

## CHEMICAL BONDING

- Chemical bonds are strong attractive force that exists between the atoms of a substance

Chemical Bonds are commonly classified into 3 types:

### 1. IONIC BONDING

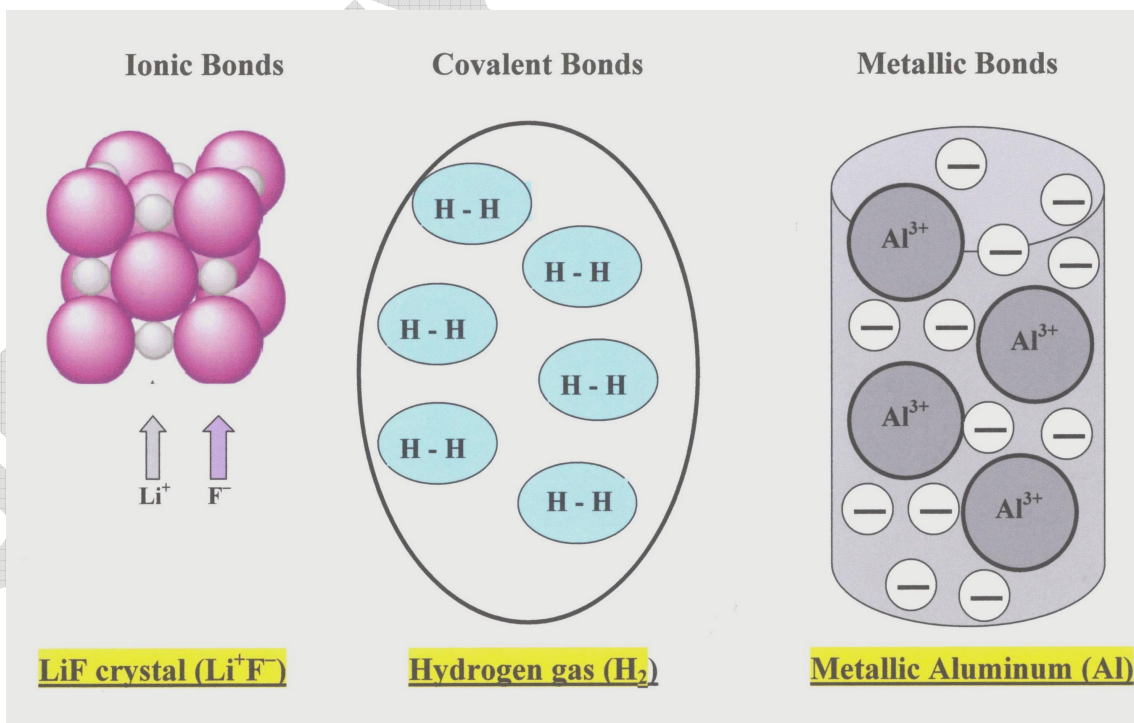
- Ionic bonds usually form between metals and nonmetals
- Bonds form by transfer of electrons
- Ex: NaCl, MgBr<sub>2</sub>, AlF<sub>3</sub>, etc

### 2. COVALENT BONDING

- Covalent bonds usually form between atoms of nonmetals
- Nonmetallic atoms may be the same or different
- Bonds form by sharing of electrons
- Ex: H<sub>2</sub>, Cl<sub>2</sub>, N<sub>2</sub>, HCl, H<sub>2</sub>O, NH<sub>3</sub>, CO<sub>2</sub>, CCl<sub>4</sub>

### 3. METALLIC BONDING

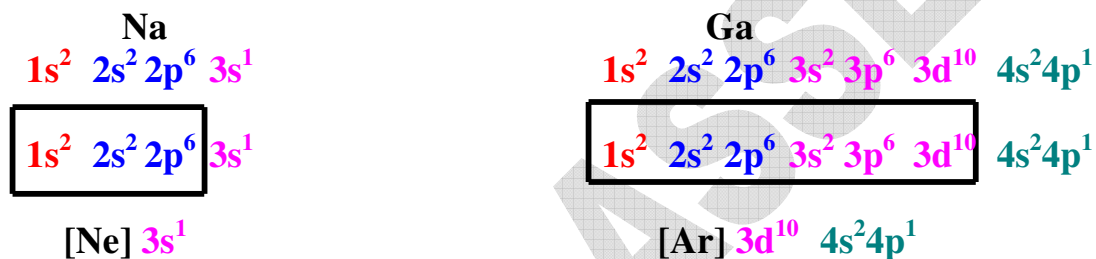
- Metallic bonds form between metallic ions
- Bonds are formed through “non-localized” sharing of electrons commonly referred to as “Sea of electrons”
- Ex: Cu, Al, Na



## LEWIS ELECTRON – DOT SYMBOLS

- When atoms bond, the bonding is usually achieved by the interaction of **valence electrons**
- Recall: Valence electrons are the “s” and “p” electrons of the outermost shell
- In Lewis Electron – Dot Symbols the dots represent valence electrons

### Examples:

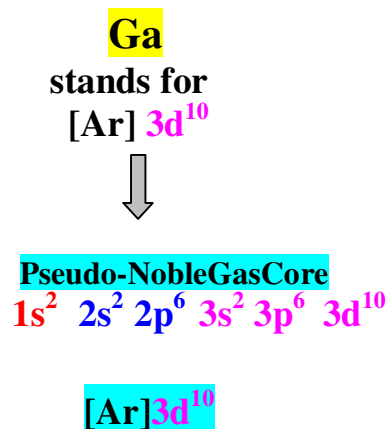
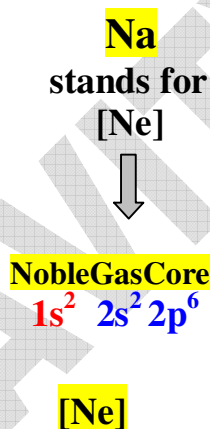


Lewis

Electron-Dot  
Symbol



Note:



### Conclusion:

In Lewis Electron – Dot Symbols:

- **dots** represent **valence electrons**
- **symbols** of elements represent **Noble Gas Cores** or **Pseudo-Noble Gas Cores**

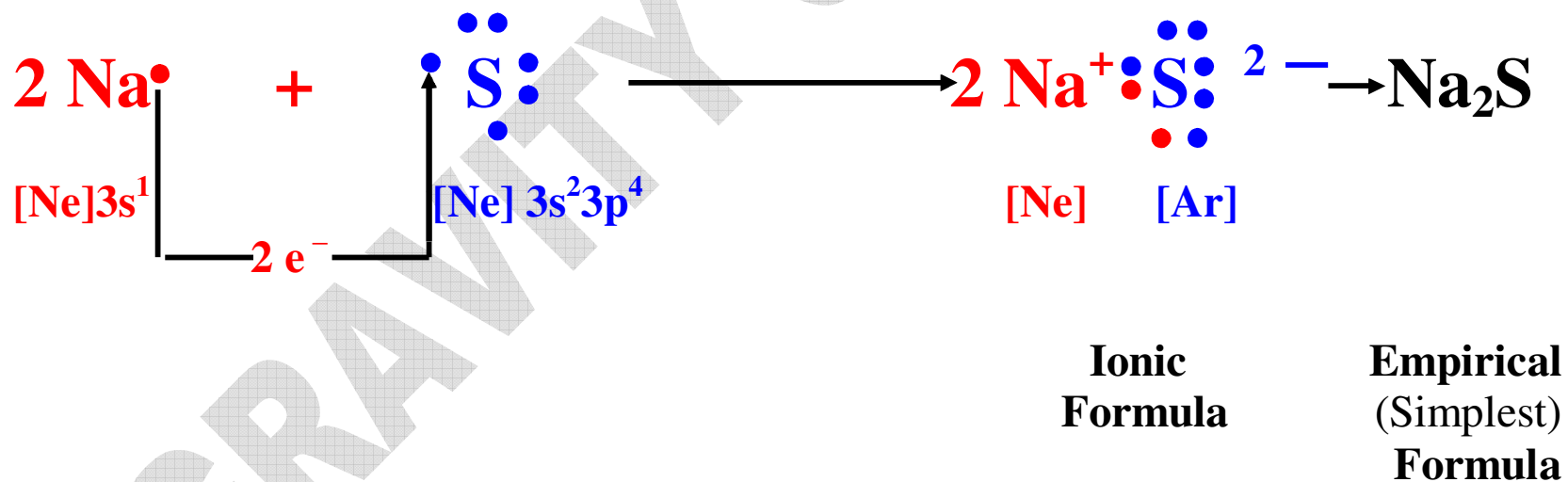
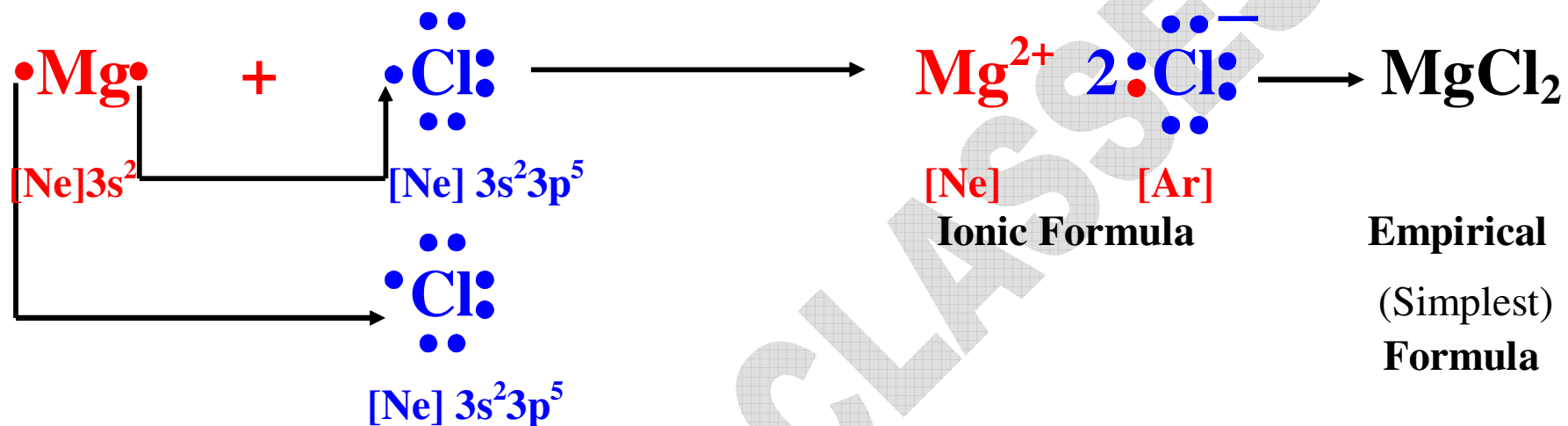
## LEWIS - ELECTRON DOT SYMBOLS FOR REPRESENTATIVE ELEMENTS

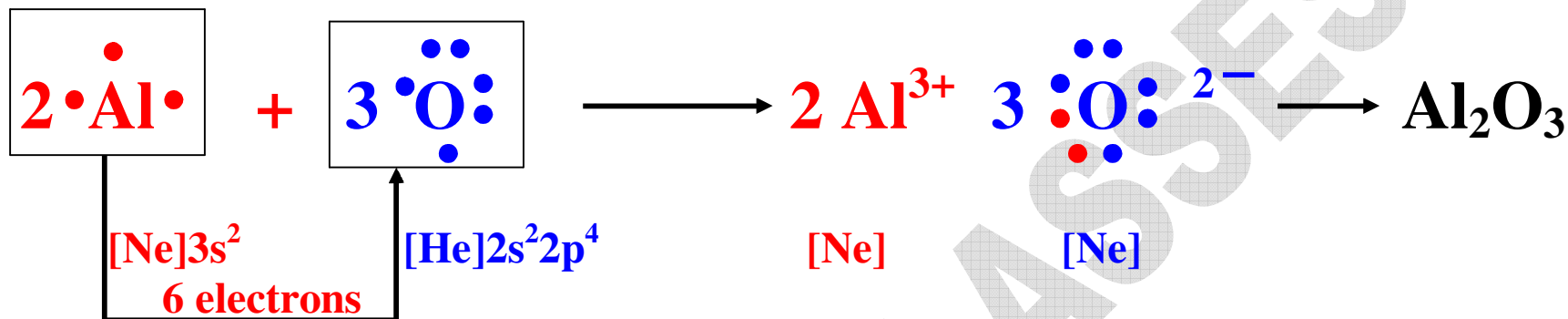
	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
2	Li [He]2s <sup>1</sup>	Be [He]2s <sup>2</sup>	B [He]2s <sup>2</sup> 2p <sup>1</sup>	C [He]2s <sup>2</sup> 2p <sup>2</sup>	N [He]2s <sup>2</sup> 2p <sup>3</sup>	O [He]2s <sup>2</sup> 2p <sup>4</sup>	F [He]2s <sup>2</sup> 2p <sup>5</sup>	Ne [He]2s <sup>2</sup> 2p <sup>6</sup>
	Li •	•Be •	•B •	•C •	•N •	•O •	•F •	•Ne •
3	Na [Ne]3s <sup>1</sup>	Mg [Ne]3s <sup>2</sup>	Al [Ne]3s <sup>2</sup> 3p <sup>1</sup>	Si [Ne]3s <sup>2</sup> 3p <sup>2</sup>	P [Ne]3s <sup>2</sup> 3p <sup>3</sup>	S [Ne] 3s <sup>2</sup> 3p <sup>4</sup>	Cl [Ne] 3s <sup>2</sup> 3p <sup>5</sup>	Ar [Ne]3s <sup>2</sup> 3p <sup>6</sup>
	Na •	•Mg •	•Al •	•Si •	•P •	•S •	•Cl •	•Ar •
	ns <sup>1</sup>	ns <sup>2</sup>	ns <sup>2</sup> np <sup>1</sup>	ns <sup>2</sup> np <sup>2</sup>	ns <sup>2</sup> np <sup>3</sup>	ns <sup>2</sup> np <sup>4</sup>	ns <sup>2</sup> np <sup>5</sup>	ns <sup>2</sup> np <sup>6</sup>
	X •	•X •	•X •	•X •	•X •	•X •	•X •	•X •

NumberofDots=NumberofValenceElectrons=GroupNumber

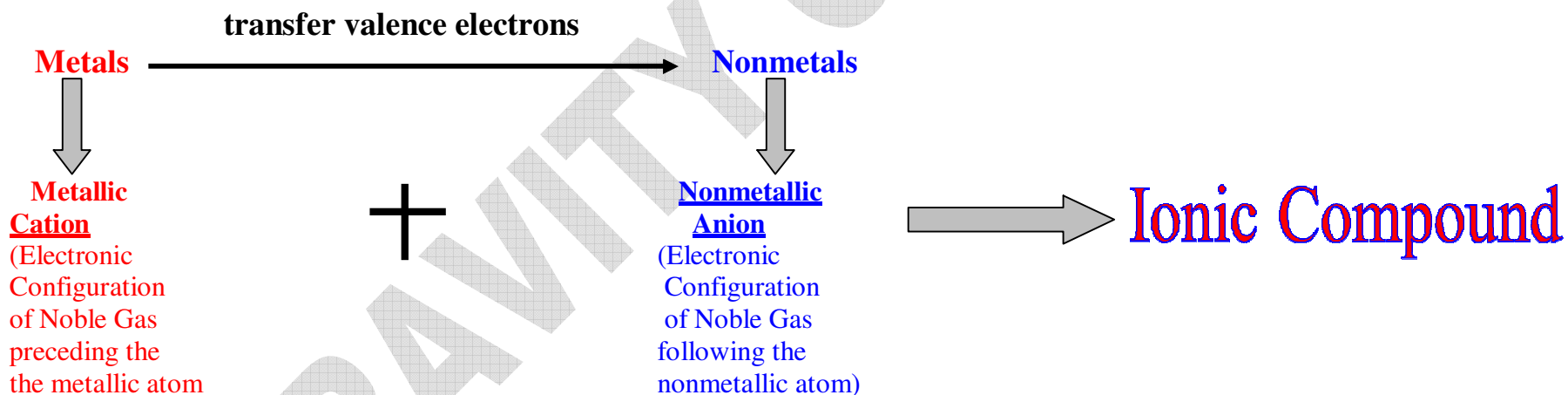


## OTHER EXAMPLES OF FORMATION OF IONIC COMPOUNDS

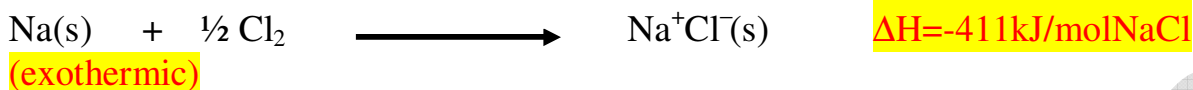




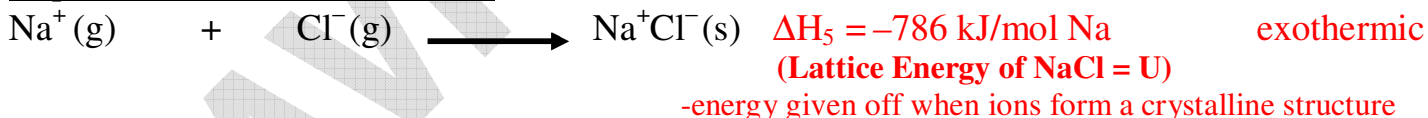
- **Ionic Bond forms by the transfer of one or more valence electrons from a metallic atom to a nonmetallic atom**



<b>ENERGY INVOLVED IN IONIC BONDING</b>
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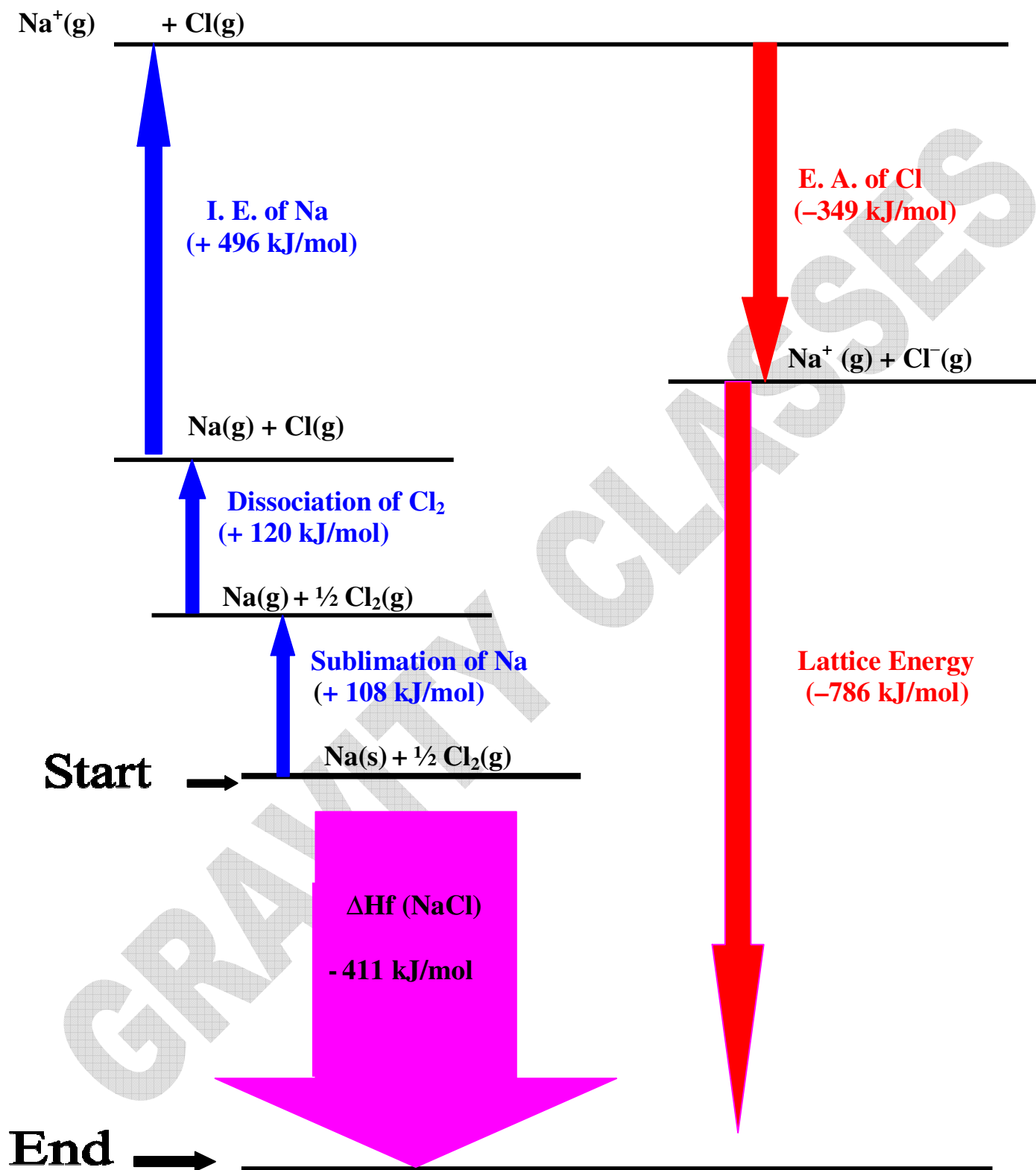


- Where is this energy coming from? The reaction can be considered to take place in 5 steps:

**Step 1: Sublimation of Sodium (solid Na → gaseous Na)****Step 2: Dissociation of Chlorine molecules (Cl<sub>2</sub> molecules → Cl atoms)****Step 3: Ionization of Sodium (Na atoms → Na<sup>+</sup> ions)****Step 4: Formation of Cl<sup>-</sup> (Cl atoms → Cl<sup>-</sup> ions)****Step 5: Formation of NaCl(s) from ions**

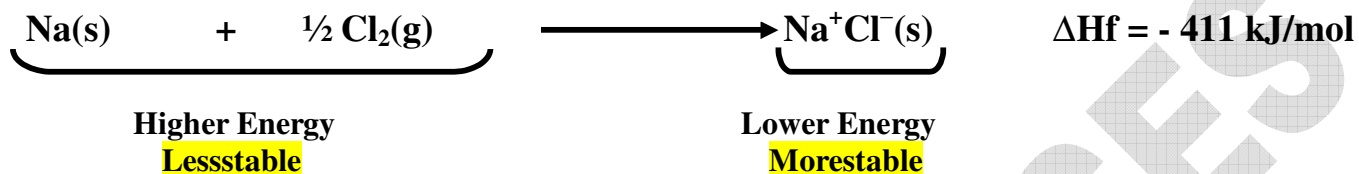
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$$\text{Net Energy Released} = +108 \text{ kJ} + 120 \text{ kJ} + 496 \text{ kJ} - 349 \text{ kJ} - 786 \text{ kJ} = -411 \text{ kJ/mol}$$



## INTERPRETING THE ENERGETICS OF IONIC BONDING

- The Net Ionic Reaction of the formation of  $\text{Na}^+\text{Cl}^-(\text{s})$  is exothermic:



### NOTE:

- The system's energy is lowered
- The larger the energy given off:
  - the more stable the ionic compound
  - the easier the ionic compound forms

### CONCLUSIONS

- Formation of an Ionic Compound is favored by any factor that increases the Net Energy Released:

- Low Ionization Energy** (characteristic of **Metals**)
- High Electron Affinity** (characteristic of **Nonmetals**)
- High Lattice Energy. The Lattice Energy depends on:
  - respective charges of ions
  - respective sizes of ions

- Most compounds formed between **Metals** and **Nonmetals** are **Ionic Compounds**

**Elements on the lower left-hand side of Periodic Table (Metals with low I.E.)**

+

**Elements on the upper right-hand side of Periodic Table (Nonmetals with high E.A.) (Noble Gases are excluded)**

→

**Ionic Compounds**

<b>IONS OF MAIN GROUP ELEMENTS</b>
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	IA	IIA	IIIA	IVA	VA	VIA	VIIA
Period 1							H <sup>-</sup>
Period 2	Li <sup>+</sup>	covalent	covalent	covalent	N <sup>3-</sup>	O <sup>2-</sup>	F <sup>-</sup>
Period 3	Na <sup>+</sup>	Mg <sup>2+</sup>	Al <sup>3+</sup>	covalent	P <sup>3-</sup>	S <sup>2-</sup>	Cl <sup>-</sup>
Period 4	K <sup>+</sup>	Ca <sup>2+</sup>	Ga <sup>3+</sup>	covalent	As <sup>3-</sup>	Se <sup>2-</sup>	Br <sup>-</sup>
Period 5	Rb <sup>+</sup>	Sr <sup>2+</sup>	In <sup>3+</sup>	Sn <sup>2+</sup> Sn <sup>4+</sup>	Sb <sup>3+</sup>	Te <sup>2-</sup>	I <sup>-</sup>
Period 6	Ca <sup>+</sup>	Ba <sup>2+</sup>	Tl <sup>3+</sup> Tl <sup>+</sup>	Pb <sup>2+</sup> Pb <sup>4+</sup>	Bi <sup>3+</sup>	covalent	radioactive

**NOTE:**1. Group I through Group IV A

- **Charge of Cation = Group Number**
- Atoms lose all their valence electrons ("s" and "p" electrons)
- Electronic Configuration of cations = Preceding Noble Gas electronic Configuration

2. Group III A through Group V A

- **Charge of Cation = Group Number - 2**
- **Atoms lose only their "p" valence electrons ("s" electrons are kept)**

Electronic Configuration of cations = **Preceding Pseudo Noble Configuration + ns<sup>2</sup>**

3. Group V A through Group VII A

- **Charge of Anion = Group Number - 8**

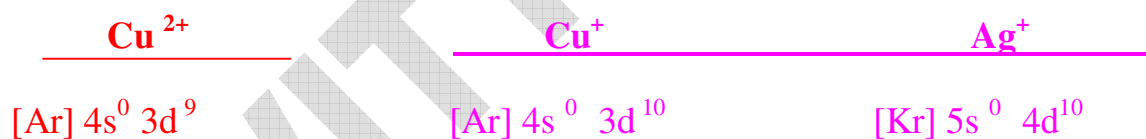
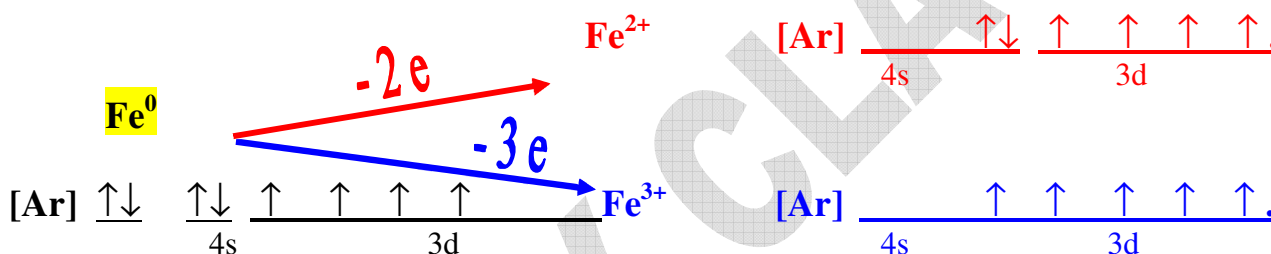
Atoms gain electrons to achieve stable octet (or doublet in the case of H)

Electronic Configuration of anions = Noble Gas electronic Configuration following the element.

## CATIONS OF TRANSITION METALS

- These ions are very often colored

Loss of electrons	Ionic Charge	Examples
The "ns" electrons are lost first (most common)	2+	Zn <sup>2+</sup> Cd <sup>2+</sup>
In some cases in addition to the "s" electrons, some "d" electrons from an inner shell (n-1) may be lost.	2+ and/or larger than 2+	Fe <sup>2+</sup> Fe <sup>3+</sup>
Unusual charges are due to irregular electronic configurations associated with the stability of half-filled or filled subshells	2+ and/or 1+	Cu <sup>2+</sup> Cu <sup>+</sup> Ag <sup>+</sup>

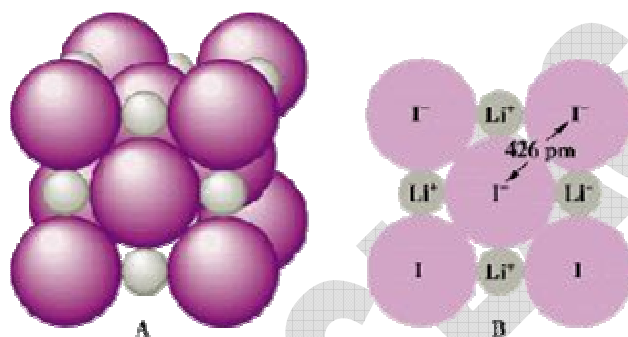


## IONIC RADII

### Ionic Radius

- is the radius of the spherical region around the nucleus within which the electrons are most likely to be found.
- is obtained from known distances between nuclei in ionic crystals (obtained by X-ray diffraction crystallography, a method that makes the positions of the nuclei visible)

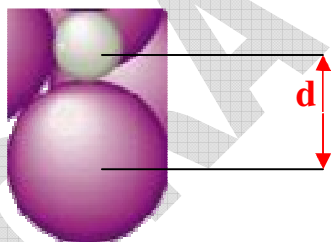
### Determining the Radius of the $I^-$ (iodide) ion in LiI (lithium iodide)



- Distance between the nuclei of two  $I^-$  ions = 426 pm

$$\text{Radius of the } I^- \text{ ion} = \frac{426 \text{ pm}}{2} = 213 \text{ pm}$$

- From this information the radius of the  $Li^+$  ion can be determined from the experimentally determined **distance between  $Li^+$  ion and the  $I^-$  ion.**



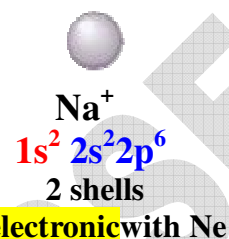
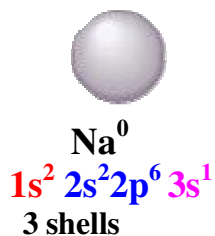
$$\text{Radius of } Li^+ = d (Li^+ - I^-) - 213 \text{ pm}$$

<b>RADII OF CATIONS</b>
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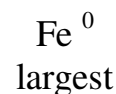
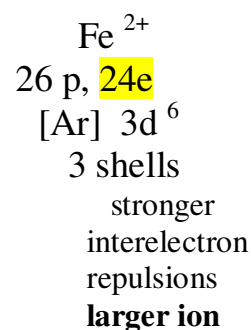
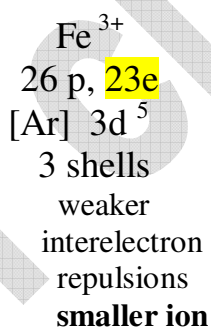
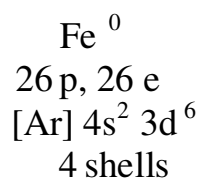
Radii of atoms

&gt;

Radii of Cations



- **Isoelectronic** = species having the same number of electrons and electron configuration.
- Arrange the following species in order of increasing size:

**Conclusion:**

- For cations of the same element (same number of protons):

**The larger the positive ionic charge, the smaller the cation**

- Reason: Larger positive ionic charge means fewer electrons

### Size of Cations in The Periodic Table (Main Group Elements)

	IA	IIA	IIIA	
2	<b>Li<sup>+</sup></b>			<b>(isoelectronic with He)</b>
3	<b>Na<sup>+</sup></b> 11 protons	<b>Mg<sup>2+</sup></b> 12 protons	<b>Al<sup>3+</sup></b> 13 protons	<b>(isoelectronic with Ne)</b>
4	<b>K<sup>+</sup></b> 19 protons	<b>Ca<sup>2+</sup></b> 20 protons	<b>Ga<sup>3+</sup></b>	<b>(isoelectronic with Ar)</b>
5	<b>Rb<sup>+</sup></b> 37 protons	<b>Sr<sup>2+</sup></b> 38 protons	<b>In<sup>3+</sup></b>	<b>(isoelectronic with Kr)</b>
6	<b>Cs<sup>+</sup></b> 55 protons	<b>Ba<sup>2+</sup></b> 56 protons	<b>Tl<sup>3+</sup></b>	<b>(isoelectronic with Xe)</b>

**Number of protons decreases**

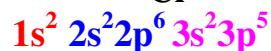
**Ionic Radius Increases**

<b>RADII OF ANIONS</b>
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Radii of Atoms

&lt;

Radii of Anions

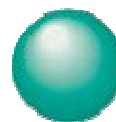
 $\text{Cl}^0$ 

3 shells

17 p, 17 e

Nuclear Charge = + 17

Holds 17 electrons strongly

 $\text{Cl}^-$ 

isoelectronic with Ne

3 shells

17p, 18 e

Nuclear Charge = + 17

Holds 18 electrons less strongly

- Arrange the following species in order of increasing size:

$\text{O}^0$   
8 p, 8e  
[He]  $2s^2 2p^4$   
2 shells  
some  
interelectron  
repulsions

$\text{O}^{2-}$   
8 p, 10e  
[He]  $2s^2 2p^6$   
2 shells  
strongest  
interelectron  
repulsions  
**larger ion**

$\text{O}^-$   
8 p, 9e  
[He]  $2s^2 2p^5$   
2 shells  
stronger  
interelectron  
repulsions  
**smaller anion**

$\text{O}^0$   
smallest

&lt;

 $\text{O}^-$ 

&lt;

$\text{O}^{2-}$   
largest


**Conclusion:**

- For anions of the same element (same number of protons):

**The larger the negative ionic charge, the larger the anion**

- Reason: Larger negative ionic charge means more electrons.

Size of Anions in the Periodic Table

<b>Radius</b>				
				
<b>Number of protons</b>				
	VA	VIA	VIIA	Noble Gases
2	<b>N<sup>3-</sup></b> 7 protons	<b>O<sup>2-</sup></b> 8 protons	<b>F<sup>-</sup></b> 9 protons	<b>(isoelectronic with Ne)</b>
3	<b>P<sup>3-</sup></b> 15 protons	<b>S<sup>2-</sup></b> 16 protons	<b>Cl<sup>-</sup></b> 17 protons	<b>(isoelectronic with Ar)</b>
4	<b>As<sup>3-</sup></b> 33 protons	<b>Se<sup>2-</sup></b> 34 protons	<b>Br<sup>-</sup></b> 35 protons	<b>(isoelectronic with Kr)</b>
5		<b>Te<sup>2-</sup></b> 52 protons	<b>I<sup>-</sup></b> 53 protons	<b>(isoelectronic with Xe)</b>

Number of protons

Atomic Number