

Ionic and Covalent Bonds

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Section #2

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The periodic table includes a legend for Manganese (Mn) with the following details:

- Group Classification: VII
- Atomic Number: 25
- Symbol: Mn
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
- Atomic Weight: 54.938025
- Name: Manganese

Below the table, a note states: "Atomic weights are based on $^{12}\text{C} = 12$ and conform to the 1995 IUPAC reported values. Number in () indicates the isotope of longest half-life."

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Compounds: Introduction to Bonding

- The noble gases - helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn) – occur in air as separate atoms
- Other molecules: O₂, N₂, H₂, S₈, C, Cu, Ag, Pt, Au
- Elements combine to form a compound in 2 general ways:
 - Transferring electrons from the atoms of one element to those of another to form *ionic compounds*
 - Sharing electrons between atoms of different elements to form *covalent compounds*.
- These processes generate **chemical bonds**, the forces that hold the atoms of elements together in a compound.

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Formation of Ionic Compounds

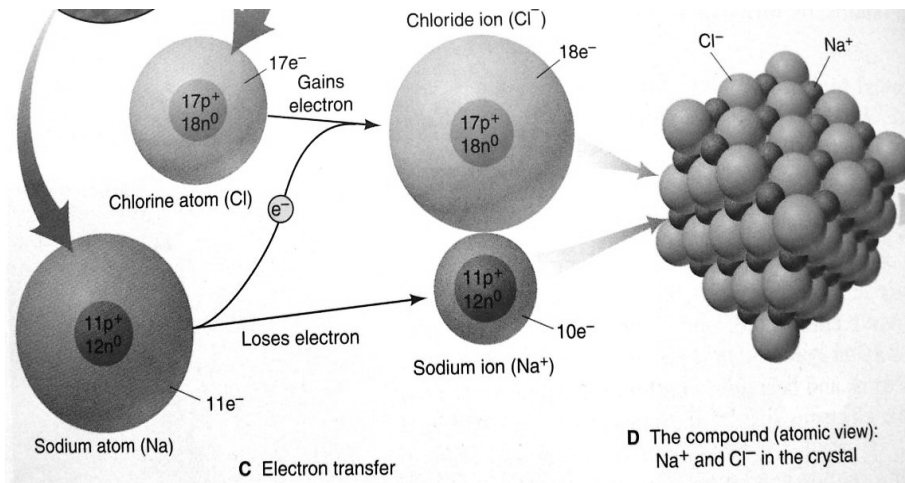
- **Ionic compounds** are composed of ions – charged particles that form when an atom (or small group of atoms) gains or loses one or more electrons.
- **Binary ionic compound**: one composed of just two elements, typically forms when a metal reacts with a nonmetal.
- Each metal atom loses a certain number of its electrons and becomes a **CATION**, a positively charged ion.
- The non metal atoms gain the electron loss by the metal atoms and becomes **ANIONS**, negatively charged ion.
- The Cations and Anions attract each other through electrostatic forces and form the ionic compounds
- All binary ionic compounds are solid
- A cation or anion derived from a single atom is called a **MONATOMIC ION**

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Example of the formation of ionic compound -- NaCl



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- Sodium atom loses 1 electron and forms a sodium cation (Na^+).
- A Chlorine atom gains the electron and becomes a chloride anion (Cl^-).
- The oppositely charge ions (Na^+ and Cl^-) attract each other
- The resulting solid aggregation is a regular array of alternating Na^+ and Cl^- ions that extends in all three dimensions

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Strength of Ionic Bonding

- depends on net strength of attractions and repulsions and is described by Coulomb's Law
- *The energy of attraction (or repulsion) between two particles is directly to the product of the charges and inversely proportional to the distance between them*
- *Ions with higher charges attract (or repel) each other more strongly than ions with lower charges*

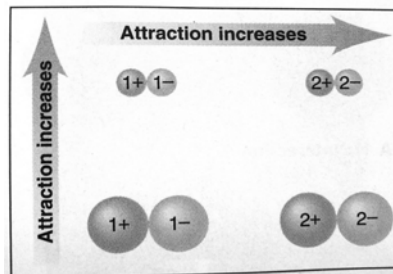
$$\text{Energy} \propto \frac{\text{charge 1} \times \text{charge 2}}{\text{distance}}$$

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- Smaller ion attract (or repel) each other more strongly than larger ions, because their charges are closer together.
- Ionic compounds are neutral, that is, they possess no net charge
- Figure 2.13



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Why Why Why Why ??????

- Why does each sodium atom give up only 1 of its 11 electrons?
- Why doesn't each chlorine atom gain two electrons, instead of just one?
- ==> PERIODIC TABLE will answer
- In general, metals lose electrons and non metals gain electrons to form ions with the same number of electrons as in the nearest noble gas [GROUP 8A(18)]

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- Noble gases have a stability (low reactivity) that is related to their number of electrons
- Sodium ion (11e-) can attain the stability of neon (10e-) by losing one electron
- By gaining one electron, a chlorine atom (17e-) attains the stability of argon (18e-).
- Thus, when an element located near a noble gas forms a monatomic ion, it gains or loses enough electrons to attain the same number as that noble gas

5A (15)	6A (16)	7A (17)	8A (18)	1A (1)	2A (2)	3A (13)
		H ⁻	He	Li ⁺		
N ³⁻	O ²⁻	F ⁻	Ne	Na ⁺	Mg ²⁺	Al ³⁺
	S ²⁻	Cl ⁻	Ar	K ⁺	Ca ²⁺	
		Br ⁻	Kr	Rb ⁺	Sr ²⁺	
		I ⁻	Xe	Cs ⁺	Ba ²⁺	

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In Periodic Table

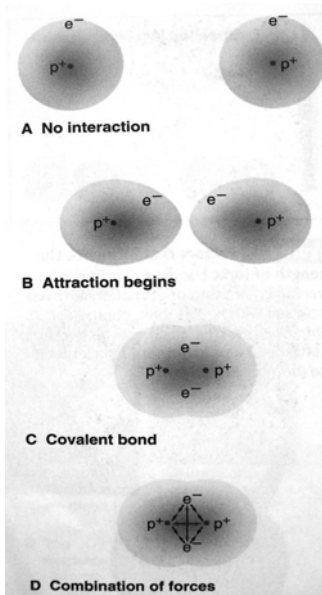
- The elements in Group 1A(1) lose one electron, those in Group 2A(2) lose two, and aluminum in Group 3A(13) loses three
- The elements in Group 7A(17) gain one electron, oxygen and sulfur in Group 6A(16) gain two, and nitrogen in Group 5A(15) gains three
- Fluorine (F, $Z=9$) has one electron fewer and sodium (Na, $Z=10$) has one electron more than the noble gas neon (Ne, $Z=10$), they form the F^- and Na^+ ions
- Similarly, oxygen (O, $Z=8$) gains two electrons and magnesium (Mg, $Z=12$) loses two to form the O^{2-} and Mg^{2+} ions and attain the same number of electrons as neon

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Formation of Covalent Compounds



- Covalent compounds form when elements share electrons, which usually occurs between nonmetals
- Example: Two Hydrogen atoms (H, $Z=1$):
- The nucleus of each atom attracts the electron of the other atom more and more strongly
- The separated atoms begin to interpenetrate each other
- At some optimum distance between the nuclei, the atoms form a **COVALENT BOND**, where a pair of electrons mutually attracted by the two nuclei

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	1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	H ₂							
2				N ₂	O ₂	F ₂		
3				P ₄	S ₈	Cl ₂		
4					Se ₈	Br ₂		
5						I ₂		
6								
7								

Diatomic molecules
 Tetratomic molecules
 Octatomic molecules

- The result is a hydrogen molecule, in which each electron no longer “belongs” to a particular H atom; the two electrons are shared by the two nuclei
- Repulsions between the nuclei and between the electrons also occur, but the net attraction is greater than the net repulsion.
- Other nonmetals that exist as diatomic molecules at room temperature are nitrogen (N₂), oxygen (O₂), and the halogens (fluorine (F₂), chlorine (Cl₂), bromine (Br₂), and iodine (I₂)).
- Phosphorus exists as tetratomic molecules (P₄), and sulfur and selenium as octatomic molecules (S₈ and Se₈).
- At room temperature, covalent compounds may be gases, liquids, or solids

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- Atoms of different elements share electrons to form the molecules of a covalent bond
- Example: Hydrogen Fluoride (HF), Water (H₂O)



- **Most covalent substances consists of molecules** : a cup of water ==> consists of individual water molecules
- **No molecules exist in a sample of ionic compound**: ==> NaCl is a continuous array of oppositely charged sodium and chloride ions, not a collection of individual NaCl molecules

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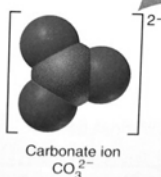
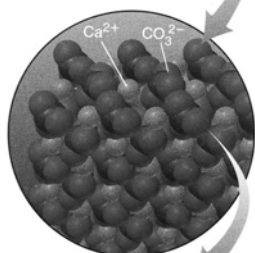
Another Key Difference

- **Covalent bonding:** involves the mutual attraction between two (positively charged) nuclei and the two (negatively charged) electrons that reside between them.
- **Ionic bonding:** involves the mutual attraction between positive and negative ions

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tion Name

Polyatomic Ions

- Many ionic compounds contain polyatomic ions, which consist of two or more atoms **bonded Covalently** and have a net positive or negative charge
- **Calcium carbonate** is a 3D array of monatomic calcium cations and polyatomic carbonate anions.
- As the bottom structure shows, each **carbonate** ion consists of four covalently bonded atoms

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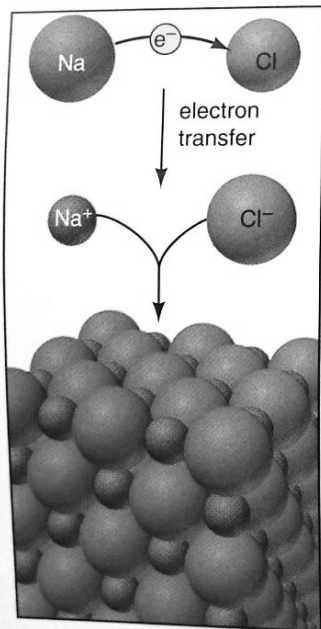
Three Types of Chemical Bonding

Key:

- Metals
- Nonmetals
- ◻ Metalloids

1A (1)	2A (2)											3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)																												
Li	Be											B	C	N	O	F	Ne																												
Na	Mg	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8) (9) (10)			1B (11)	2B (12)	Al	Si	P	S	Cl	Ar																												
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr																												
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe																												
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn																												
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	112		114		116																														
		<table border="1"> <tr> <td>Ce</td><td>Pr</td><td>Nd</td><td>Pm</td><td>Sm</td><td>Eu</td><td>Gd</td><td>Tb</td><td>Dy</td><td>Ho</td><td>Er</td><td>Tm</td><td>Yb</td><td>Lu</td> </tr> <tr> <td>Th</td><td>Pa</td><td>U</td><td>Np</td><td>Pu</td><td>Am</td><td>Cm</td><td>Bk</td><td>Cf</td><td>Es</td><td>Fm</td><td>Md</td><td>No</td><td>Lr</td> </tr> </table>																Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu																																
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr																																

- Why atoms bond at all ?
- *Bonding lowers the potential energy between positive and negative particles.*
- *Just as the electron configuration and the strength of the nucleus-electron attraction determine the properties of an atom,*
- *the type and strength of chemical bonds determine the properties of a substance*



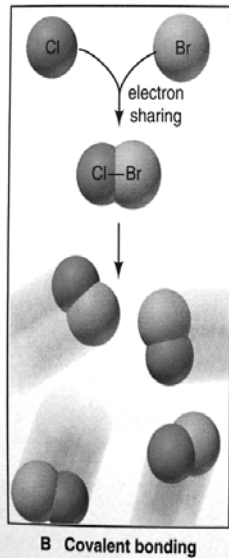
A Ionic bonding

name

- (1). ***Metal with nonmetal***: electron transfer and ionic bonding
- Typically as IONIC, BONDING
- Such difference occur between reactive metals (1A(1) and 2A(2)) and nonmetals (7A(17)) and the top of Group 6A (16).
- The metal atom loses its one or two valence electrons, whereas the nonmetal atom gains the electron(s).
- Electron transfer from metal to nonmetal occurs, and each atom forms an ion with a noble gas electron configuration

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- The electrostatic attraction between these positive and negative ions draws them into the 3D array of an ionic solid
- whose chemical formula represents the cation-to-anion ratio (empirical formula)

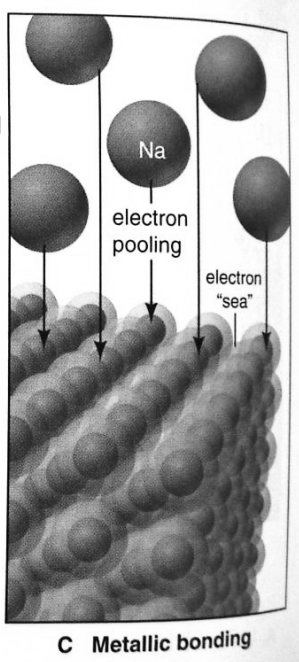


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- **Nonmetal with nonmetal:** electron sharing and covalent bonding.
- when two atoms have a small difference in their tendencies to lose or gain electrons, we observe *electron sharing* and *covalent bonding*
- This type of bonding most commonly occurs between nonmetal atom, Although a pair of metal atoms can sometimes form a covalent bond also
- The attraction of each nucleus for the valence electrons of the other draws the atoms together.
- A shared electron pair is considered to be “localized” between the atoms because it spends most of its time there, linking them in a covalent bond of a particular length and strength.

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C Metallic bonding

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- (3). **Metal with metal:** *electron pooling* and *metallic bonding*
- In general, metal atoms are relatively large, and their view outer electrons are well shielded by filled inner levels
- Thus, they lose outer electrons comparatively easily but do not gain them very readily.
- These properties lead large numbers of metal atoms to share their valence electrons, but in a way that differs from covalent bonding

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- All the metal atoms in a sample *pool* their valence electrons into an evenly distributed “*sea*” of electrons that “flows” between and around the metal-ion cores (nucleus plus inner electrons) and attracts them, thereby holding them together
 - Unlike the localized electrons in covalent bonding, electrons in metallic bonding are “*delocalized*”, *moving freely throughout the piece of metal*.
 - In other word: Metallic bonding occurs when many metal atoms pool their valence electrons in a delocalized electron sea that holds all the atoms together.

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Notes

- You cannot always predict the bond type solely from the elements' positions in the periodic table
- All binary compounds contain a metal and a nonmetal, but all metals do not form binary ionic compounds with all nonmetals
- Example: metal beryllium (Group 2A(2)) combines with the nonmetal chlorine (Group 7A(17)), the bonding fits the covalent model better than the ionic model.

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LEWIS ELECTRON-DOT SYMBOLS

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

- The element symbol represents the nucleus and the inner electrons
- The dots around it represent valence electrons, either paired or unpaired
- The number of unpaired dots indicates the number of electrons a metal atom loses, or the number of nonmetal atom gains, or the number of covalent bonds a nonmetal atom usually forms

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Lewis Symbol

- Note its **A-group number (1A to 8A)**, which gives the number of valence electrons
- Place one dot at a time on the four sides (top, right, bottom, left) of the element symbol
- Keep adding dots, pairing the dots until all are used up
- The placement of dots is not important
- The number and pairing of dots provide information about an element's bonding behavior:
 - For a metal, the total number of dots is the maximum number of electrons an atom loses to form a cation
 - For a nonmetal, the number of unpaired dots is the number of electrons that become paired either through electron gain or through electron sharing.
 - Thus the number of unpaired dots equals either the number of electrons a nonmetal atom gains in becoming an anion or the number of covalent bonds it usually forms

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Lewis : ==> OCTET RULE

- when atoms bond, they lose, gain, or share electrons to attain a filled outer level of eight (or two) electrons
- The octet rule holds for nearly all of the compounds of Period 2 elements and a large number of others as well

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