Chem 103, Section F0F
Unit III - Combining Atoms to Make Compounds, Part I: Ionic Compounds
Lecture 10

- The atomic properties and chemical bonds
- Ionic bonding.
- Naming ionic compounds and representing them with chemical formulas

Lecture 10 - Chemical Bonding and Ionic Compounds

- Reading in Silberberg
  - Chapter 9, Section 1 *Atomic Properties and Chemical Bonds*
  - Chapter 2, Section 7 *Compounds: Introduction to Bonding*
  - Chapter 9, Section 2 *The Ionic Bonding Model*
  - Chapter 2, Section 8 (pp. 64-70) *Ionic Compounds: Formulas and Names*

Lecture 10 - Introduction

We will look first at chemical bonding
- With a focus on ionic bonds and ionic compounds.

We will then look at the source of energy needed to form ionic compounds.

We will also look into how chemical formulas are used to represent compounds
- And the systematic way of naming ionic compounds.

Lecture 10 - Atomic Properties and Chemical Bonds

Nearly all naturally occurring substances consist of atoms or ions bonded to other atoms or ions.
- There are very few elements that occur uncombined in nature
- Most exist combined with other elements to form compounds.

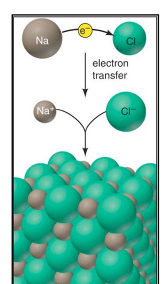
Lecture 10 - Atomic Properties and Chemical Bonds

Chemical bonding allows atoms to lower their energy.

Types of chemical bonding include:
- Ionic bonding (metals with nonmetals)
- Covalent bonding (nonmetals with nonmetals)
- Metallic bonding (metals with metals)

Lecture 10 - Atomic Properties and Chemical Bonds

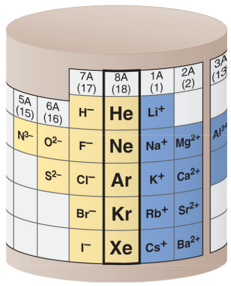
Ionic bonding
- Occurs when metal atoms transfer valance electrons to nonmetal atoms.
  - The resulting attraction of the opposing charges leads to the formation of ionic solids.



A Ionic bonding

Lecture 10 - Atomic Properties and Chemical Bonds

Ionic bonding
- For monatomic ions
  - The metals lose and nonmetals gain that number of electrons that will make them *isolectronic* with the nearest noble gas.

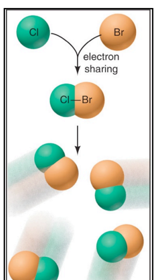


Lecture 10 - Question

What monatomic ions do barium (Z = 56) and sulfur (Z = 16) form?

Lecture 10 - Atomic Properties and Chemical Bonds

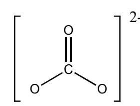
Covalent bonding
- Occurs most commonly between nonmetals.
  - In covalent molecules, the atoms share pairs of valence electrons.



B Covalent bonding

Lecture 10 - Atomic Properties and Chemical Bonds

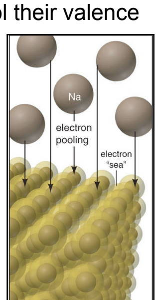
Covalent bonding
- We will come back later to look a covalent molecules
- However, when considering ionic compounds we will encounter polyatomic ions
  - These are a group of atoms that are covalently bonded to one another, but have a net positive or negative charge.
    - They therefore also participate in ionic bonding.



Carbonate Ion (CO_3^{2-}

Lecture 10 - Atomic Properties and Chemical Bonds

Metallic bonding
- Occurs when many metal atoms pool their valence electrons.
  - This pool of electrons holds the metal atoms together.



C Metallic bonding

Lecture 10 - Question

Are molecules of $MgBr_2$ present in a sample of $MgBr_2$?

Explain

Lecture 10 - Question

Are ions present in a sample of P_4O_6?

Explain

Lecture 10 - Question

Predict the type of bonding, *ionic*, *covalent*, or *metallic*, you would predict for each of the following:

A) $CsF_{(s)}$

B) $H_2S_{(g)}$

C) $N_2O_{(g)}$

D) $CaO_{(s)}$

E) $BrO_{2(g)}$

F) $Cr_{(s)}$

Lecture 10 - Atomic Properties and Chemical Bonds

Lewis dot structures are used to represent the valence electrons for an atom

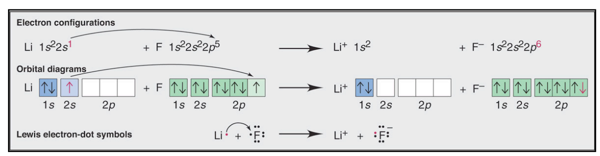
- They are usefull for illustrating the formation of both ionic and covalent compounds.

Period	1A(1)	2A(2)	3A(13)	4A(14)	5A(15)	6A(16)	7A(17)	8A(18)
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
2	Li	Be	B	C	N	O	F	Ne
3	Na	Mg	Al	Si	P	S	Cl	Ar

Lecture 10 - Atomic Properties and Chemical Bonds

Lewis dot structures are used to represent the valence electrons for an atom

- They are useful for illustrating the formation of both ionic and covalent bonded molecules.



Forming lithium fluoride from lithium metal and fluorine gas

Lecture 10 - Question

Draw a Lewis electron-dot symbol for

A) Ba

B) Kr

C) Se

D) P

E) Rb

Lecture 10 - Question

Give the group number an general elecron configuration of an element with each electron-dot symbol

A) :Ẍ·

B) ·X·

C) ·Ẋ·

D) ·X

Lecture 10 - Question

Identify the main group to which X belongs in each ionic compound formula:

A) XPO_4

B) MgX_2

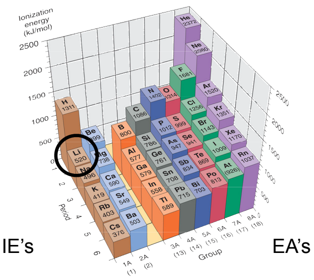
C) Na_3X

D) Al_2X_3

Lecture 10 - Ionic Bonding

Ion formation by itself requires energy

- The negative Electron Affinities (EA's) of the nonmetals are not sufficient to compensate for the positive Ionization Energies (IE's) of the metals.



IE's

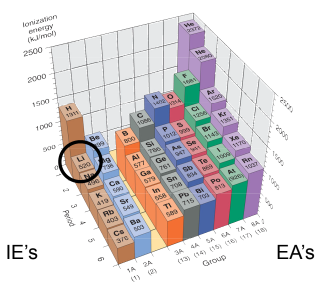
1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (0.0)
Li −59.6	Be ≤0	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)
Na −52.9	Mg ≤0	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)
K −48.4	Ca −2.37	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr −5.03	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)
Cs −45.5	Ba −13.95	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

EA's

Lecture 10 - Ionic Bonding

Ion formation by itself requires energy

- For the formation of lithium fluoride
  - IE_1 (lithium) + EA_1 (fluoride) = 520 - 328 = 192 kJ/mol



IE's

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EA's

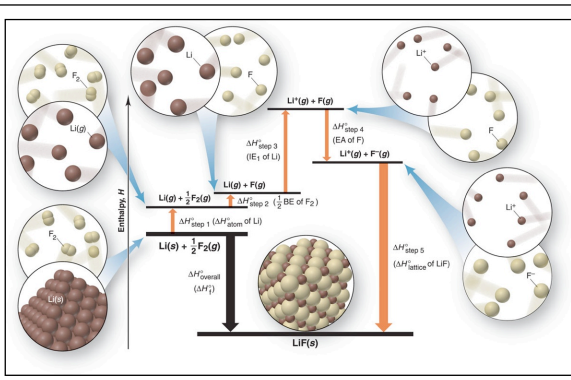
Lecture 10 - Ionic Bonding

Ion formation by itself requires energy

- The required energy for forming ionic compounds comes from the lattice energy.
  - The lattice energy is the energy absorbed when an ionic solid separates into gaseous ions.
    - It is very large and is the major reason that ionic solids exist.
  - The lattice energy depends on ionic size and charge and can be calculated using **Hess's Law** in the **Born-Haber cycle**.

- **Hess's Law** - *the enthalpy or energy change of an overall process is the sum of the enthalpy or energy changes of its individual steps.*

Lecture 10 - Ionic Bonding

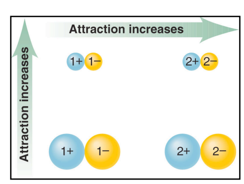


Lecture 10 - Ionic Bonding

Ion formation by itself requires energy

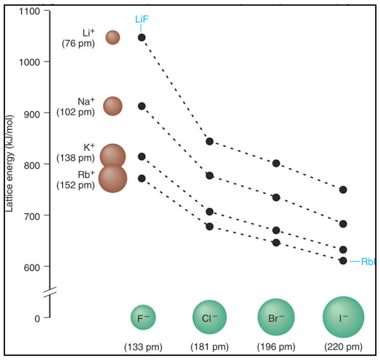
- The required energy for forming ionic compounds comes from the lattice energy.
  - The ionic interactions can be described by Coulomb's law

$$E_{\text{electrostatic}} = \frac{\text{charge A} \times \text{charge B}}{\text{distance}}$$



Lecture 10 - Ionic Bonding

The smaller the ions, the stronger the interaction.



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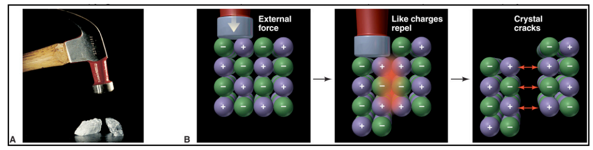
Lecture 10 - Ionic Bonding

The ionic bonding model pictures oppositely charged ions held together by strong electrostatic interactions.

- This model explains why ionic solids tend to fracture when struck.



As layer slip past on another, charge repulsion leads to a fracture

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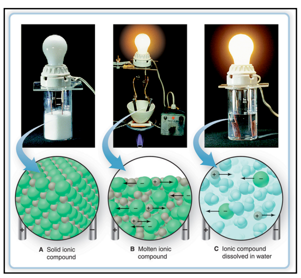
Lecture 10 - Ionic Bonding

The ionic bonding model pictures oppositely charged ions held together by strong electrostatic interactions.

- This model explains why they conduct electricity only when melted or dissolved.



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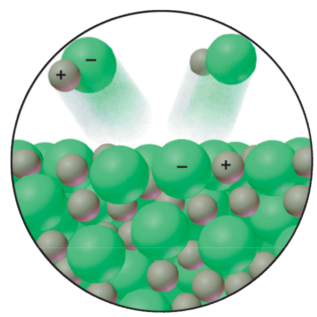
Lecture 10 - Ionic Bonding

The ionic bonding model pictures oppositely charged ions held together by strong electrostatic interactions.

- Gaseous ion pairs require very high temperatures.



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Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Chemical formulas describe

- the simplest atom ratio(empirical formula)
- the actual atom number (molecular formula)
- The atom arrangement (structural formula) of one unit of a compound.

Type of Bonding	*Name*	*Empirical Formula*	*Molecular Formula*	*Structural Formula*
Ionic	*Magnesium Chloride*	*$MgCl_2$ (formula unit)*		
Covalent	*Hydrogen Peroxide*	*HO*	*H_2O_2*	*H-O-O-H*

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Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- For metal cations that can have more than one charge state, Roman numerals are used to indicate the charge.
- Monatomic anions add the -ide suffix to the root name of the atom.
- Oxoanions (polyatomic) use suffixes and sometime prefixes to indicate the relative number oxygen atoms.
- Some ionic compounds form crystal lattices which contain water molecules.
  - These crystals are called hydrates.
  - The number of waters is included in the chemical formula and in the name.
- Acid names are based on the anion names.

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Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- For metal cations fromed from Groups 1A, 2A and 3A, the charge is understood.

Table 2.3 Common Monatomic Ions*

Charge	Formula	Name
Cations		
1+	H^+	hydrogen
	Li^+	lithium
	Na^+	sodium
	K^+	potassium
	Cs^+	cesium
	Ag^+	silver
2+	Mg^{2+}	magnesium
	Ca^{2+}	calcium
	Sr^{2+}	strontium
	Ba^{2+}	barium
	Zn^{2+}	zinc
	Cd^{2+}	cadmium
3+	Al^{3+}	aluminum

Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- For metal cations that can have more than one charge state, Roman numerals are used to indicate the charge.

Table 2.4 Some Metals That Form More Than One Monatomic Ion*

Element	Ion Formula	Systematic Name	Common (Trivial) Name
Chromium	Cr^{2+}	chromium(II)	chromous
	Cr^{3+}	**chromium(III)**	chromic
Cobalt	Co^{2+}	cobalt(II)	
	Co^{3+}	cobalt(III)	
Copper	Cu^+	**copper(I)**	cuprous
	Cu^{2+}	**copper(II)**	cupric
Iron	Fe^{2+}	**iron(II)**	ferrous
	Fe^{3+}	**iron(III)**	ferric
Lead	Pb^{2+}	**lead(II)**	
	Pb^{4+}	lead(IV)	
Mercury	Hg_2^{2+}	mercury(I)	mercurous
	Hg^{2+}	**mercury(II)**	mercuric
Tin	Sn^{2+}	**tin(II)**	stannous
	Sn^{4+}	tin(IV)	stannic

*Listed alphabetically by metal name; those in **boldface** are most common.

Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- Monatomic anions add the *-ide* suffix to the root name of the atom.

Table 2.3 Common Monatomic Ions*

Charge	Formula	Name
Anions		
1−	H^-	hydride
	F^-	**fluoride**
	Cl^-	**chloride**
	Br^-	**bromide**
	I^-	**iodide**
2−	O^{2-}	**oxide**
	S^{2-}	**sulfide**
3−	N^{3-}	nitride

*Listed by charge; those in **boldface** are most common.

Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- Oxoanions (polyatomic) use suffixes and sometime prefixes to indicate the relative number oxygen atoms.

Table 2.5 Common Polyatomic Ions*

Formula	Name	Formula	Name
Cations		MnO_4^-	**permanganate**
NH_4^+	**ammonium**	CO_3^{2-}	**carbonate**
H_3O^+	**hydronium**	HCO_3^-	**hydrogen carbonate** (or **bicarbonate**)
Anions			
CH_3COO^- (or $C_2H_3O_2^-$)	**acetate**	CrO_4^{2-}	chromate
		$Cr_2O_7^{2-}$	**dichromate**
CN^-	cyanide	O_2^{2-}	peroxide
OH^-	**hydroxide**	PO_4^{3-}	**phosphate**
ClO^-	hypochlorite	HPO_4^{2-}	hydrogen phosphate
ClO_2^-	chlorite		
ClO_3^-	**chlorate**	$H_2PO_4^-$	dihydrogen phosphate
ClO_4^-	**perchlorate**		
NO_2^-	nitrite	SO_3^{2-}	sulfite
NO_3^-	**nitrate**	SO_4^{2-}	**sulfate**
MnO_4^-	**permanganate**	HSO_4^-	hydrogen sulfate (or bisulfate)

*Boldface ions are most common.

No. of O atoms	Prefix	Root	Suffix
	per	*root*	ate
		root	ate
		root	ite
	hypo	*root*	ite

Lecture 10 - Chemical Formulas and Names for Ionic Compounds

Ionic compounds are named by listing the cation first followed by the anion.

- Oxoanions (polyatomic) use suffixes and sometime prefixes to indicate the relative number oxygen atoms.
- The charge on a polyatomic ion is less straight forward to predict.
  - Oxoanions formed from the same non metal tend to have the same charge
    - nitrite (NO_2^-), nitrate (NO_3^-)
    - sulfite (SO_3^{2-}), sulfate (SO_4^{2-})
    - hypochlorite (ClO^-), chlorite (ClO_2^-), chlorate (ClO_3^-), perchlorate (ClO_4^-)

Lecture 10 - Question

Give the name and formula of the compound formed from the following elements

A) $_{12}$L and $_{9}$M

B) $_{17}$L and $_{38}$M

Lecture 10 - Question

Give the systematic names for the following compounds

A) $CuSO_4 \cdot 5H_2O$
B) CoO
C) $Ba_3(PO_4)_2$
D) $Pb(NO_3)_4$
E) $Pb(NO_3)_2$

Unit III - Up Next

Unit IV - Chemical Bookkeeping: Stoichiometry

- The concept of a mole
- Determining the chemical formulas for a compound
- Writing and balancing a chemical equation for a chemical reaction

The End