Bonding – General Concepts

Coulomb's Law, Lattice Energy, Bond E

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

- below 1.0
- 1.0–1.4
- 1.5–1.9
- 2.0–2.4
- 2.5–2.9
- 3.0–4.0

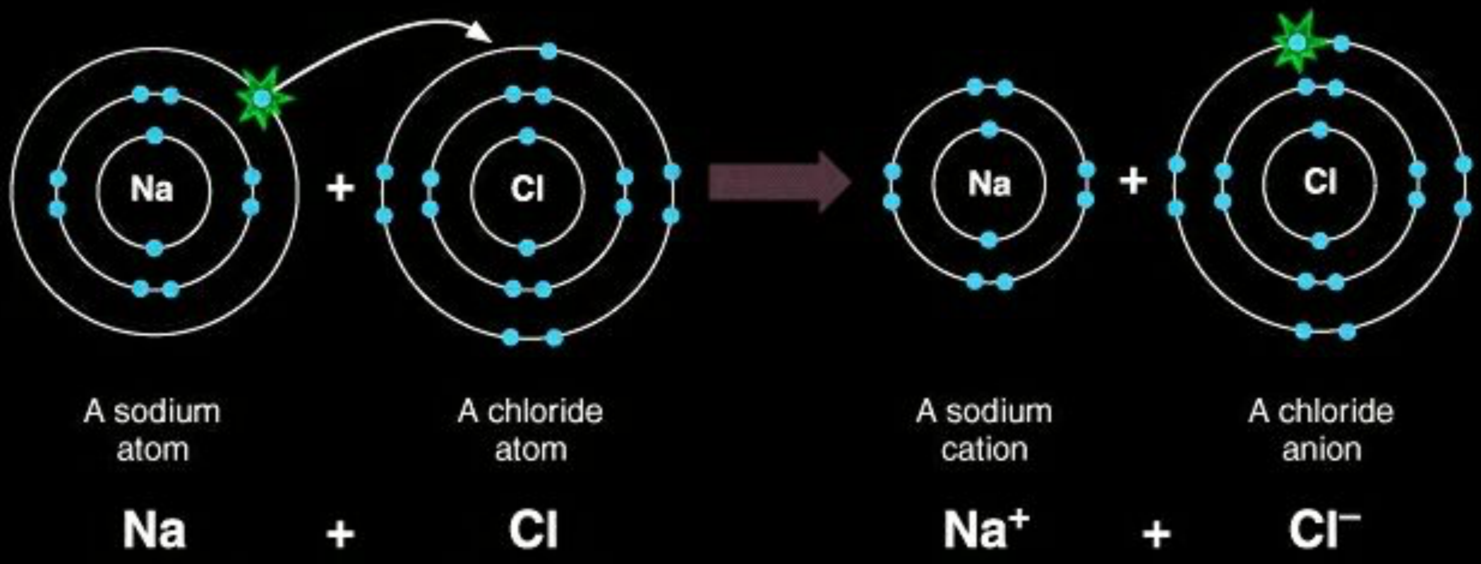
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr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

*Lanthanides: 1.1–1.3

†Actinides: 1.3–1.5

Ionic Bonds

- Electrons are transferred

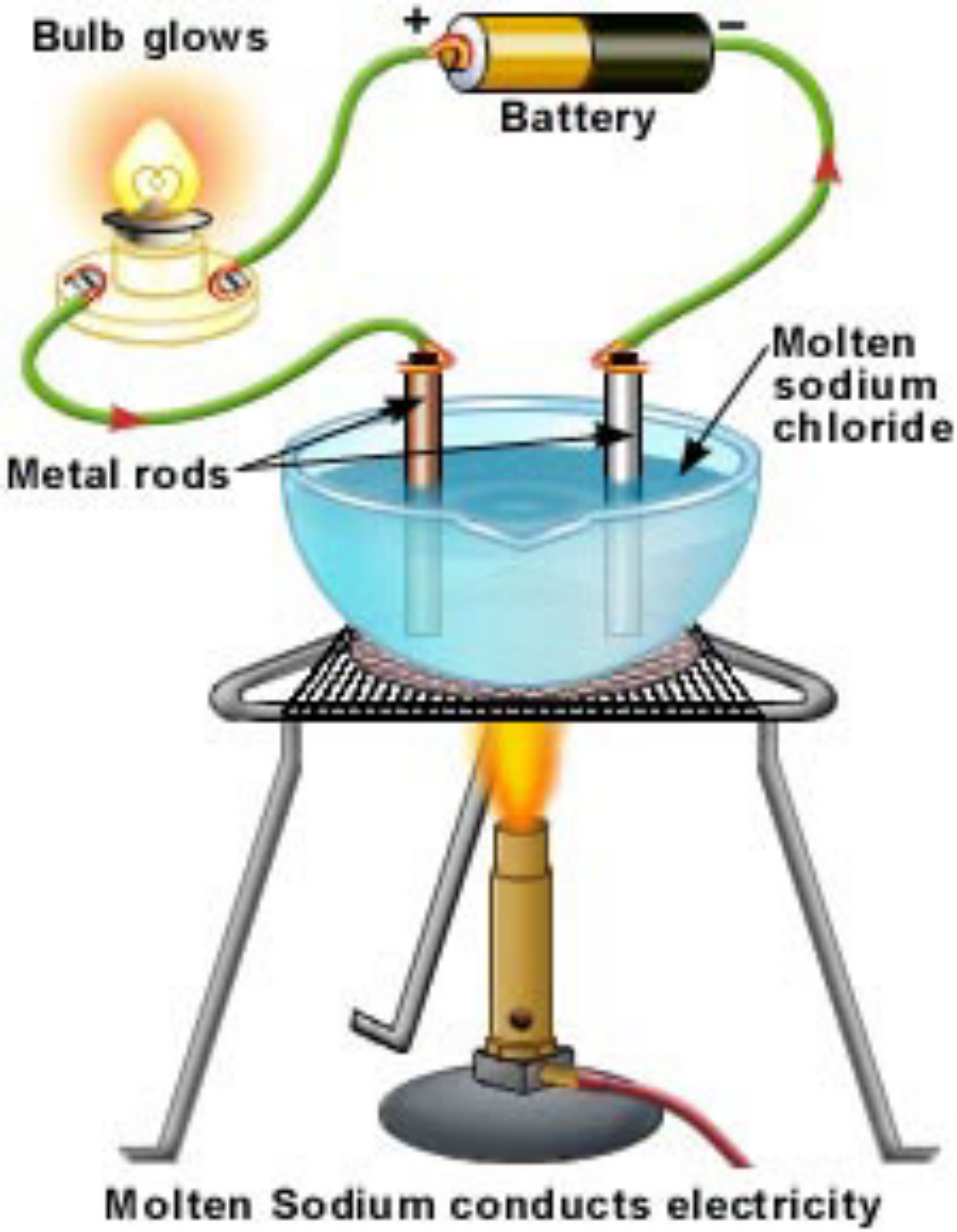


- Electronegativity differences are generally greater than 1.7
- The formation of ionic bonds is always exothermic!

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

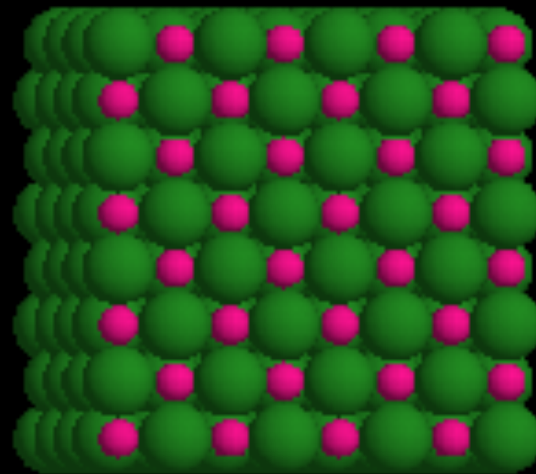


Coulomb's Law

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

- Coulomb's law describes the attractions and repulsions between charged particles.
- For like charges, the potential energy (E) is positive and decreases as the particles get farther apart as *r* increases. r should be squared
- For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.
- 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.

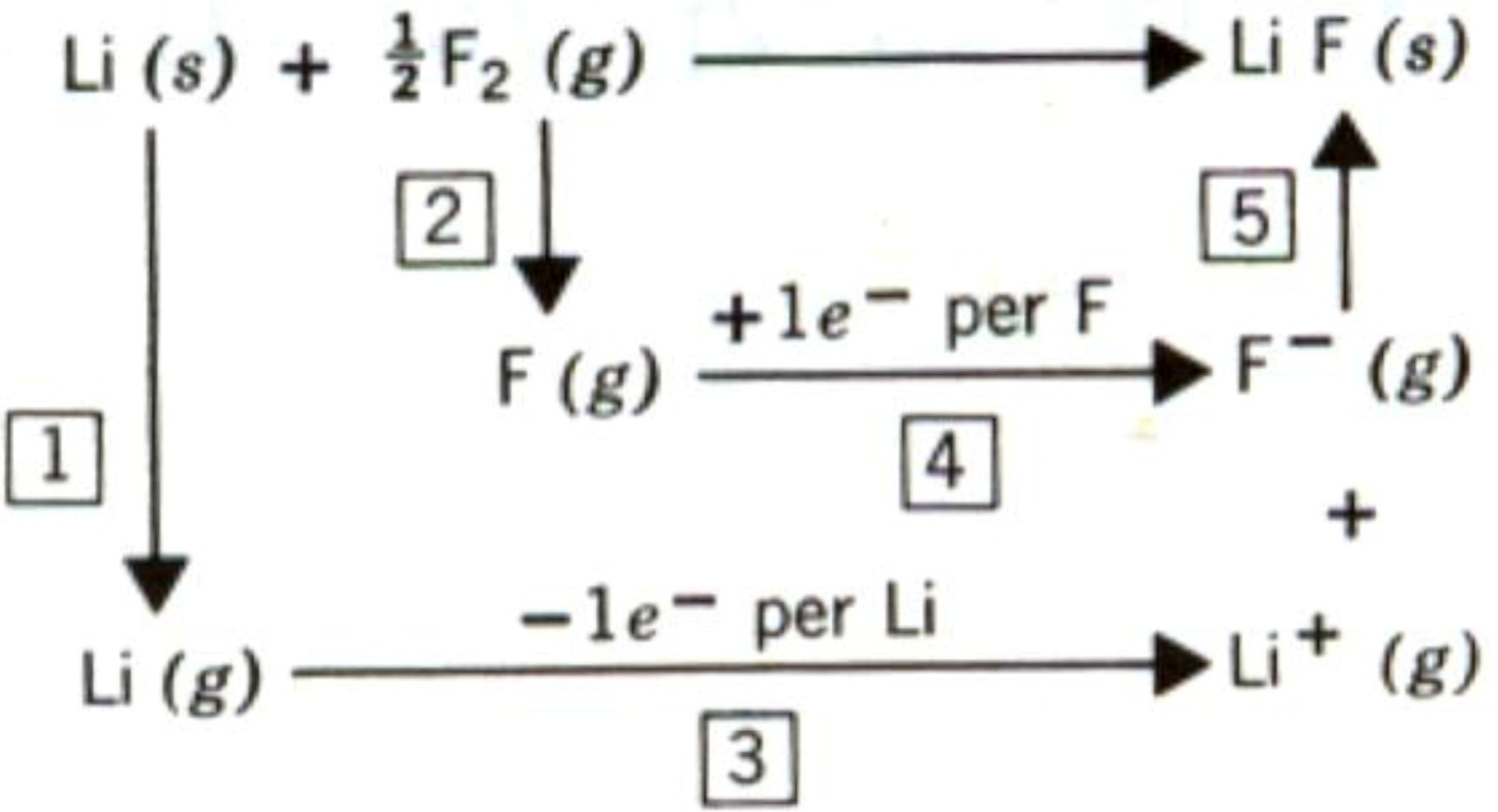
Sodium Chloride Crystal Lattice



Ionic compounds form solids at ordinary temperatures.

Ionic compounds organize in a characteristic crystal lattice of alternating positive and negative ions.

Lattice Energy Steps

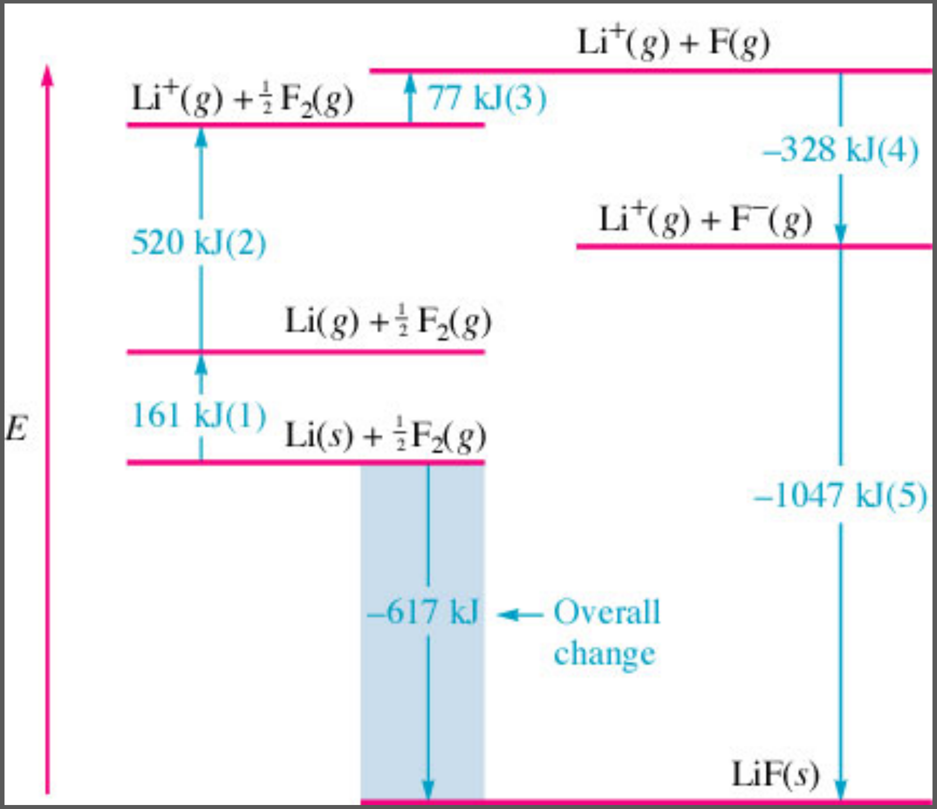


Lattice Energy cont…

Lattice energy can be represented by a modified form of Coulomb's Law: *k is a* proportionality constant that depends on the structure of the solid and the electron configurations of the ions.

k is not the rate constant from this point of view

$$LatticeEnergy = k\left(\frac{Q_1Q_2}{r}\right)$$



Enthalpy of dissociation

--energy required to decompose an ion pair (from a lattice) into ions ∴ a measure of the strength of the ionic bond

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

Reminder: Q are the ion charges

Estimate ΔH_f for Sodium Chloride

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

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

$Na(s) \rightarrow \cancel{Na}(g)$	+ 109 kJ
$\cancel{Na}(g) \rightarrow \cancel{Na^+}(g) + \cancel{e^-}$	+ 495 kJ
$\frac{1}{2}\ Cl_2(g) \rightarrow \cancel{Cl}(g)$	+ $\frac{1}{2}$(239 kJ)
$\cancel{Cl}(g) + \cancel{e^-} \rightarrow \cancel{Cl^-}(g)$	- 349 kJ
$\cancel{Na^+}(g) + \cancel{Cl^-}(g) \rightarrow NaCl(s)$	-786 kJ
$Na(s) + \frac{1}{2}\ Cl_2(g) \rightarrow NaCl(s)$	**-412 kJ/mol**

Covalent Bonds

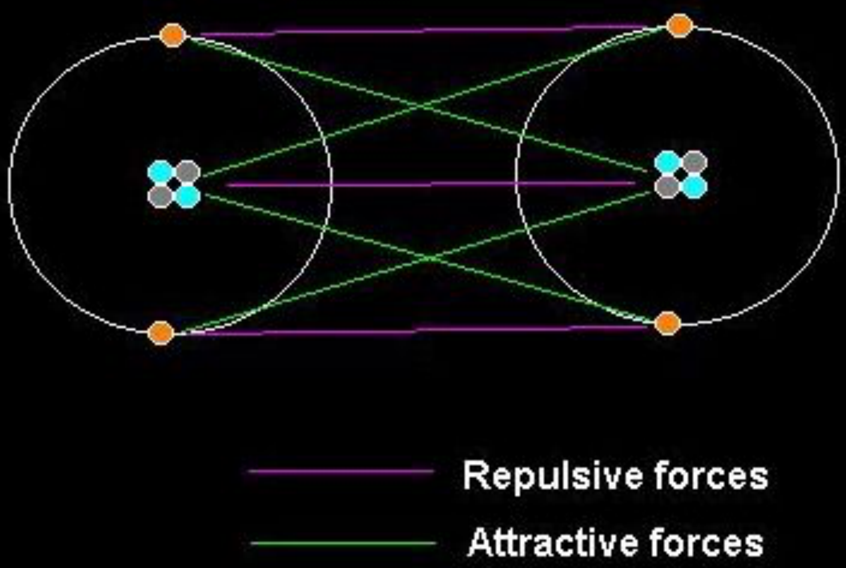
Polar-Covalent bonds

- Electrons are unequally shared
- Electronegativity difference between .3 and 1.7

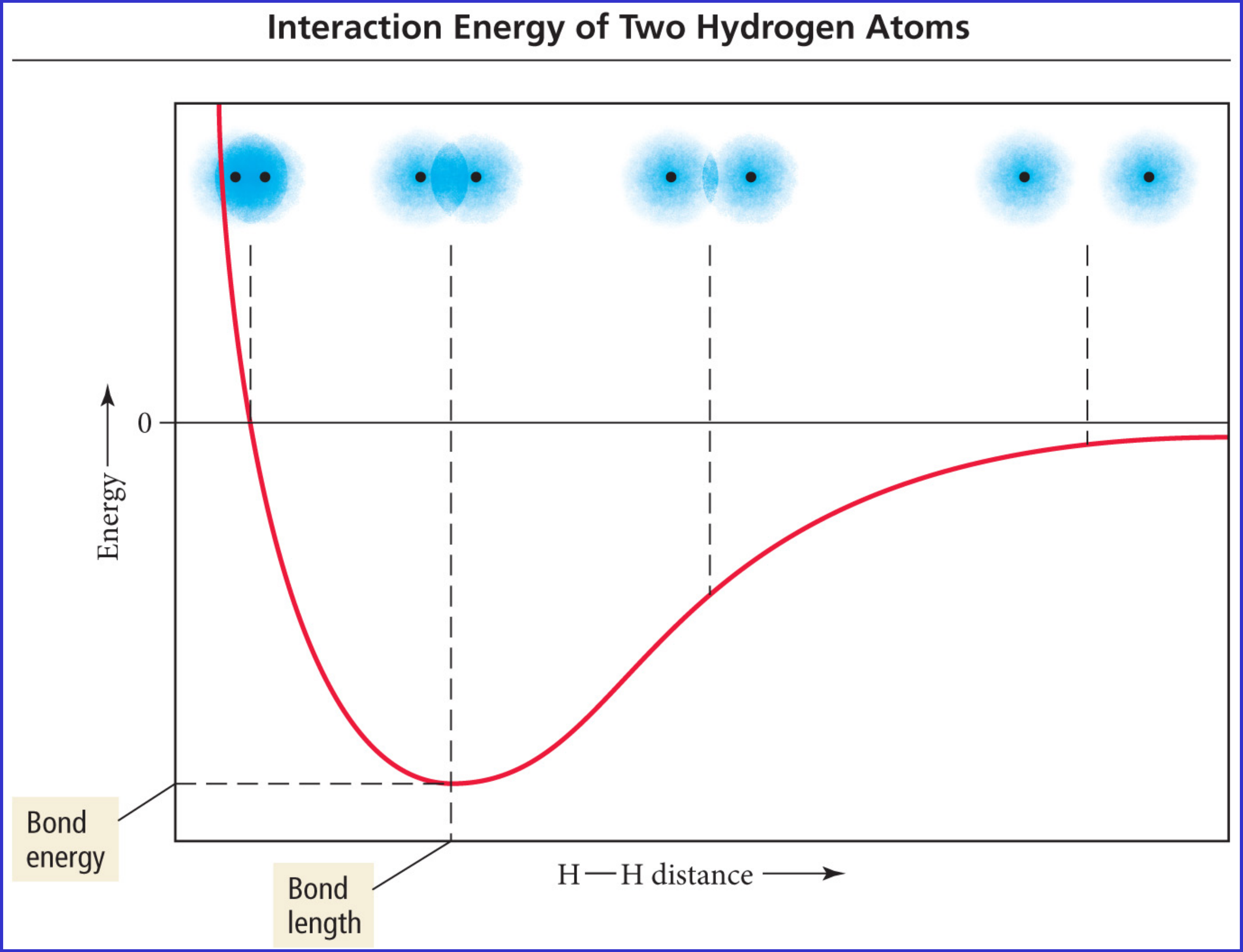
Nonpolar-Covalent bonds

- Electrons are equally shared
- Electronegativity difference of 0 to 0.3

Covalent Bonding Forces

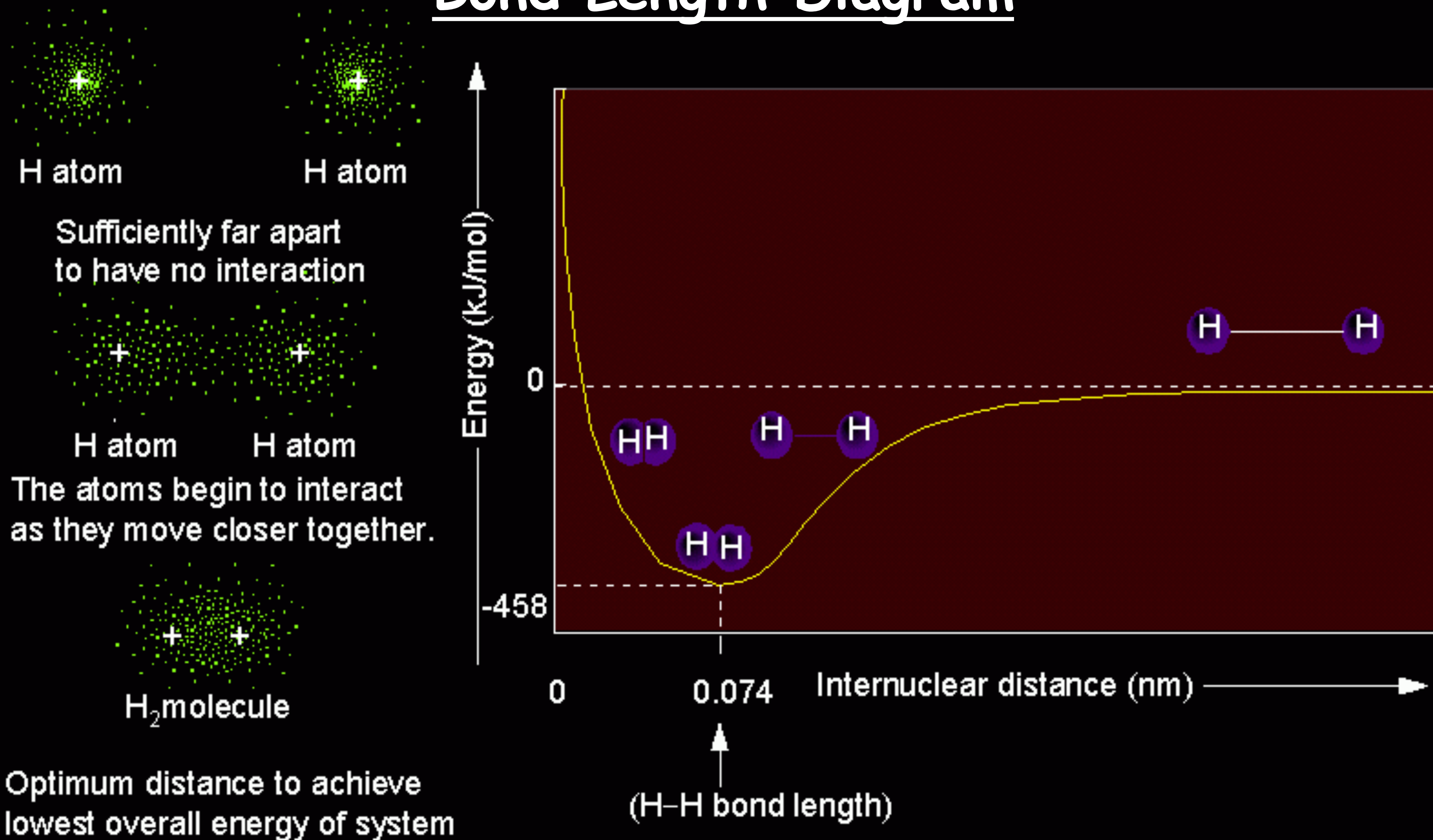


- Electron – electron repulsive forces = Bad
- Proton – proton repulsive forces = bad
- Electron – proton attractive forces = good



Interaction Energy of Two Hydrogen Atoms
Energy
0
Bond energy
Bond length
H—H distance

Bond Length Diagram



Bond Length and Energy

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

Bonds between elements become shorter and stronger as multiplicity increases.

Bond Energy and Enthalpy

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

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