

# **N2O – Bonding**

## **Energy of Bonding**

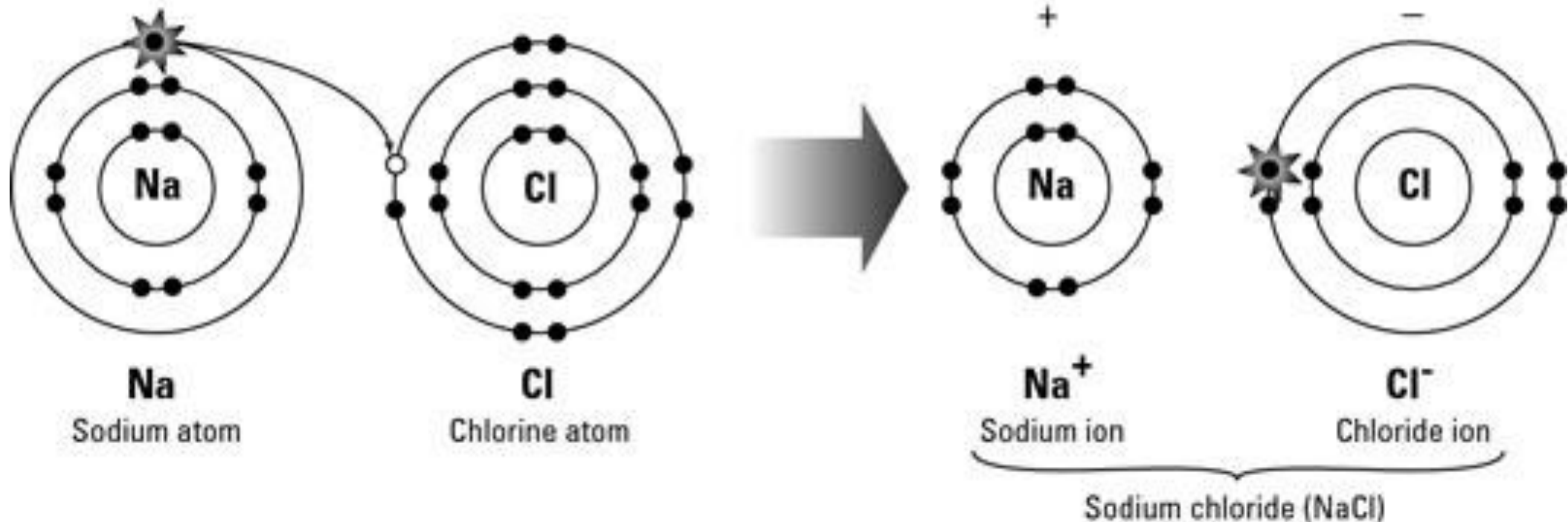
# Electronegativity

**The ability of an atom in a molecule to attract shared electrons to itself.**

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

# Ionic Bonds

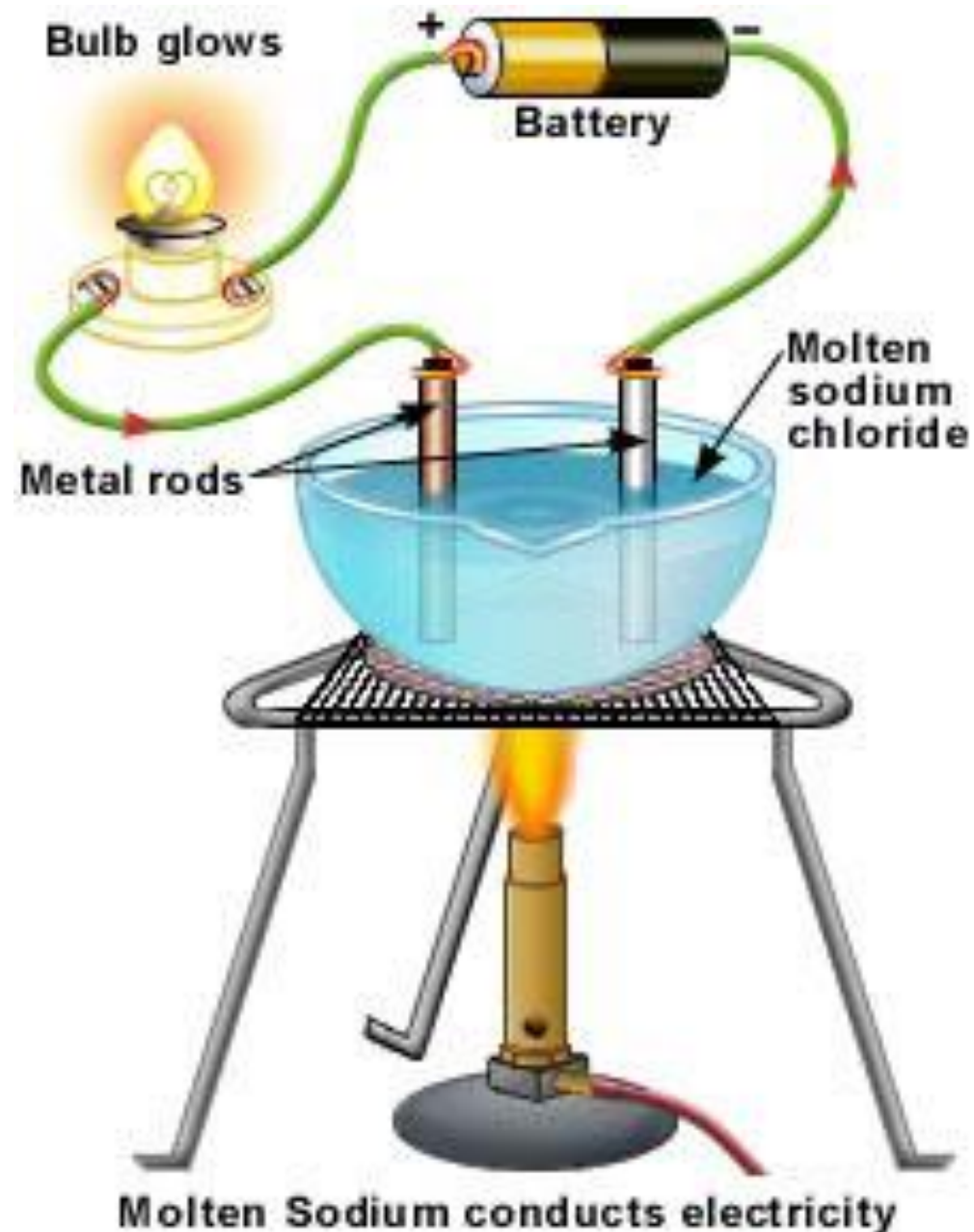
- Electrons are transferred
- Electronegativity differences are generally greater than 1.7 – large difference
- The formation of ionic bonds is always exothermic!



# Determination of Ionic Character

Electronegativity difference is not the final determination of ionic character

Compounds are ionic if they conduct electricity in their molten state



# 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

$$F = k \frac{q_1 q_2}{r^2} \quad E = \frac{1}{4\pi\epsilon_0} \frac{q_1 q_2}{r^2}$$

k and the  $\frac{1}{4\pi\epsilon_0}$  are Coulomb's constant  
which varies based on what substance  
the objects are in

# Effect of Distance Between Particles

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

- Remember, 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 – )**

- Remember, like 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 Charge

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

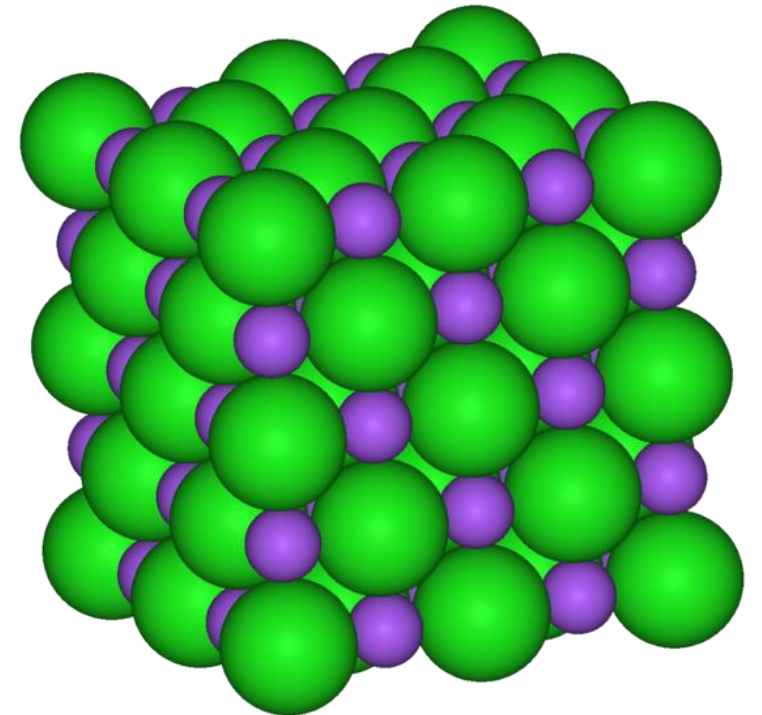
## Therefore...

- **Strongest ionic bond would be:**
  - Large charge magnitude (*example: +2 versus +1, or -3 versus -2*)
  - AND
  - Small ionic radius (*example:  $\text{Li}^+$  versus  $\text{Cs}^+$ , or  $\text{Cl}^-$  versus  $\text{I}^-$* )



# Sodium Chloride Crystal Lattice

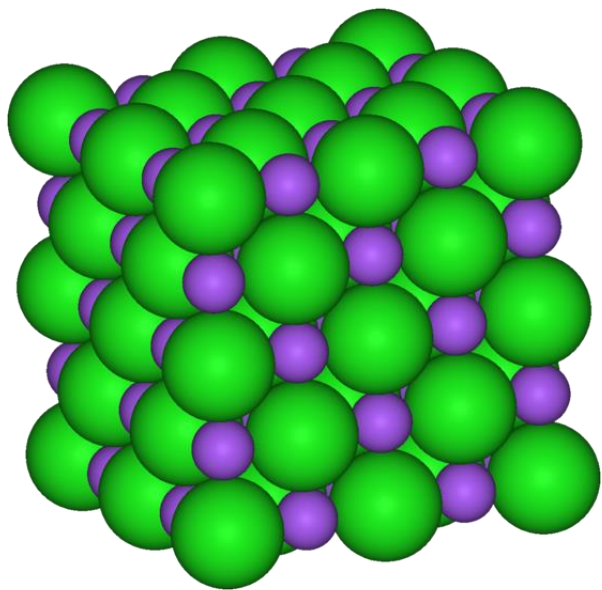
- Ionic compounds form solids at ordinary temperatures.
- Ionic compounds organize in a characteristic crystal lattice of alternating positive and negative ions.





# Lattice Dissociation Energy

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

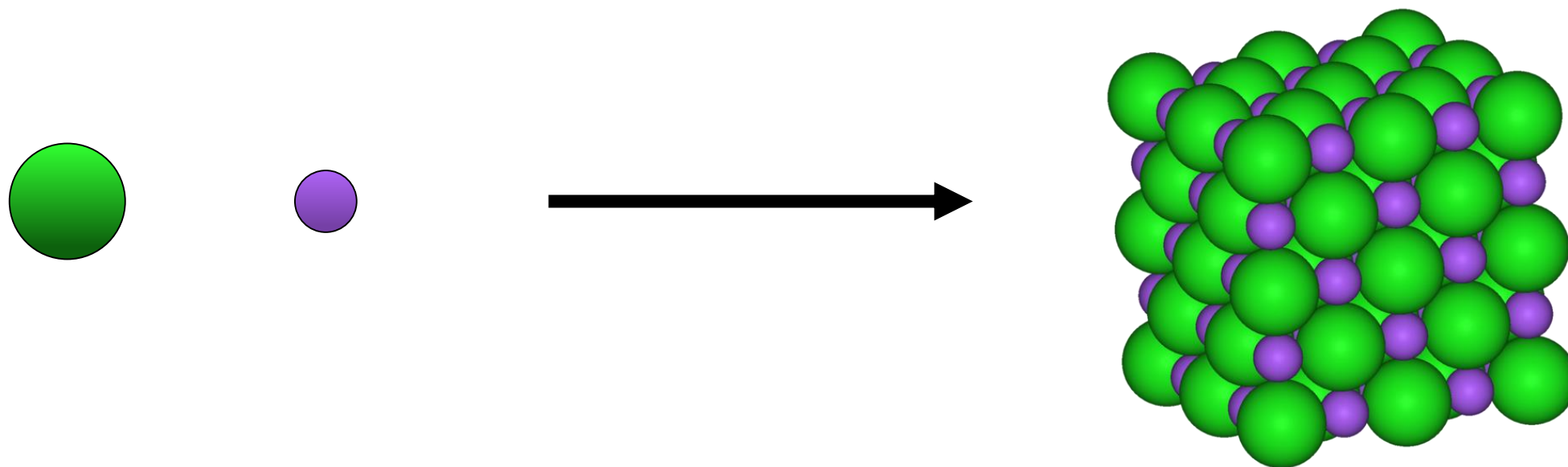


*Pretend there are a mole  
of each element here 😊*

# Lattice Formation Energy

**The amount of energy involved to form a mole of solid ionic compound from its gaseous ions**

- Usually just called “The Lattice Energy”

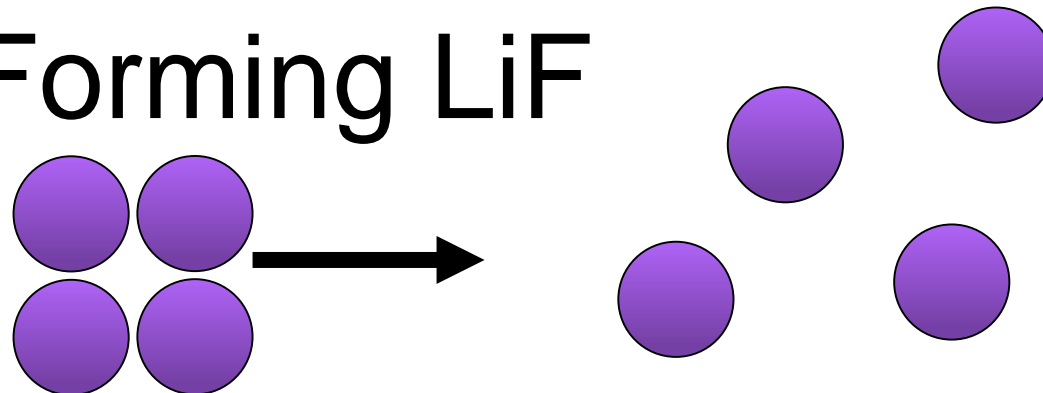


*Pretend there are a mole  
of each element here 😊*

# Example: Steps for Forming LiF

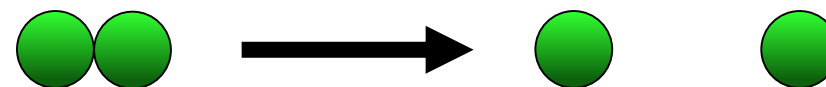
1) Turn solid Li into a gas

- Sublimation



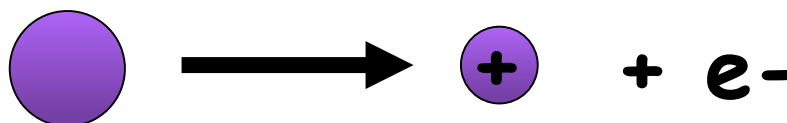
2) Break the  $\text{F}_{2(g)}$  bond to get  $\text{F}_{(g)}$

- Bond energy



3) Ionize  $\text{Li} \rightarrow \text{Li}^+$

- Ionization energy



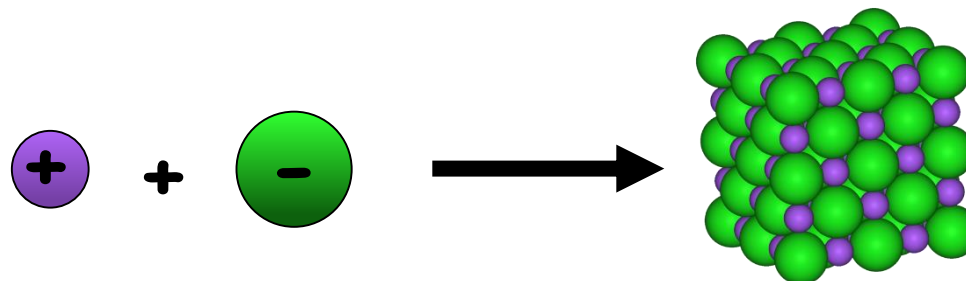
4) Add an electron to  $\text{F} \rightarrow \text{F}^-$

- Electron affinity



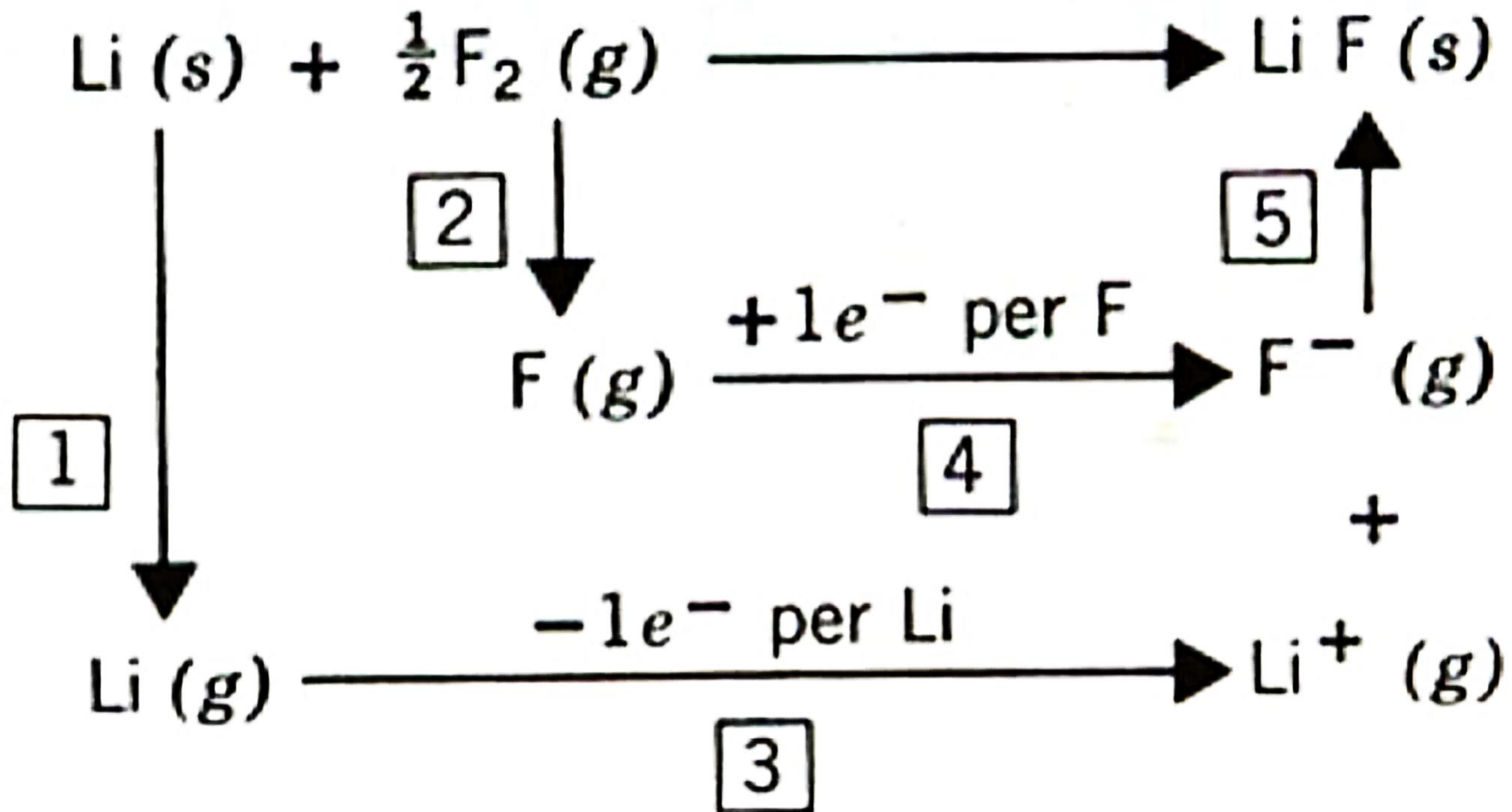
5) Form the ionic bond

- Lattice energy



*Pretend there are a mole of each element here ☺*

# Often see diagrams similar to this



# Lattice Energy cont...

Often see Lattice energy simplified into “a modified form of Coulomb’s Law” with  $r$  instead of  $r^2$

- $k$  is a proportionality constant - depends on structure of the solid and the electron configurations of the ions.

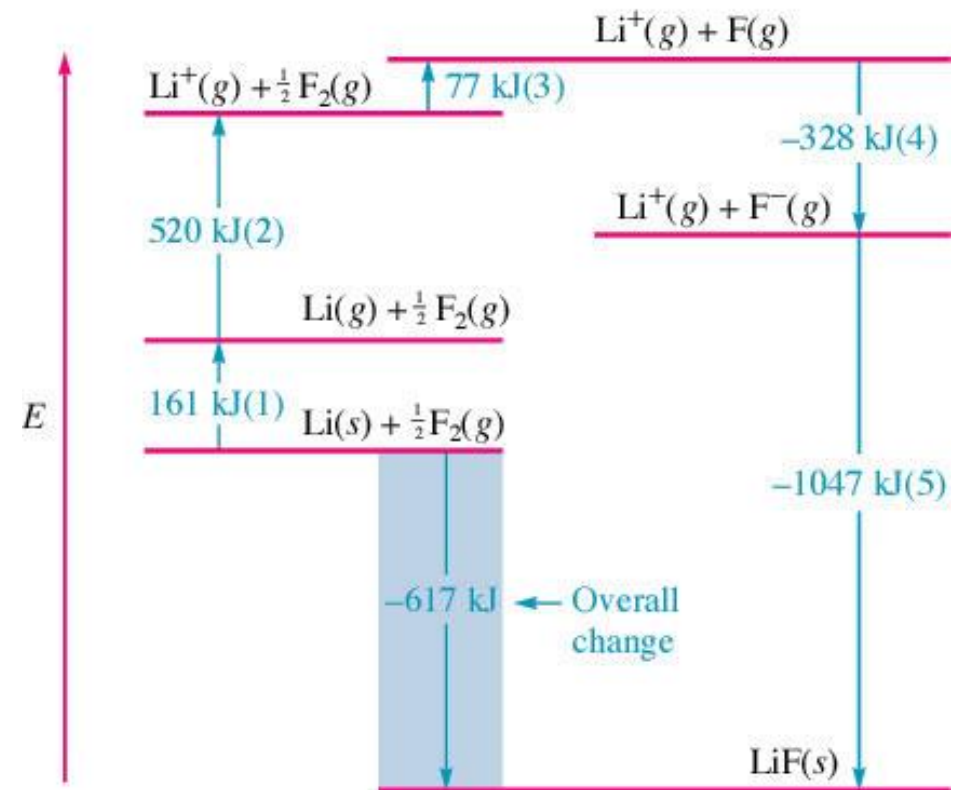
*$k$  is not the rate constant*

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

# How Strong is the Bond?

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

- Often use the “Enthalpy of Dissociation” to discuss this bond strength
- Equal but opposite sign as the Lattice Formation Energy.

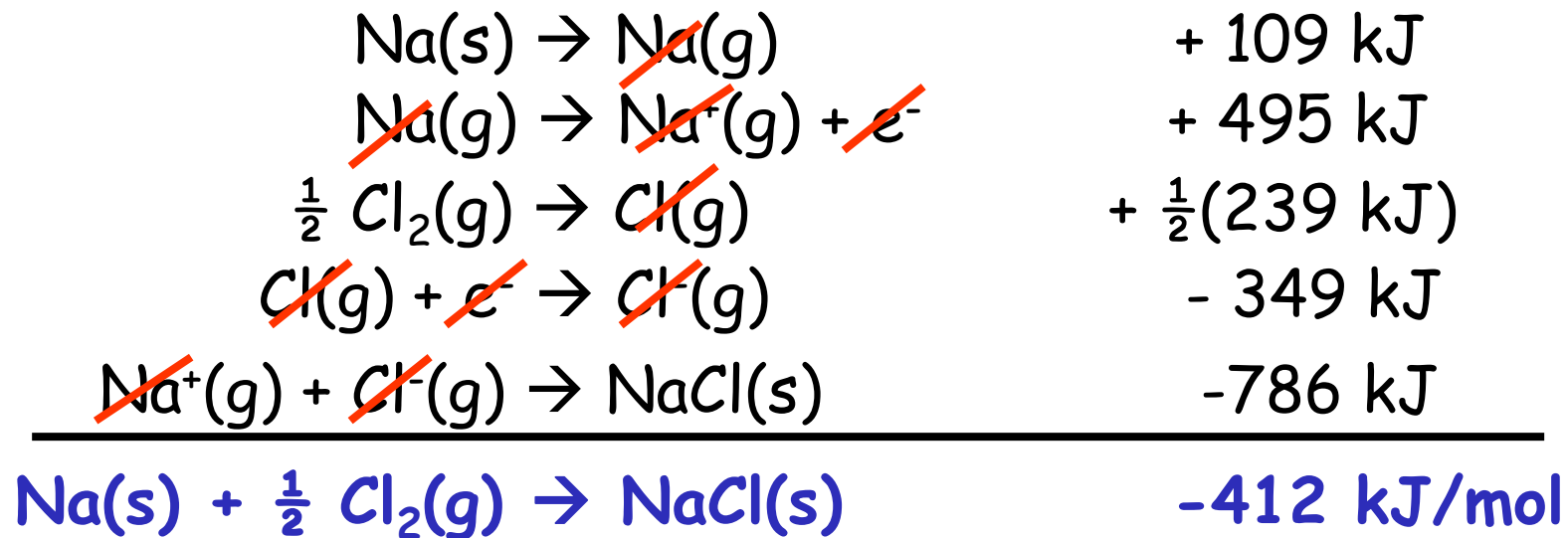


$$E = \Delta H_{\text{dissociation}} \propto \left( \frac{Q_1 Q_2}{r} \right)$$

# Estimate $\Delta H_f$ for Sodium Chloride



Lattice Energy	-786 kJ/mol
Ionization Energy for Na	495 kJ/mol
Electron Affinity for Cl	-349 kJ/mol
Bond energy of $\text{Cl}_2$	239 kJ/mol
Enthalpy of sublimation for Na	109 kJ/mol





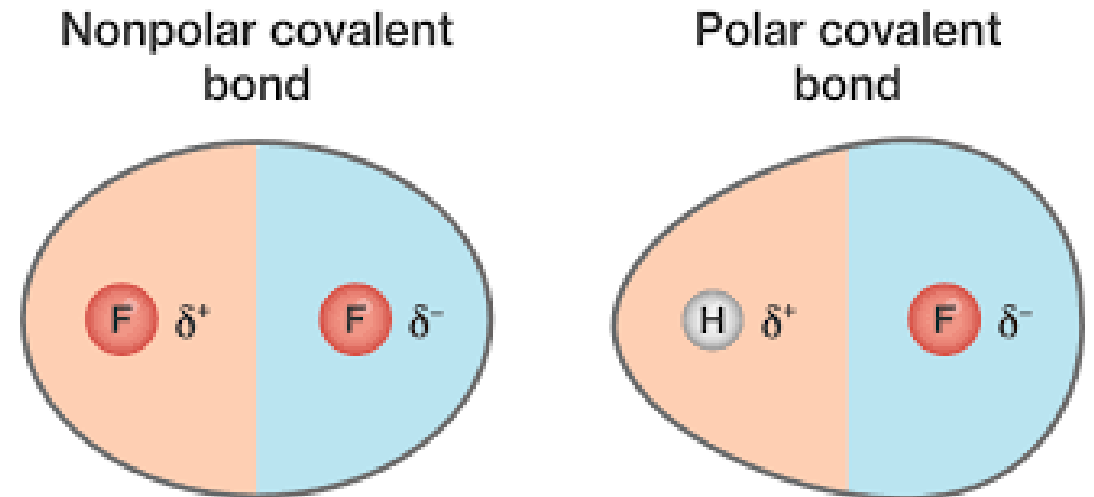
# Covalent Bonds

## Polar-Covalent bonds

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

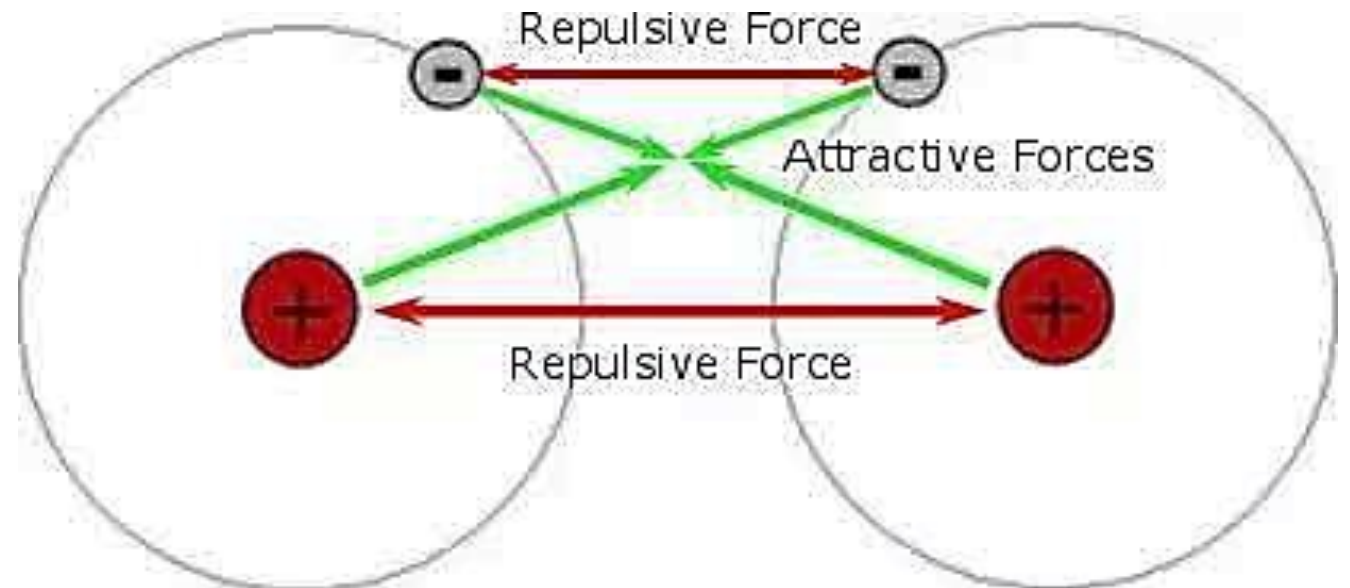
## Nonpolar-Covalent bonds

- 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” ?

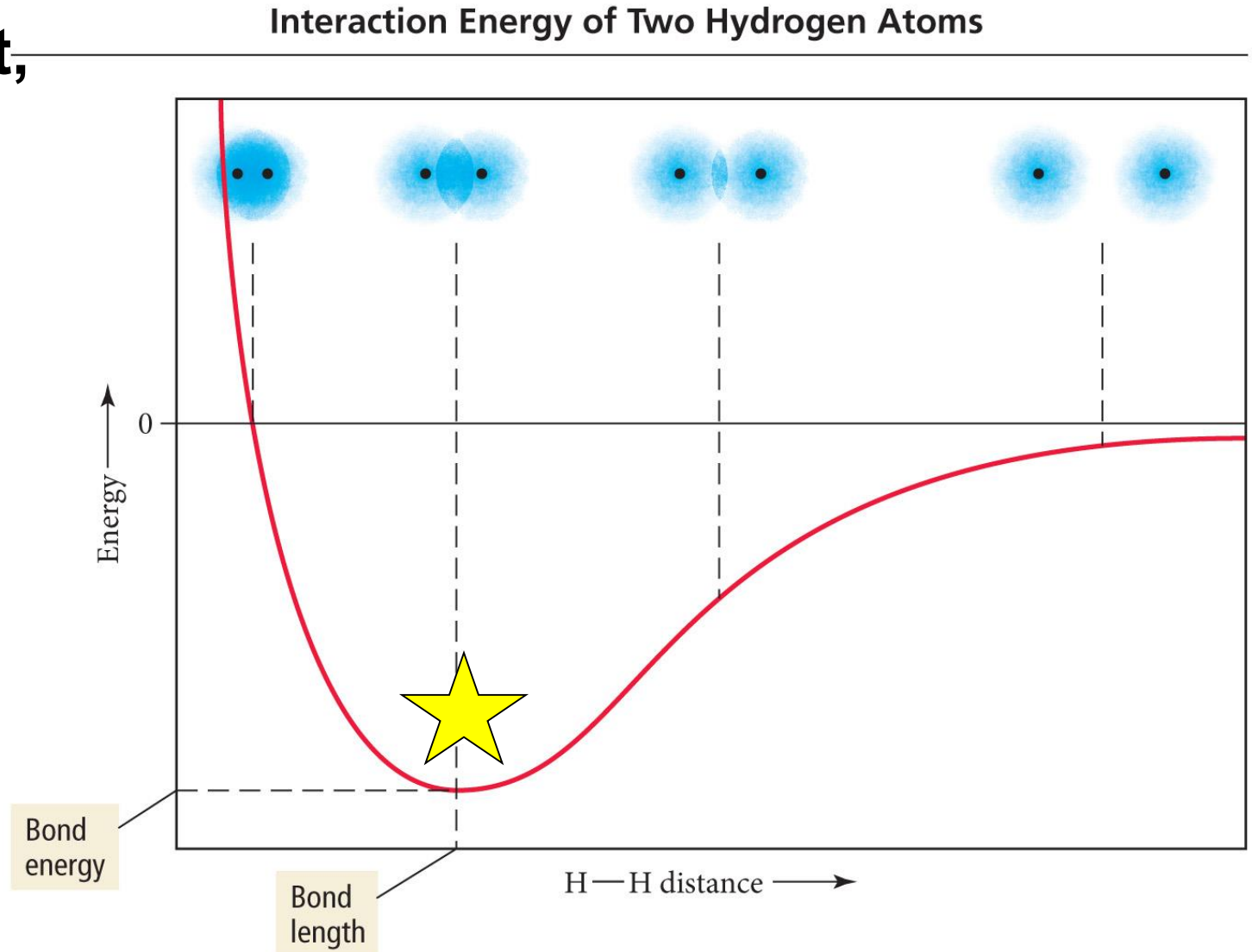
“Bonded” when at lowest,  
most stable energy.

**Goldie Locks...**

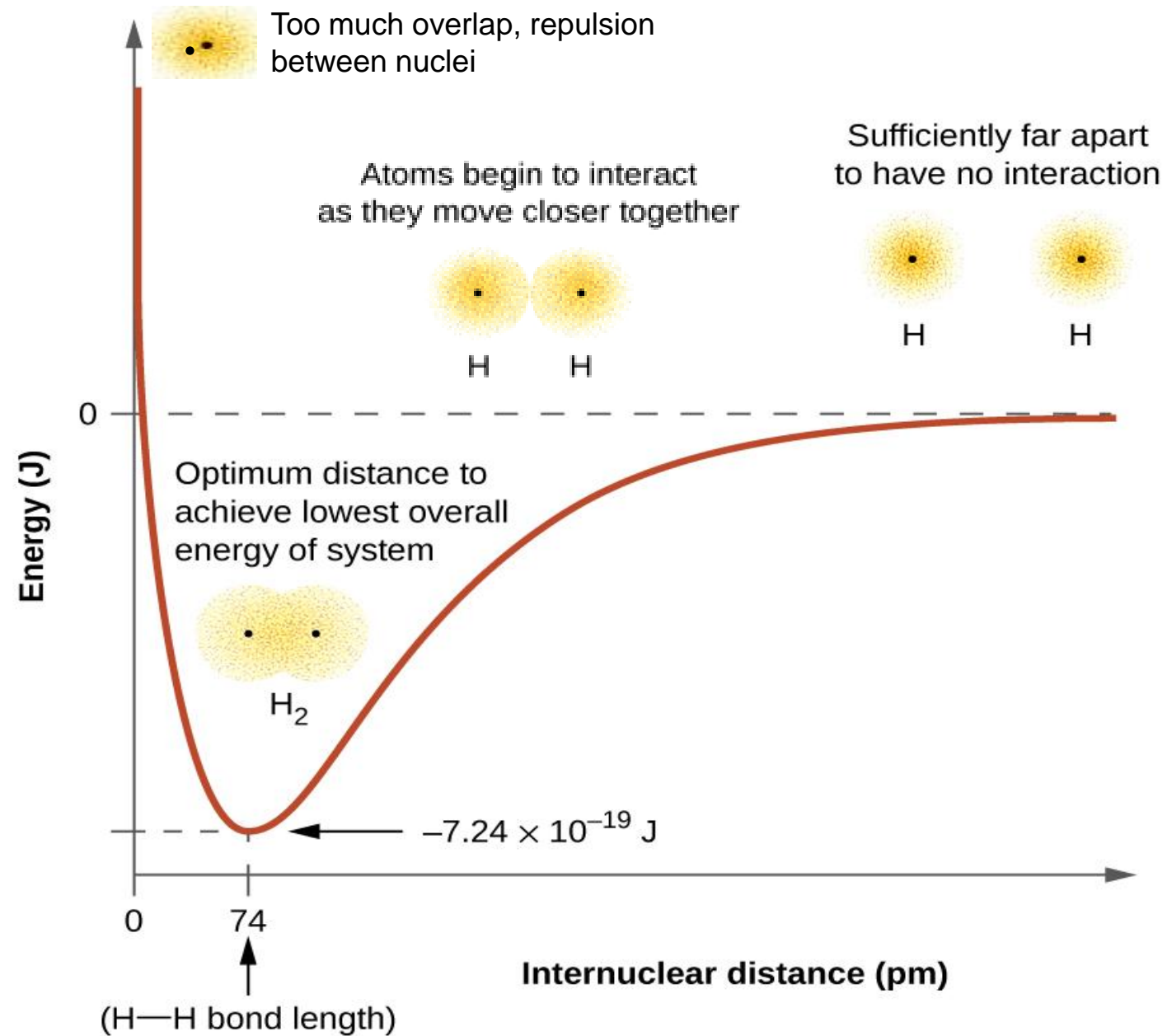
Too far = bad

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)
$\text{C} - \text{C}$	Single	154	347
$\text{C} = \text{C}$	Double	134	614
$\text{C} \equiv \text{C}$	Triple	120	839
$\text{C} - \text{O}$	Single	143	358
$\text{C} = \text{O}$	Double	123	745
$\text{C} - \text{N}$	Single	143	305
$\text{C} = \text{N}$	Double	138	615
$\text{C} \equiv \text{N}$	Triple	116	891

# Bond Energy and Enthalpy

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

Energy required

Energy released

$D$  = Bond energy per mole of bonds

**Breaking bonds always requires energy**

Breaking = endothermic

**Forming bonds always releases energy**

Forming = exothermic

# Bond Energy and Enthalpy

“Takes to Break” = + endo  
“Frees to Form” = - exo

How much energy does it take to break  $2\text{H}_2\text{O}$  into  $2\text{H}_2$  and  $\text{O}_2$ ?

Bond energies: O-H 463 kJ/mol, H-H 436 kJ/mol, O=O 498 kJ/mol

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

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

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