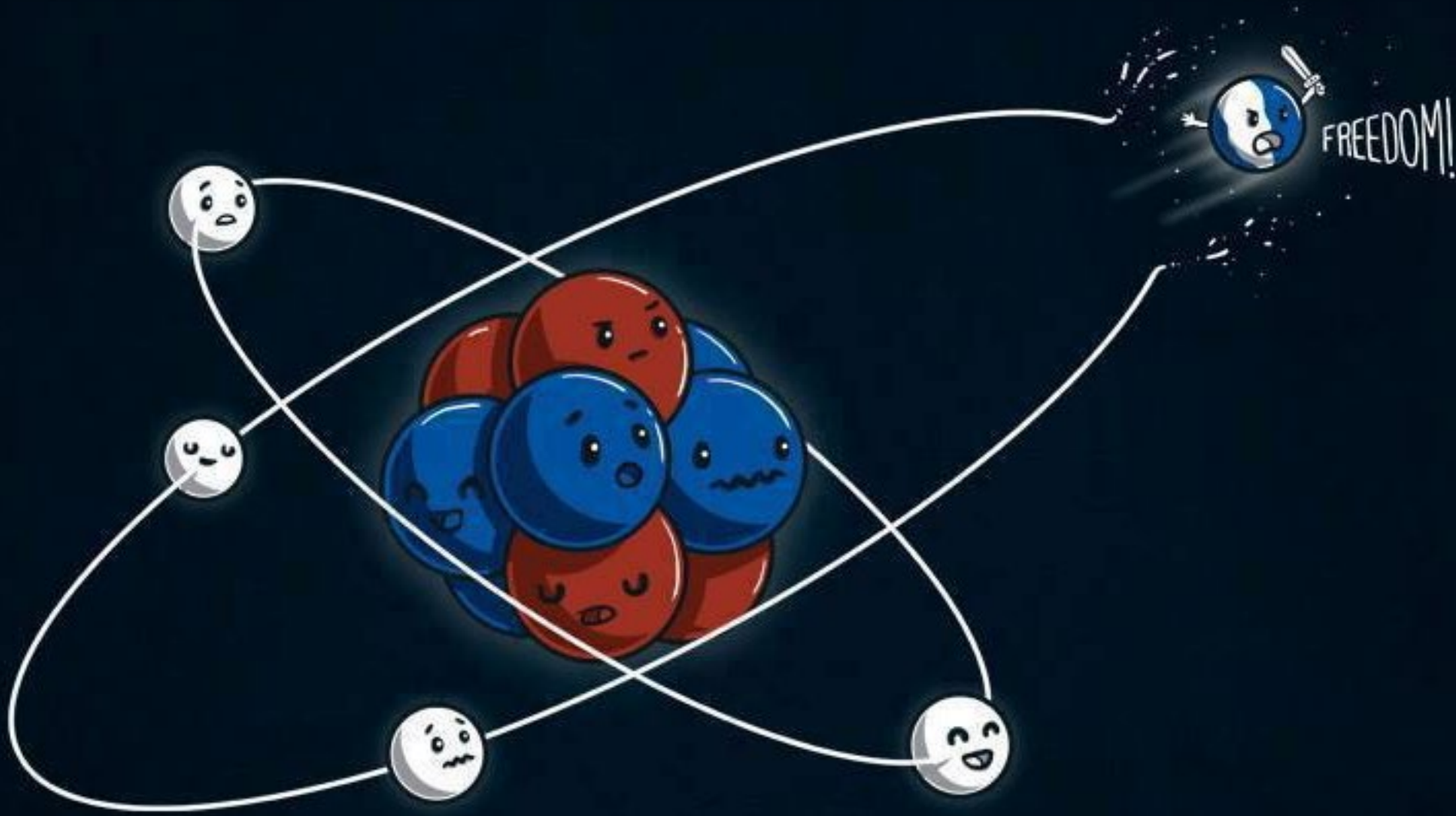


WHY DO ATOMS BOND TO FORM COMPOUNDS?



Chemical Bonds:

the attractive (intramolecular) forces, ie. “the glue”, strong enough to maintain a group of atoms together for an indefinite amount of time.

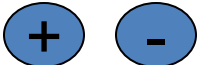
There are three basic types of bonds: Ionic, Covalent, & Metallic.

Bonding

Bonds

“glue”

Ionic

- e- transfer
- $\Delta EN > 1.6$
- Crystal lattice
-  ions
- $\text{Na} \cdot + \cdot \text{Cl} \cdot \rightarrow \text{Na}^+ \text{Cl}^-$
 $3s^1 \quad 3s^2 3p^5 \quad \text{Cl}^-$
- “salts”
- very strong bonds generally between metals and non metals.
- Solids at room temp.
- Poor conductors of electricity in a solid state

Covalent

- sharing
- $1.6 > \Delta EN > 0$
- “molecules”
- hypothetical charge
- $\text{H} \cdot + \cdot \text{H} \rightarrow \text{H} \cdot \text{H}$
 $1s^1 \quad 1s^1 \quad 1s^2$
- orbital overlap
- Bonding e- are localized between two atoms
- Formed between two nonmetals

Metallic

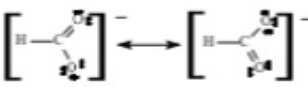
- “sea” of e-
- e- are delocalized.

Section _____

Name _____

Report Sheet

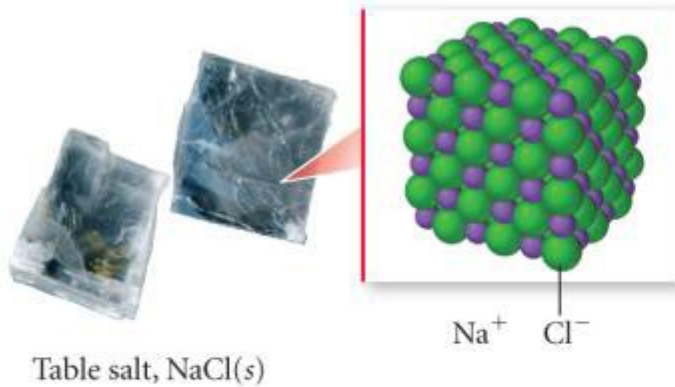
Complete the table following the pattern given in the example.

3-D molecular drawing, with Lewis resonance structures, if necessary	VSEPR geometry, AB_xE_y hybridization	Molecular shape, bond angles	Polarity, bond lengths
HCO_2^- 	Trigonal planar AB ₃ E ₀ Carbon is sp ² hybridized	Trigonal planar HCO=120° OCO=120°	Polar C-O = 1.23Å H-C = 1.14Å
CF ₄			
CHI ₃			
HCl			

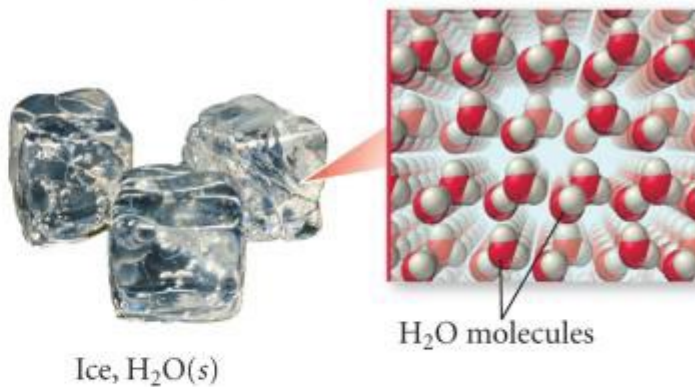
EXPERIMENT 11

Types of Bonding

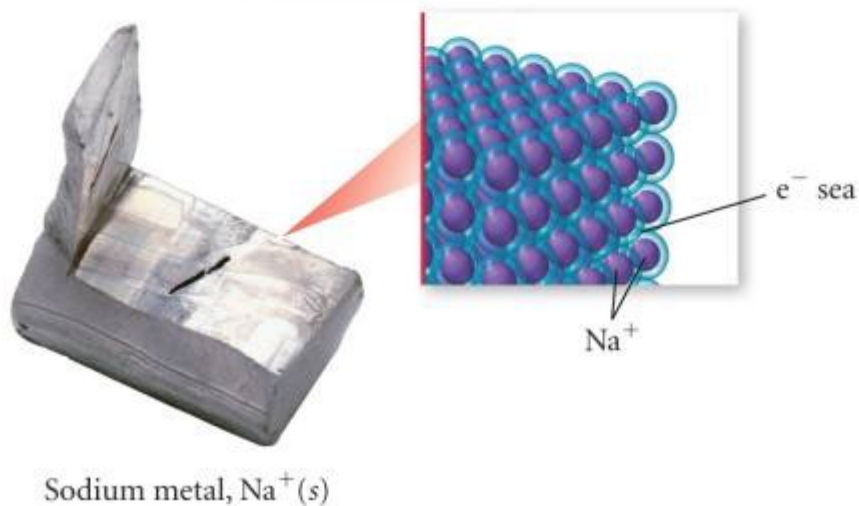
Ionic bonding



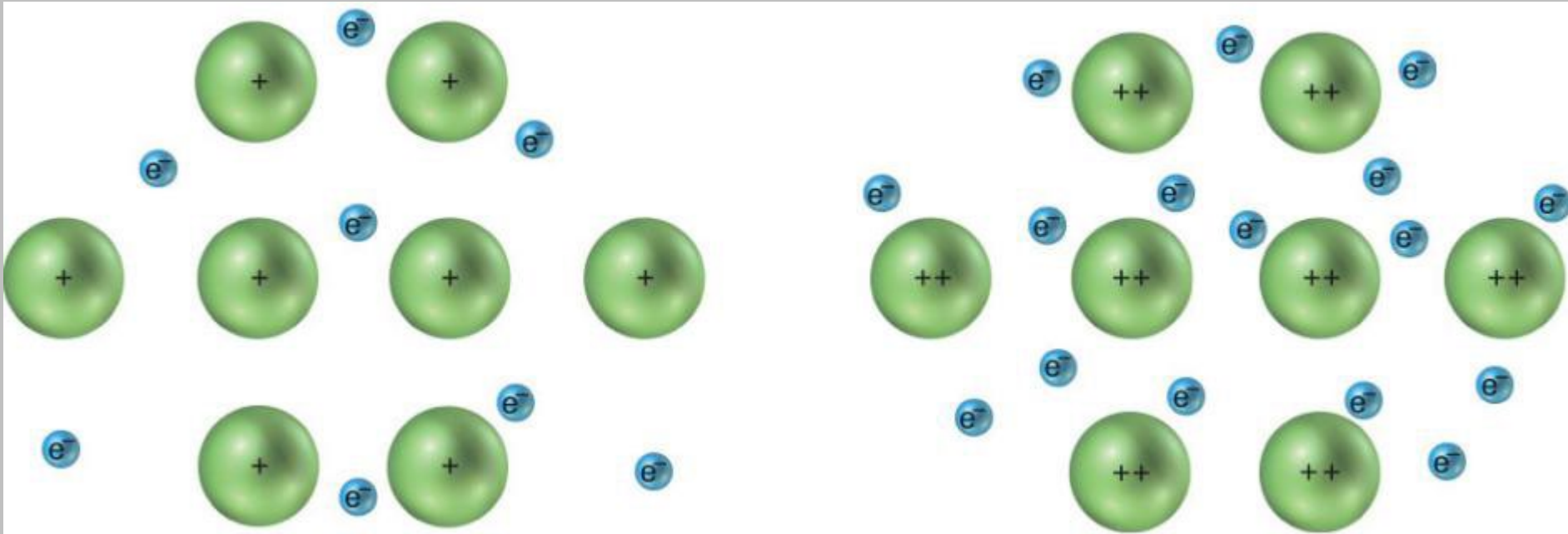
Covalent bonding



Metallic bonding



METALLIC BONDS



Metallic Bonds

- **low ionization energy of metals allows them to lose electrons easily**
- **the simplest theory of metallic bonding involves the metals atoms releasing their valence electrons to be shared by all to atoms/ions in the metal**
 - **an organization of metal cation islands in a sea of electrons**
 - **electrons delocalized throughout the metal structure**
- **bonding results from attraction of cation for the delocalized electrons**

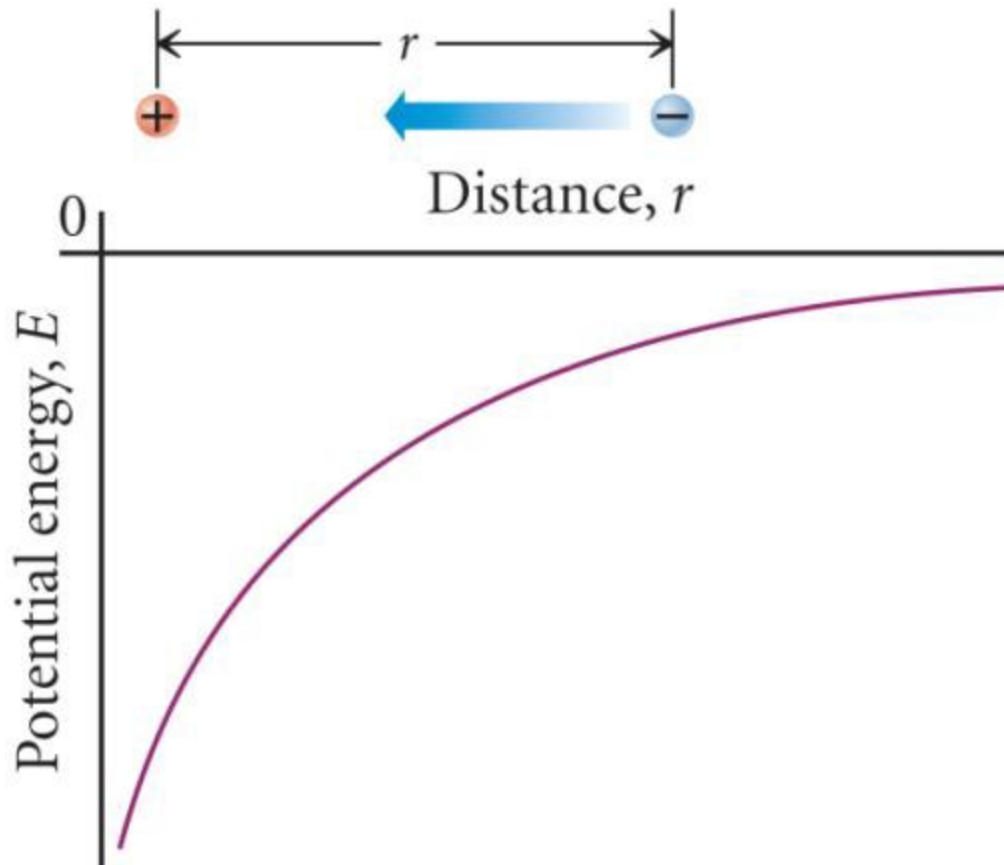
Metallic Bonds

- **metallic solids conduct electricity & heat**
- **as temperature increases, electrical conductivity decreases**
- **metallic solids reflect light**
- **metallic solids are malleable and ductile**
- **metals generally have high melting points and boiling points**
 - **all but Hg are solids at room temperature**
 - **melting points generally increase to right across period**
 - **melting points generally decrease down column**

Why Do Atoms Bond?

- processes are spontaneous if they result in a system with lower potential energy
- chemical bonds form because they lower the potential energy between the charged particles that compose atoms
- the potential energy between charged particles is directly proportional to the product of the charges
- the potential energy between charged particles is inversely proportional to the distance between the charges

Potential Energy Between Charged Particles



The attraction between oppositely charged particles increases as the particles get closer together. Bringing them closer together requires the potential energy of the system.

Bonding

- **a chemical bond forms when the potential energy of the bonded atoms is less than the potential energy of the separate atoms**
- **have to consider following interactions:**
 - **nucleus-to-nucleus repulsion**
 - **electron-to-electron repulsion**
 - **nucleus-to-electron attraction**

Chemical Bonding: An Introduction

1. chemical bond:

- ionic bond:

- covalent bond:

 - Electronegativity:

 - Polar Covalent:

2. octet rule:

EXCEPTIONS

3. Lewis dot structure

4. Formal Charge:

5. Resonance:

Isomers:

4. Types of Bonds:

- Single vs. double
vs. triple bonds:

- Bond Strength:

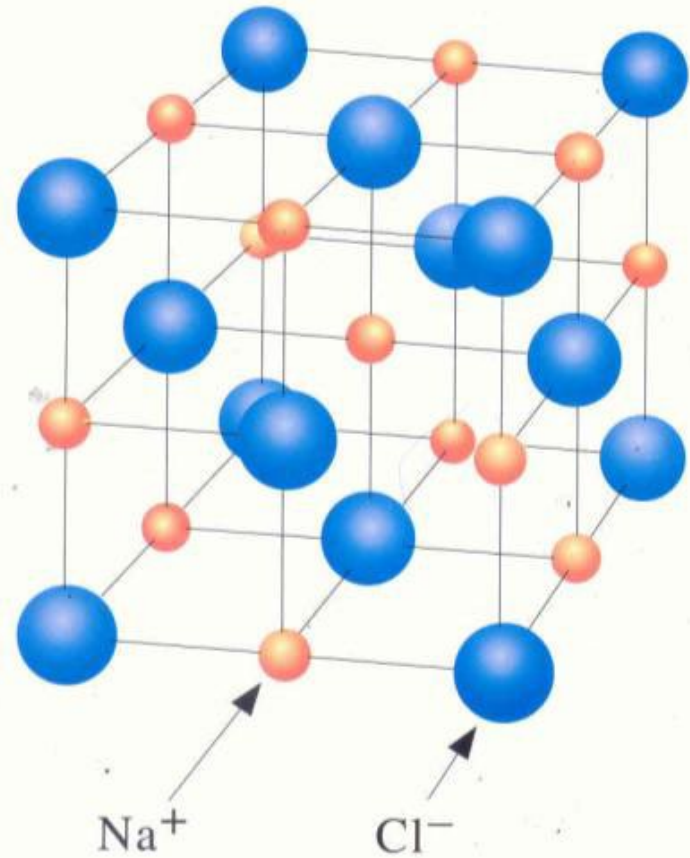
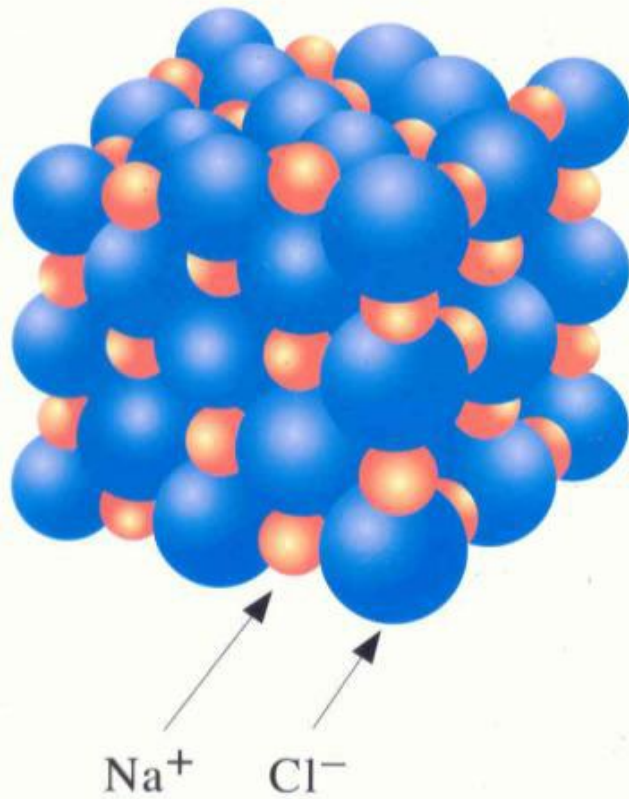
- Bond Length:

5. VSEPR Model

6. Polarity of Molecules

IONIC BONDS

Sodium chloride crystal



IONIC RADIUS

The ionic radius is a measure of the size of the spherical region (electron probability) around the nucleus of an ion.

Isoelectronic: refers to different species having the same electron configuration

- In general; the ionic radius of cations decreases going left to right across a period until the ion becomes an anion (Group V) then there is an abrupt increase in ionic radius.

From Group V to Group VII the radius once again decreases.

- The ionic radius increases going down a group.


Ionic Radius
cation


Ionic Radius
anion

H^+
0.0

$[\text{OH}_3]^+$
95

H^-
145

$[\text{NH}_4]^+$
148

N^{3-}
171

N
74

$[\text{OH}]^-$
136

O^{2-}
140

O
74

F^-
136

F
72

Li^+
68

Be^{2+}
35

B
80

C
77

Na^+
97

Mg^{2+}
66

Al^{3+}
51

P^{3-}
212

S^{2-}
184

Cl^-
181

K^+
133

Ca^{2+}
99

Ga^{3+}
62

Se^{2-}
198

Br^-
195

Rb^+
147

Sr^{2+}
122

In^{2+}
81

Te^{2-}
221

I^-
216

Cs^+
167

Ba^{2+}
134

Tl^+
147

Tl^{3+}
95


Covalent Radius
100 picometres

Sc^{3+}
81

Ti^{3+}
76

V^{2+}
88

Cr^{2+}
74

Mn^{2+}
80

Fe^{2+}
74

Co^{2+}
72

Ni^{2+}
69

Cu^+
96

Zn^{2+}
74

Ti^{4+}
68

V^{3+}
74

Cr^{3+}
63

Mn^{3+}
66

Fe^{3+}
64

Co^{3+}
63

Cu^{2+}
72

Pd^{2+}
80

Ag^+
126

Cd^{2+}
97

Pt^{2+}
80

Au^+
137

Hg^{2+}
110

IONIC BONDS

- *EXIST BETWEEN IONS OF OPPOSITE CHARGE*
- *ARE THE RESULT OF ELECTON TRANSFER BETWEEN IONS*
- *ARE HELD TOGETHER BY ATTRACTIVE ELECTROSTATIC FORCES*
$$E = \frac{k Q_1 Q_2}{r}$$
- *GENERALLY INVOLVE BONDING BETWEEN GROUP I & GROUP II CATIONS WITH GROUP VI & VII ANIONS.*
- *OCCURS BETWEEN IONS WITH VERY DIFFERENT ELECTRONEGATIVITIES*

Potential Energy Between Charged Particles

$$E_{\text{potential}} = \frac{1}{4\pi \epsilon_0} \left(\frac{q_1 \bullet q_2}{r} \right)$$

- ϵ_0 is a constant
– $= 8.85 \times 10^{-12} \text{ C}^2/\text{J}\cdot\text{m}$
- for charges with the same sign, $E_{\text{potential}}$ is + and the magnitude gets less positive as the particles get farther apart
- for charges with the opposite signs, $E_{\text{potential}}$ is – and the magnitude gets more negative as the particles get closer together
- remember: the more negative the potential energy, the more stable the system becomes

ATTRACTIVE FORCES

Coulomb's law for NaBr gives:

$$E = \frac{-kQ_1Q_2}{r} = \frac{-(8.99 \times 10^9 \frac{\text{J}}{\text{C}^{\text{M}}})(1.6 \times 10^{-19} \text{C})^2}{\approx 3.5 \times 10^{-10} \text{m}}$$

$$E = -6.59 \times 10^{-19} \text{J}$$

The molar value is

$$\begin{aligned} E \cdot N_A &= -6.59 \times 10^{-19} \text{ J} \cdot 6.022 \times 10^{23} \left(\frac{\text{kJ}}{1000 \text{J}} \right) \\ &= -397 \text{ kJ/mol} \end{aligned}$$

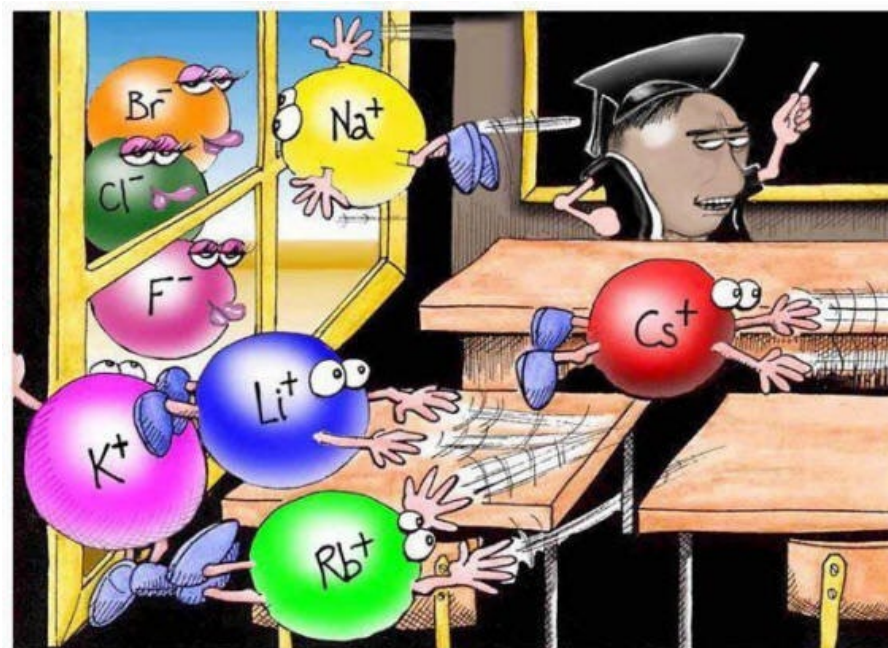
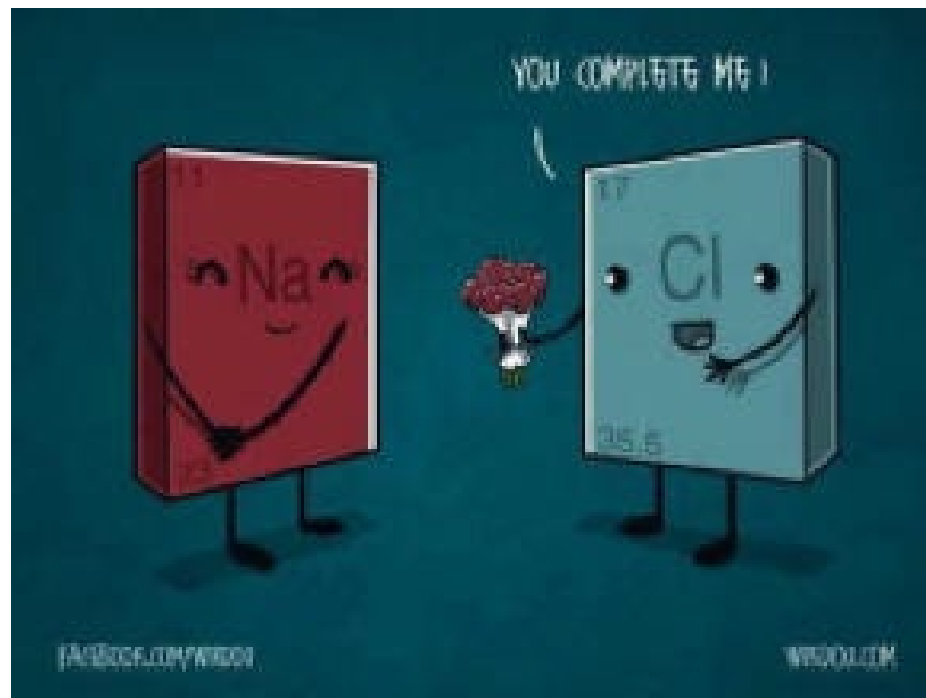
This is the amount of energy released for electrostatic attraction



BUT it does not tell the whole story. Additional energy is actually released during the formation of the Salt, NaBr.

This means the formation of a crystal is not as simple as moving two opposing charged Ions near each other.

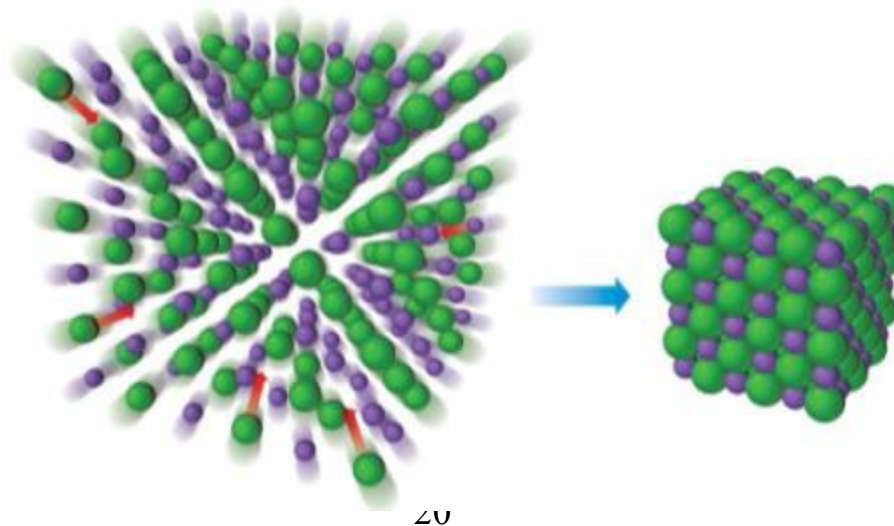
LATTICE ENERGY & Born-Haber Cycle



"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you all find so attractive...?"

Lattice Energy

- the **lattice energy** is the energy released when the solid crystal forms from separate ions in the gas state
 - always exothermic
 - hard to measure directly, but can be calculated from knowledge of other processes
- lattice energy depends directly on size of charges and inversely on distance between ions



Lattice Energy (2 of 2)

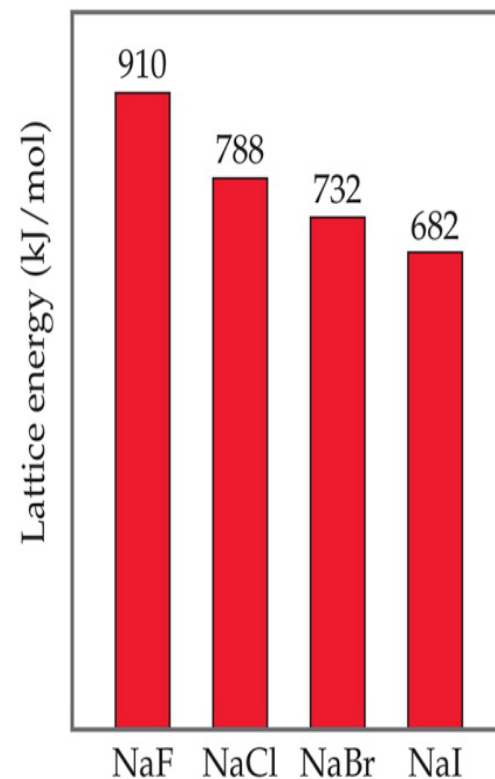
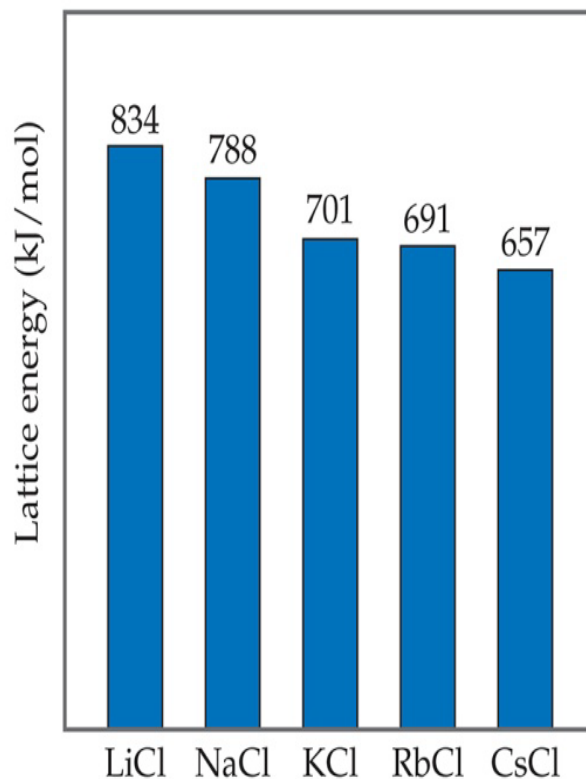
Table 8.1 Lattice Energies for Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
LiF	1030	MgCl ₂	2526
LiCl	834	SrCl ₂	2127
LiI	730		
NaF	910	MgO	3795
NaCl	788	CaO	3414
NaBr	732	SrO	3217
NaI	682		
KF	808	ScN	7547
KCl	701		
KBr	671		
CsCl	657		
CsI	600		

Trends in Lattice Energy

- Lattice energy increases with:
 - Increasing charge on the ions
 - Decreasing size of ions

$$E_{\text{el}} = \frac{\kappa Q_1 Q_2}{d}$$



Lattice Energy

THE ENERGY REQUIRED TO COMPLETELY SEPARATE ONE MOLE OF AN IONIC SOLID INTO ITS GASEOUS IONS.

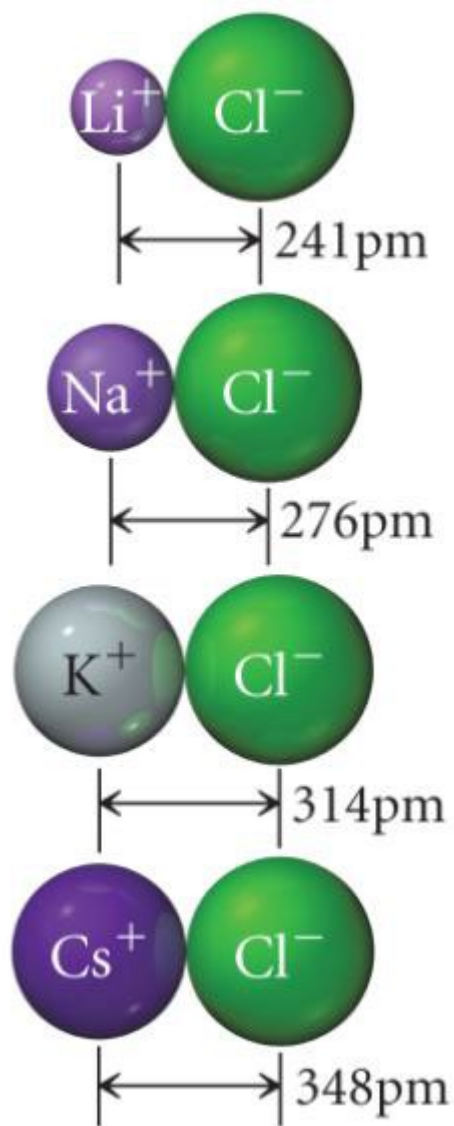


- the magnitude of the lattice energy depends on the amount of charge on the Ionic species
(low charge = low lattice energy)

larger ion ► weaker attraction ► smaller lattice energy

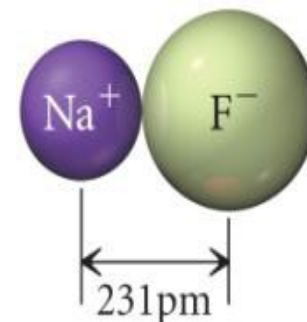
larger charge ► stronger attraction ► larger lattice energy

- the attractive interactions Increase with increasing charge and decreasing distances.

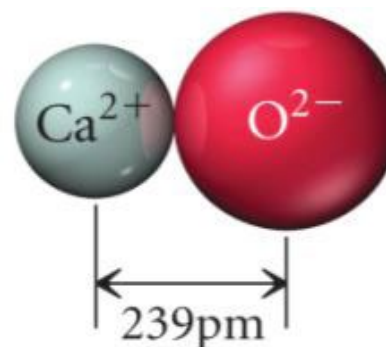


**larger ion ►
weaker
attraction ►
smaller lattice
energy**

Metal Chloride	Lattice Energy (kJ/mol)
LiCl	-834
NaCl	-787
KCl	-701
CsCl	-657



Lattice Energy =
-910 kJ/mol



Lattice Energy =
-3414 kJ/mol

**larger charge ►
stronger attraction
► larger lattice
energy**

Example 9.2 – Order the following ionic compounds in order of increasing magnitude of lattice energy.

CaO, KBr, KCl, SrO

First examine the ion charges and order by product of the charges

Ca^{2+} & O^{2-} , K^{+} & Br^{-} ,
 K^{+} & Cl^{-} , Sr^{2+} & O^{2-}

$(\text{KBr}, \text{KCl}) < (\text{CaO}, \text{SrO})$

Then examine the ion sizes of each group and order by radius; larger < smaller

(KBr, KCl) same cation,
 $\text{Br}^{-} > \text{Cl}^{-}$ (same Group)

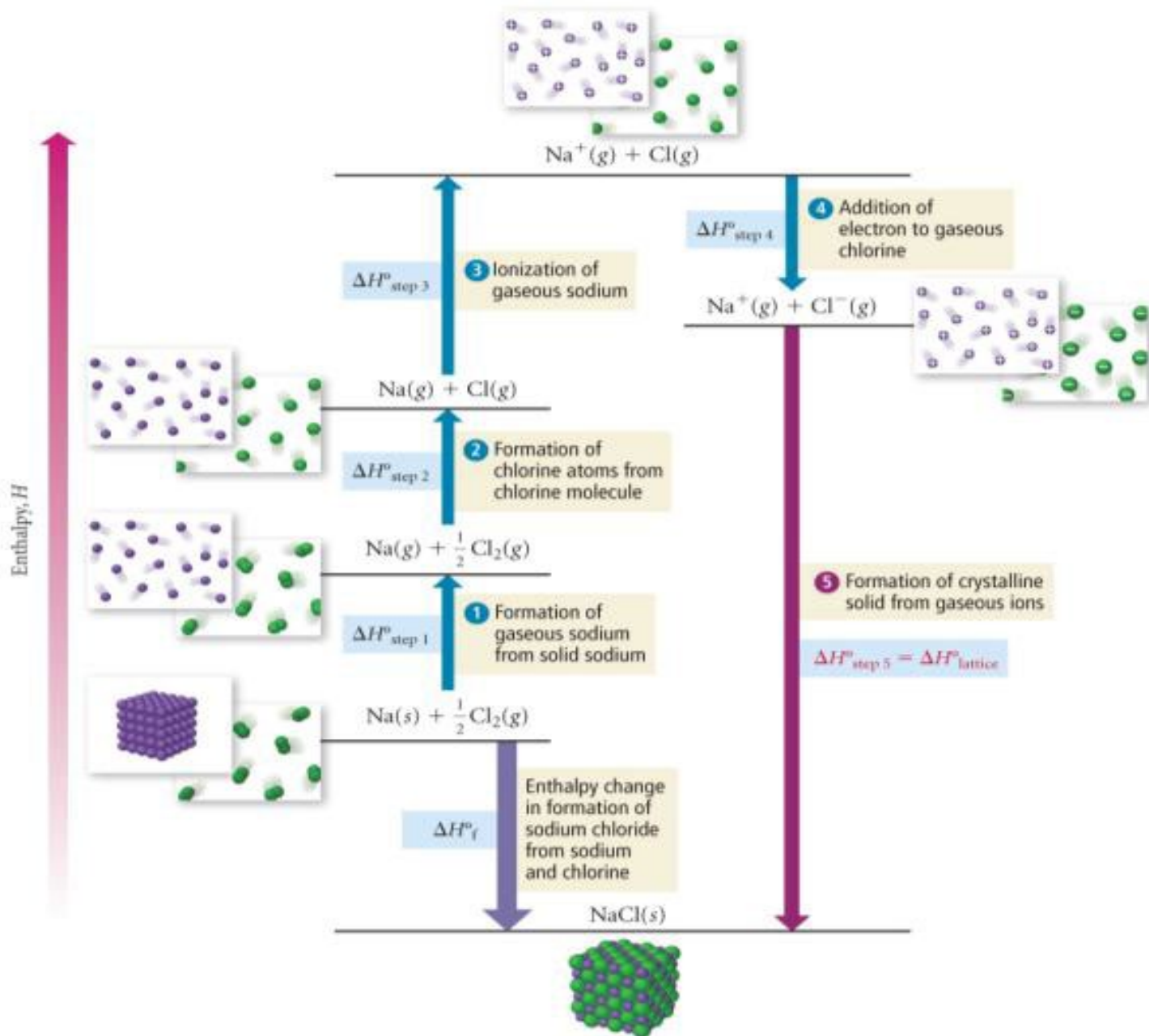
(CaO, SrO) same anion,
 $\text{Sr}^{2+} > \text{Ca}^{2+}$ (same Group)

$\text{KBr} < \text{KCl} < \text{SrO} < \text{CaO}$

Born-Haber Cycle

- method for determining the lattice energy of an ionic substance by using other reactions
 - use Hess's Law to add up heats of other processes
- $\Delta H^\circ_f(\text{salt}) = \Delta H^\circ_f(\text{metal atoms, g}) + \Delta H^\circ_f(\text{nonmetal atoms, g}) + \Delta H^\circ_f(\text{cations, g}) + \Delta H^\circ_f(\text{anions, g}) + \Delta H^\circ_f(\text{crystal lattice})$
 - $\Delta H^\circ_f(\text{crystal lattice}) = \text{Lattice Energy}$
 - metal atoms (g) \rightarrow cations (g), $\Delta H^\circ_f =$ ionization energy
 - don't forget to add together all the ionization energies to get to the desired cation
 - $M^{2+} = 1^{\text{st}} \text{ IE} + 2^{\text{nd}} \text{ IE}$
 - nonmetal atoms (g) \rightarrow anions (g), $\Delta H^\circ_f =$ electron affinity

Born-Haber Cycle for NaCl



Born-Haber cycle for NaCl

- The transfer of an electron from an Na atom to a Cl atom is not in itself energetically favorable; it requires 147 kJ/mol of energy (step 1).
- 493 kJ of energy is released when these oppositely charged ions come together to form ion pairs (step 2a)
- Additional energy 293 kJ is released when these ion pairs form the solid crystal (step 2b)
- The lattice energy released when 1 mol of Na^+ & Cl^- ions react to produce NaCl (s) is 786 kJ/mol
- The overall process of NaCl formation is energetically favorable, releasing 639 kJ/mol if gaseous Na & Cl atoms are initially used.

Born-Haber cycle for NaCl

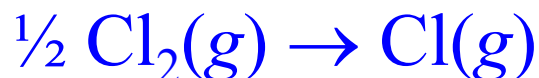
- $\text{Na(s)} \rightarrow \text{Na(g)}$ $\Delta H_1 = 108 \text{ kJ}$
- $\frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{Cl(g)}$ $\Delta H_2 = 120 \text{ kJ}$
- $\text{Na(g)} \rightarrow \text{Na}^+(\text{g}) + \text{e}^-(\text{g})$ $\Delta H_3 = 496 \text{ kJ}$
- $\text{Cl(g)} + \text{e}^-(\text{g}) \rightarrow \text{Cl}^-(\text{g})$ $\Delta H_4 = -349 \text{ kJ}$
- $\text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{NaCl(s)}$ $\Delta H_5 = -786 \text{ kJ}$
- ---
- $\text{Na(s)} + \frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{NaCl(s)}$ $\Delta H^\circ_f = ?$
-411 kJ/mol

****Step 1: Sublimation of sodium, Step 2: dissociation of chlorine, Step 3: ionization of sodium, Step 4: formation of chloride ion, Step 5: formation of sodium chloride from ions.**

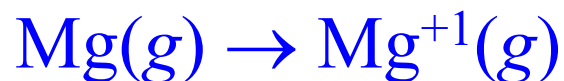
Practice - Given the Information Below, Determine the Lattice Energy of MgCl_2



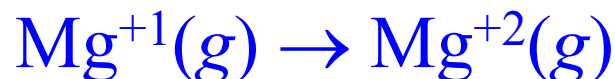
$$\Delta H_1^\circ_f = +147.1 \text{ kJ/mol}$$



$$\Delta H_2^\circ_f = +121.3 \text{ kJ/mol}$$



$$\Delta H_3^\circ_f = +738 \text{ kJ/mol}$$



$$\Delta H_4^\circ_f = +1450 \text{ kJ/mol}$$



$$\Delta H_5^\circ_f = -349 \text{ kJ/mol}$$

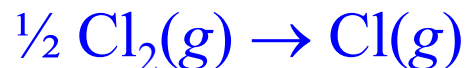


$$\Delta H_6^\circ_f = -641.3 \text{ kJ/mol}$$

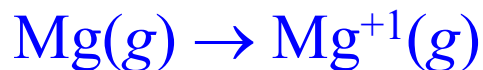
Practice - Given the Information Below, Determine the Lattice Energy of MgCl_2



$$\Delta H_{1f}^{\circ} = +147.1 \text{ kJ/mol}$$



$$\Delta H_{2f}^{\circ} = +121.3 \text{ kJ/mol}$$



$$\Delta H_{3f}^{\circ} = +738 \text{ kJ/mol}$$



$$\Delta H_{4f}^{\circ} = +1450 \text{ kJ/mol}$$



$$\Delta H_{5f}^{\circ} = -349 \text{ kJ/mol}$$



$$\Delta H_{6f}^{\circ} = -641.3 \text{ kJ/mol}$$

$$\Delta H_{6f}^{\circ} = \Delta H_{1f}^{\circ} + 2\Delta H_{2f}^{\circ} + \Delta H_{3f}^{\circ} + \Delta H_{4f}^{\circ} + 2\Delta H_{5f}^{\circ} + \Delta H_f^{\circ}(\text{lattice energy})$$

$$\Delta H_f^{\circ}(\text{lattice energy}) = \Delta H_{6f}^{\circ} - \left(\Delta H_{1f}^{\circ} + 2\Delta H_{2f}^{\circ} + \Delta H_{3f}^{\circ} + \Delta H_{4f}^{\circ} + 2\Delta H_{5f}^{\circ} \right)$$

$$\Delta H_f^{\circ}(\text{lattice energy}) = (-641.3 \text{ kJ}) - ((+147.1 \text{ kJ}) + 2(+121.3 \text{ kJ}) + (+738 \text{ kJ}) + (+1450 \text{ kJ}) + 2(-349 \text{ kJ}))$$

$$\Delta H_f^{\circ}(\text{lattice energy}) = -2521 \text{ kJ}$$

The Born-Haber Cycle

1. For a given solid, the difference in molar enthalpy between the solid and a gas of widely separated ions is called the lattice enthalpy (energy) of the solid. That is, lattice enthalpy (or energy) is the energy required to separate gaseous ions from a solid.



A. The lattice enthalpy of a solid CANNOT be measured directly.

B. However, it can be calculated by utilizing the first law of thermodynamics and, in particular, the fact that enthalpy is a state function.

C. This calculation invokes the use of a Born-Haber cycle, a closed path of steps, one of which is the formation of a solid lattice from its gaseous ions. The cycle traces the enthalpy changes that occur:

1. beginning with the pure elements,
2. atomize the pure elements to form gaseous atoms,
3. ionize the atoms to form gaseous ions,
4. allow the ions to form an ionic solid,
5. AND convert the solid back into the pure elements.

ACTIVITY on Born-Haber Process

Problem #1: Devise a Born-Haber cycle to calculate the lattice enthalpy of barium iodide.

Problem #2: Devise a Born-Haber cycle to calculate the lattice enthalpy of calcium fluoride.

Problem #3: Devise a Born-Haber cycle to calculate the lattice enthalpy of magnesium bromide.

Chemical Bonding: An Introduction

1. chemical bond:

- ionic bond:

- covalent bond:

 - Electronegativity:

 - Polar Covalent:

2. octet rule:

EXCEPTIONS

3. Lewis dot structure

4. Formal Charge:

5. Resonance:

Isomers:

4. Types of Bonds:

- Single vs. double
vs. triple bonds:

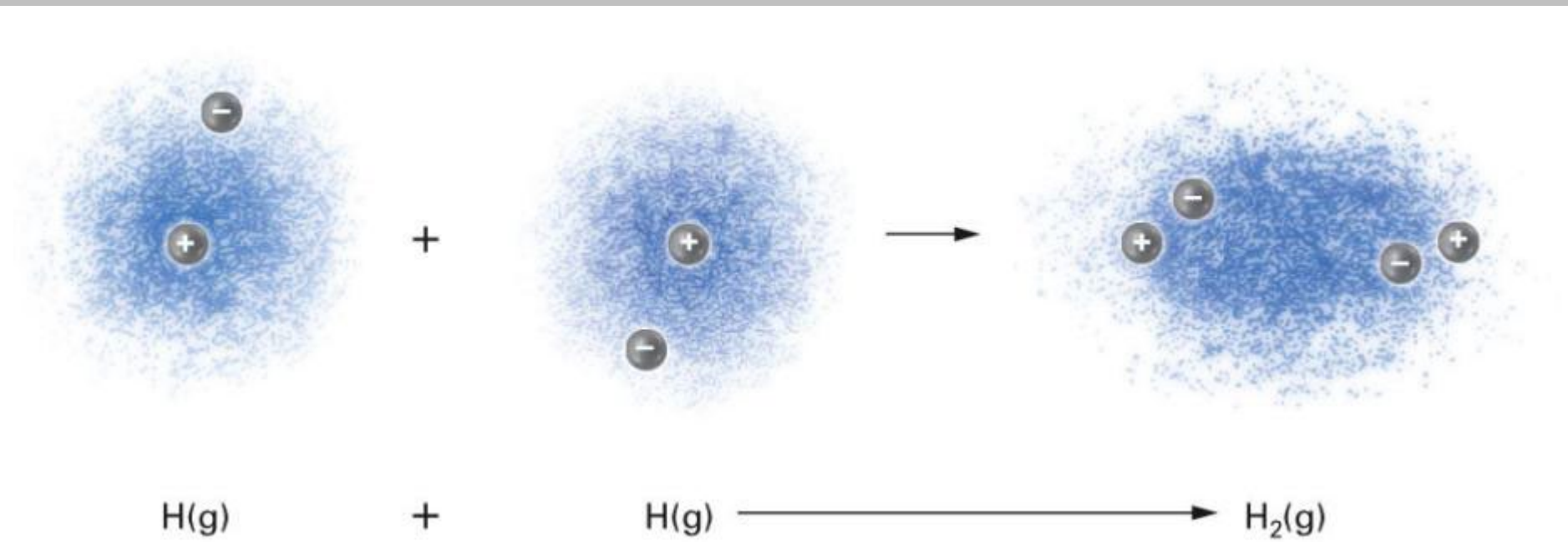
- Bond Strength:

- Bond Length:

5. VSEPR Model

6. Polarity of Molecules

COVALENT BONDS



ELECTRONEGATIVITY

- A PROPERTY OF ALL ATOMS
- IS THE ABILITY OF AN ATOM TO ATTRACT A PAIR OF SHARED ELECTRONS.
- METALS GENERALLY HAVE LOW ELECTRONEGATIVITY VALUES, METALS TEND TO LOSE ELECTRONS.
- NONMETALS GENERALLY HAVE HIGH ELECTRONEGATIVITY VALUES, NONMETALS TEND TO GAIN ELECTRONS.
- FLUORINE HAS THE HIGHEST VALUE (4.0) AND FRANCIUM HAS THE LOWEST (0.7)
- GENERAL TREND
INCREASES GOING ACROSS A PERIOD AND
DECREASES GOING DOWN A COLUMN

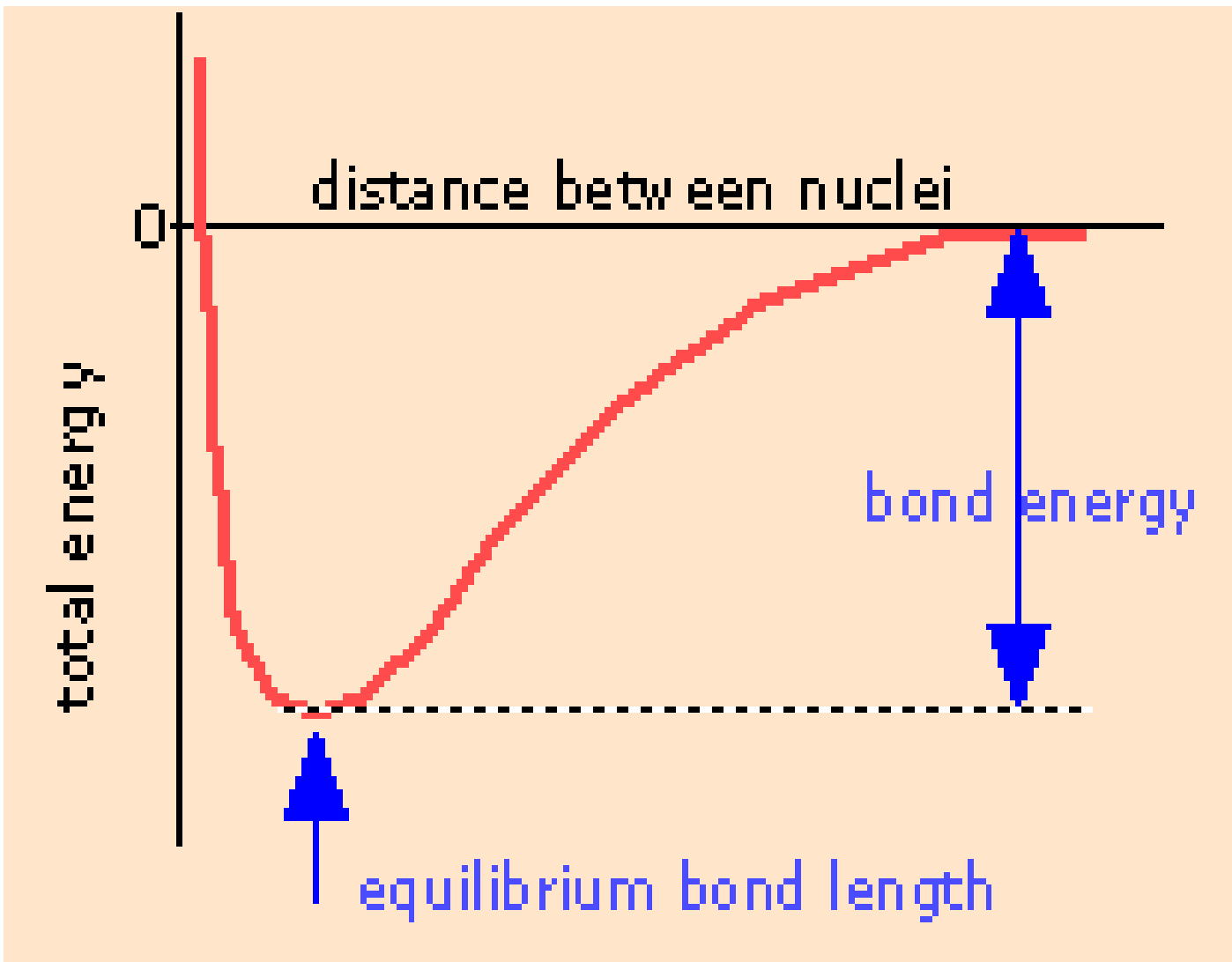
CHEMICAL BONDS

COVALENT BONDS

- ARE FORMED BY THE SHARING OF A PAIR OF ELECTRONS BETWEEN TWO ATOMS.
- INVOLVES THE OVERLAPPING OF THE VALENCE ORBITALS.
- THE PROBABILITY OF FINDING THE VALENCE ELECTRON OF ONE ATOM IS HIGHEST BETWEEN THE TWO NUCLEI.
- EACH VALENCE ELECTRON IS EQUALLY ATTRACTED TO BOTH NUCLEI, WHICH STABILIZES THE MOLECULE.
- GENERALLY FULFILLS THE OCTET RULE FOR BOTH ATOMS.
- OCCURS BETWEEN ATOMS OF SIMILAR ELECTRONEGATIVITIES ($\Delta \approx 0$ OR VERY SMALL)

Potential Energy Diagram

- The overall energy term involved is due to the net potential energy, which results from the attractive & repulsive forces between charged particles, and the Kinetic energy due to the motion of the electrons.
- The zero point energy is defined as where the atoms are at infinite separation.
- At very short distances the energy rises steeply due to the repulsive forces that exist when atoms are close together.
- The bond length is at a distance where the system has minimal energy.



BOND LENGTH

- "THE DISTANCE BETWEEN NUCLEI AT ITS MINIMUM ENERGY.
- In order to break covalent bonds the **BOND DISSOCIATION ENERGY** must be added to the system.
- In general the larger the **BDE**, the stronger the bond.
- Average bond lengths can be estimated by using the trends in the periodic table.
- $C-C$ 1.54 Å $C=C$ 1.34 Å $C\equiv C$ 1.20 Å

BOND ORDER

- Describes the number of electron pairs being shared between any two bonded atoms.
- Single bonds \equiv B.O. of 1
- Double bonds \equiv B.O. of 2
- Triple bonds \equiv B. O. of 3

DIPOLE MOMENT & POLARITY

Polarity is a term that describes the charge distribution about either a bond or a molecule.

If the electrons involved in a bond are unequally distributed along the bond axis and the electrons reside closer to the more electronegative atom, this bond is referred to as a **POLAR covalent bond**.

If the electrons are equally distributed (shared) along the bond axis then this type of bond is called **NONPOLAR covalent bonds**.

A **dipole moment** can be used to visualize the polarity of the bond and is represented by an arrow pointing towards the more electronegative atom and a “+” on the tail.

POLARITY OF A BOND

+☐

H-Cl

2.1 - 3.0

$\Delta = 0.9$

+☐

S - N

2.5 - 3.0

$\Delta = 0.5$

+☐

C - O

2.5 - 3.5

$\Delta = 1.0$

If bonds were formed between any of the above examples, the bond would be referred to as a **polar covalent bond**. A dipole moment can be drawn in each case because the difference in electronegativity is greater than zero.

The example below has zero difference in electronegativity and no dipole can be drawn.

S - C

2.5 - 2.5

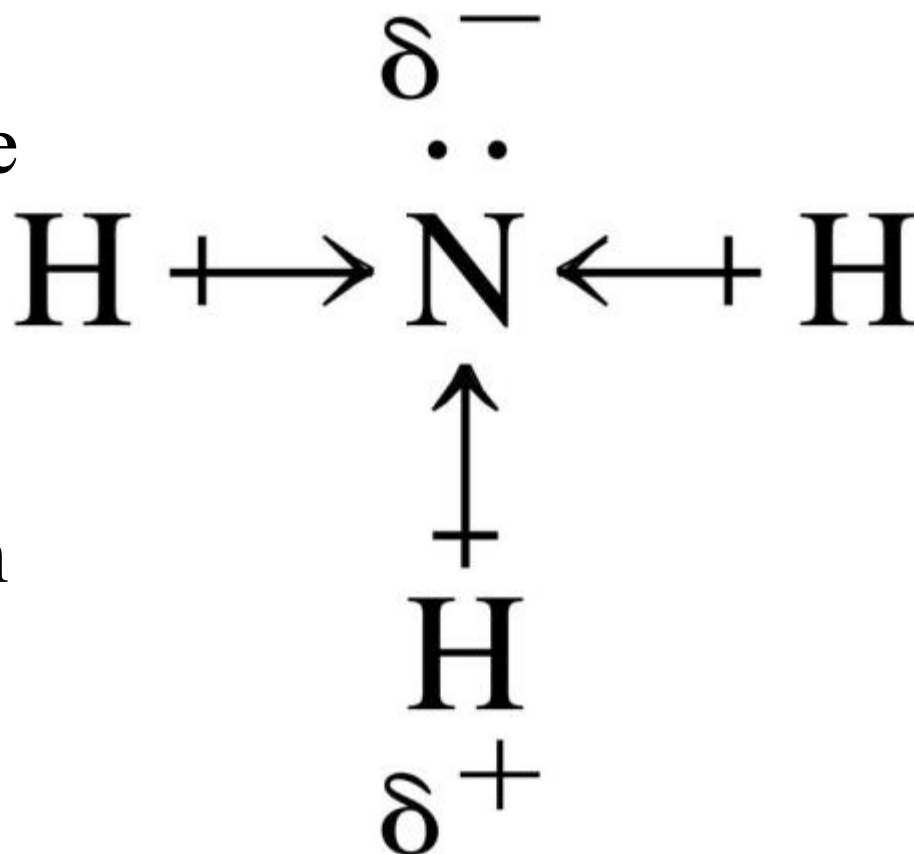
nonpolar covalent bond

☐ $\Delta = 0$

dipole moment, μ , is a measure of bond polarity

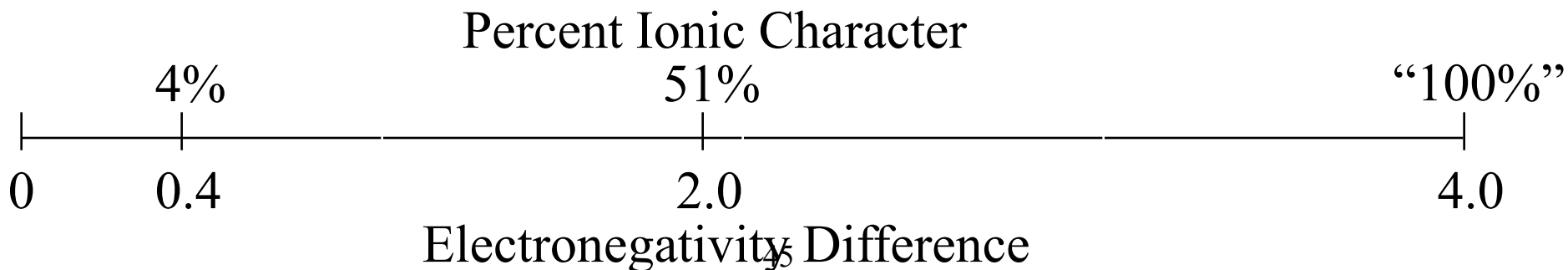
-it is directly proportional to the size of the partial charges and directly proportional to the distance between them

$\mu = (q)(r)$
(not Coulomb's Law)
measured in Debyes, D



Electronegativity and Bond Polarity

- If difference in electronegativity between bonded atoms is 0, the bond is **pure covalent**
 - equal sharing
- If difference in electronegativity between bonded atoms is 0.1 to 0.4, the bond is **nonpolar covalent**
- If difference in electronegativity between bonded atoms 0.5 to 1.9, the bond is **polar covalent**
- If difference in electronegativity between bonded atoms larger than or equal to 2.0, the bond is **ionic**



Putting it all together QUIZ

1) Identify the type of bonding involved in the following:



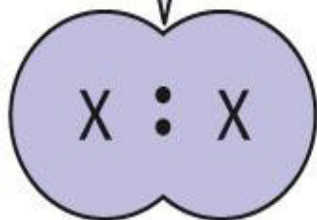
2) Determine the polarity of each bond given in #1

3) Create the Born-Haber cycle for SrO

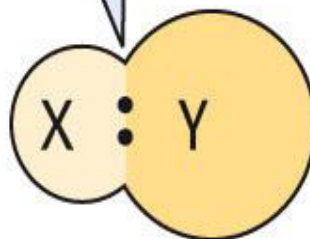
4) If the lattice energy of SrO is 3217kJ/mol, calculate the approximate bonding radius?

Chemical Bonds

Nonpolar covalent bond; equal sharing of bonding pair



Polar covalent bond; unequal sharing of bonding pair



Ionic bonding; transfer of electron

