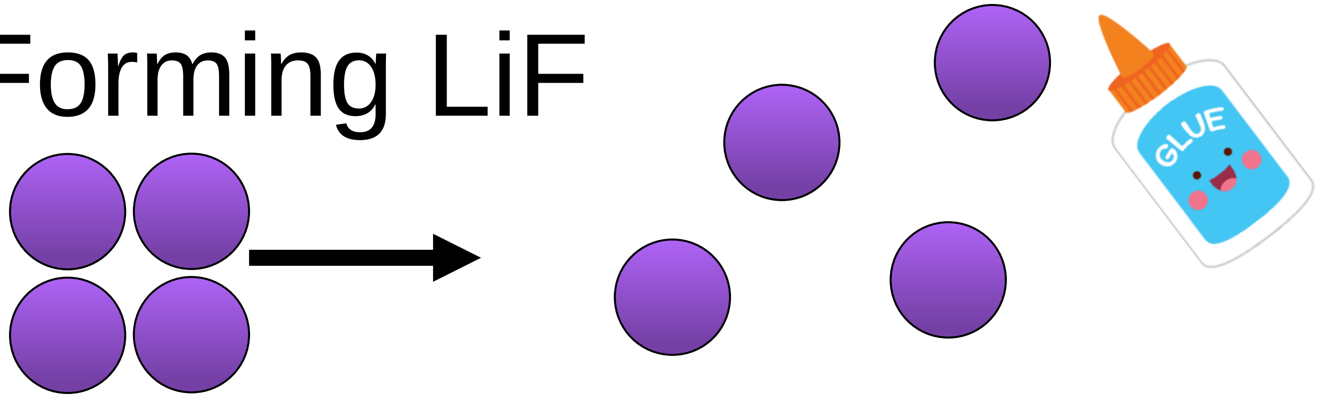
The Sr^{2+} ion is larger than the Mg^{2+} ion because it has additional occupied energy levels (or shells). Coulomb's law states that the force of attraction between cation and anion is inversely proportional to the square of the distance between them. Since the distance between Mg^{2+} and OH^- is shorter than the distance between Sr^{2+} and OH^-, the attractive forces in $Mg(OH)_2$ are stronger and, therefore, its lattice energy is greater.	1 point is earned for the correct comparison of cation sizes. 1 point is earned for indicating that smaller interionic distances lead to a greater lattice energy.

Lattice Energy vs. ΔH

Lattice Energy is just *one step* in an entire process. It involves the elements already being ions in a gaseous state.

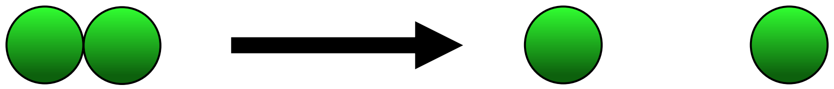
ΔH is the TOTAL amount of energy needed for the ENTIRE process to happen if the elements are NOT already in their ionic and gaseous states.

Example: Steps for Forming LiF

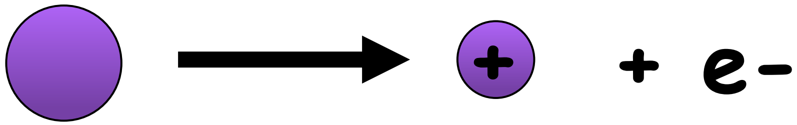


1) **Turn solid Li into a gas**
   - Sublimation

2) **Break the $F_{2\,(g)}$ bond to get $F_{(g)}$**
   - Bond energy

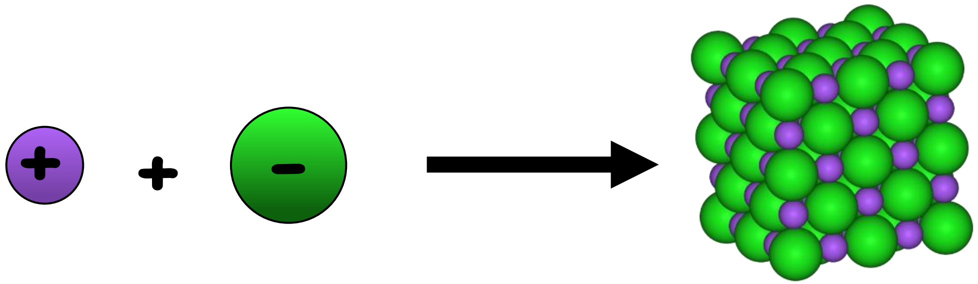


3) **Ionize Li → Li^+**
   - Ionization energy



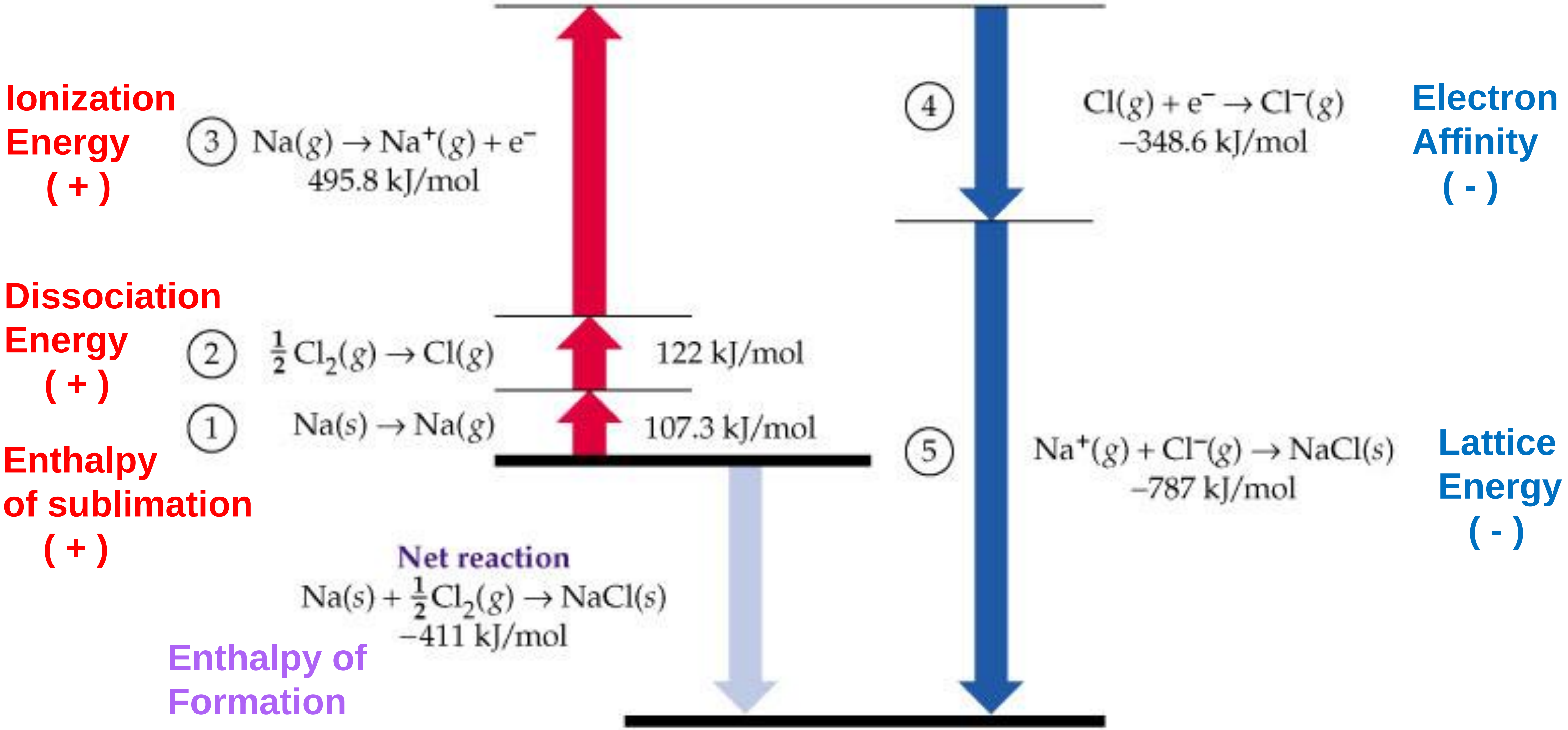
4) **Add an electron to F → F^-**
   - Electron affinity

5) **Form the ionic bond**
   - Lattice energy



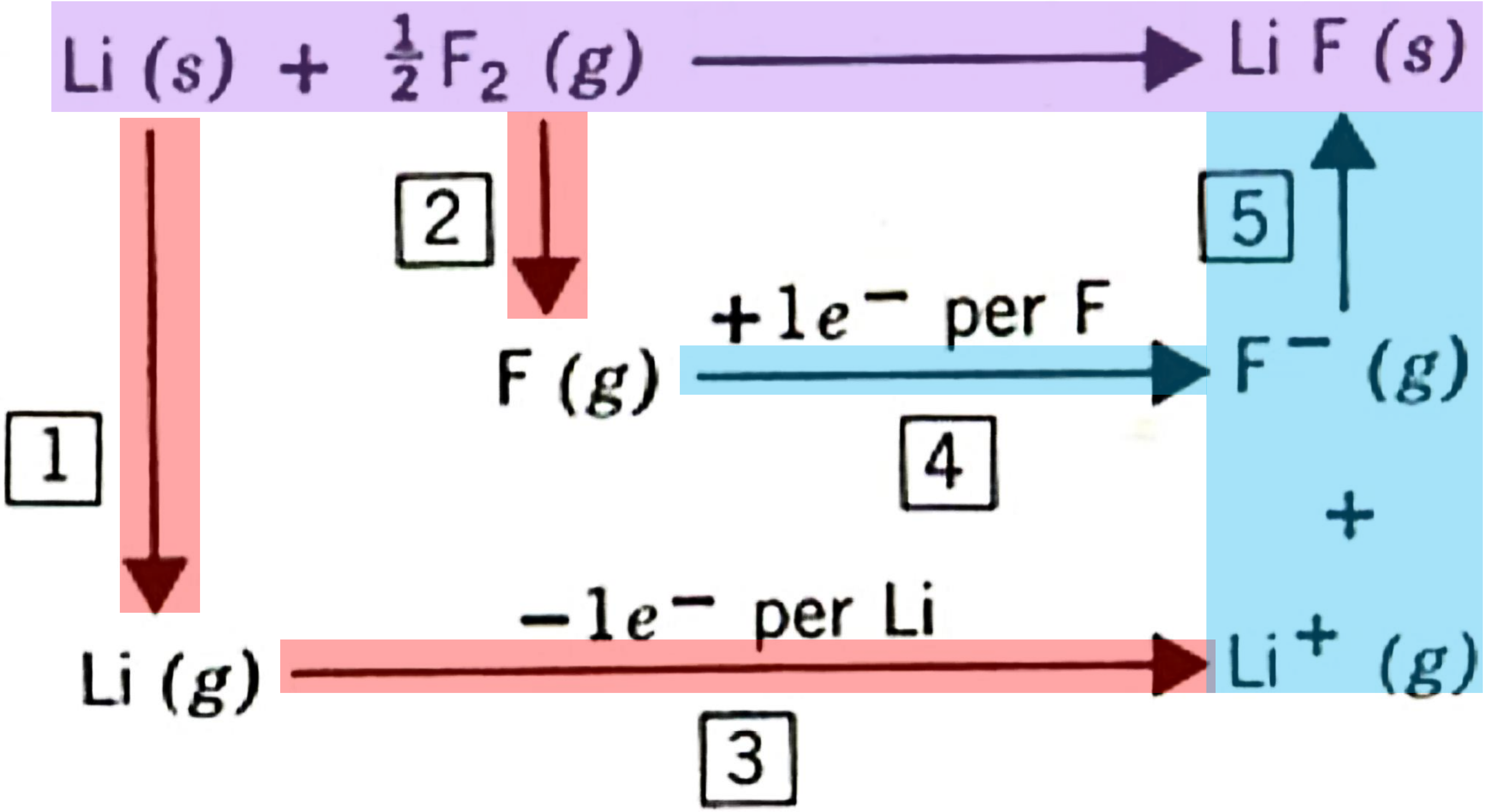
Pretend there is a mole of the solid here ☺

Often see diagrams similar to this



Often see diagrams similar to this

The overall enthalpy of formation can be + or –
depending on exact numbers for the rest of the steps!



ΔH_f for Sodium Chloride

$$Na(s) + \frac{1}{2} Cl_2(g) \rightarrow NaCl(s)$$

Sodium needs to:

1. Sublimate (s → g)
2. Loose e- ($Na \rightarrow Na^+$)

Chlorine needs to:

1. Break apart ($Cl_2 \rightarrow Cl$)
2. Gain e- ($Cl \rightarrow Cl^-$)

Ions need to:

1. Form Lattice ($Na^+ + Cl^- \rightarrow NaCl$)

Remember Hess's Law!

Lattice Energy	-786 kJ/mol
Ionization Energy for Na	495 kJ/mol
Electron Affinity for Cl	-349 kJ/mol
Bond energy of Cl_2	239 kJ/mol
Enthalpy of sublimation for Na	109 kJ/mol

$$Na(s) \rightarrow \cancel{Na}(g) \qquad + 109\ kJ$$

$$\cancel{Na}(g) \rightarrow \cancel{Na^+}(g) + \cancel{e^-} \qquad + 495\ kJ$$

$$\frac{1}{2} Cl_2(g) \rightarrow \cancel{Cl}(g) \qquad + \frac{1}{2}(239\ kJ)$$

$$\cancel{Cl}(g) + \cancel{e^-} \rightarrow \cancel{Cl^-}(g) \qquad - 349\ kJ$$

$$\cancel{Na^+}(g) + \cancel{Cl^-}(g) \rightarrow NaCl(s) \qquad -786\ kJ$$

$$Na(s) + \frac{1}{2} Cl_2(g) \rightarrow NaCl(s) \qquad -412\ kJ/mol$$

YouTube Link to Presentation:

https://youtu.be/bJm5LGLngfo