CHEMICAL BONDING: IONIC BONDS

Chemical bonding is a fundamental concept in chemistry that explains how atoms bond to form compounds and molecules. Understanding these bonds helps explain the properties of different substances and how they interact in chemical reactions.

Noble Gases

Noble gases have low chemical reactivity and are monotonic, colorless, odorless, and nonflammable. Noble gases have little tendency to receive or lose electrons due to their complete valence electron shells, so they are incredibly stable and unlikely to form chemical bonds.

These elements are found at the rightmost end of the periodic table. They belong to Group 18, composed of the following elements:

Element	Symbol	Atomic Number	Electrons per shell
Helium	He	2	2 **duplet*
Neon	Ne	10	2,8
Argon	Ar	18	2,8,8
Krypton	Kr	36	2,8,18,8
Xenon	Xe	54	2,8,18,18,8
Radon	Rn	86	2,8,18,32,18,8

Table 1. The Noble Gases

Duplet and Octet Rule

These chemistry rules describe how atoms achieve stability based on their electron configurations.

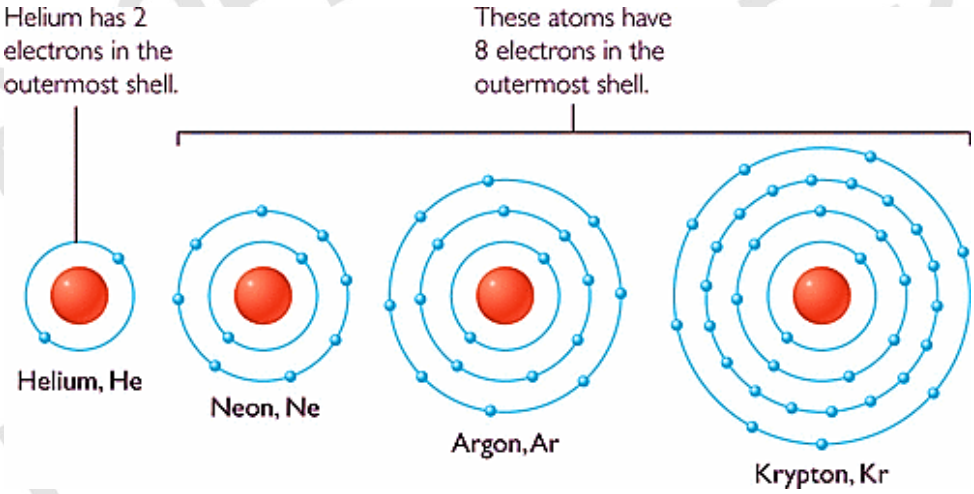


Figure 1.Duplet and octet rule. (n.d.). Brilliant.org. https://brilliant.org/wiki/octet-rule/

Duplet Rule:

Hydrogen and Helium, the elements with the lowest atomic numbers, are stable when they have two (2) electrons in their outermost shell.

Octet Rule:
The octet rule states that atoms with more than two electrons tend to gain, lose, or share electrons to have eight electrons in their outermost shell.

Atoms can achieve duplet or octet electron configuration by:

1. Giving away excess electrons
2. Receiving electrons from another atom to complete the octet
3. Sharing electrons with another atom to form an octet

Ionic Bonds

Ionic bonding occurs when one atom donates an electron to another atom, forming positively and negatively charged ions. The electrostatic attraction between these oppositely charged ions holds them together.

For example, sodium chloride (table salt) is formed by an ionic bond between sodium (Na) and chlorine (Cl).

Periodic Elements and Ionic Charges

Understanding the periodic table and ionic charges helps predict how elements will interact in chemical reactions and the types of compounds they will form. The number of valence electrons indicates the possible bonds and which elements could form ionic bonds with each other.

Main Group Elements

Group	Element	Symbol	Simple Ion Charge	Valence Electron(s)
Group IA	Hydrogen	H	+1	1
	Lithium	Li	+1	1
	Sodium	Na	+1	1
	Potassium	K	+1	1
	Rubidium	Rb	+1	1
	Cesium	Cs	+1	1
Group IIA	Beryllium	Be	+2	2
	Magnesium	Mg	+2	2
	Calcium	Ca	+2	2
	Strontium	Sr	+2	2
	Barium	Ba	+2	2
Group IIIA	Boron	B		3
	Aluminum	Al	+3	3
Group IVA	Carbon	C		4
	Silicon	Si		4
	Tin	Sn	+2	4
	Lead	Pb	+2	4

Group VA	Nitrogen	N	-3 nitride	5
	Phosphorus	P	-3 phosphide	5
	Arsenic	As		5
Group VIA	Oxygen	O	-2 oxide	6
	Sulfur	S	-2 sulfide	6
Group VIIA	Fluorine	F	-1 fluoride	7
	Chlorine	Cl	-1 chloride	7
	Bromine	Br	-1 bromide	7
	Iodine	I	-1 iodide	7

Table 2.The valence electron(s) of the main group elements and its charges in ionic form.

Transition Metals

Transition metals, found in the middle of the periodic table, have unique properties when forming ions. Unlike main group elements, which typically have fixed ionic charges, transition metals can exhibit a variety of charges. This variability arises because they can lose different numbers of electrons from their outermost and next-to-outermost electron shells.

Element	Symbol	Simple Ion Charge
Chromium	Cr	Chromium (II) – +2
		Chromium (III) – +3
Manganese	Mn	+2
Iron	Fe	Ferrous – +2
		Ferric – +3
Cobalt	Co	Cobaltous – +2
		Cobaltic – +3
Nickel	Ni	+2
Copper	Cu	Cuprous – +2
		Cupric – +1
Zinc	Zn	+2
Silver	Ag	+1
Cadmium	Cd	+2
Mercury	Hg	+2

Table 3. The ionic charges of transition metals.

Polyatomic Ions

Polyatomic ions consist of two or more atoms bonded together, carrying an overall charge.

Examples include:

- Nitrate (NO_3^-)
- Sulfate (SO_4^{2-})
- Ammonium (NH_4^+)

Naming compounds that contain polyatomic ions follows specific rules, similar to naming other ionic compounds. Below are common polyatomic ions and their names.

Ion	Name	Ion	Name
NH4+	Ammonium	CO32-	Carbonate
NO2-	Nitrite	HCO3-	Bicarbonate
NO3-	Nitrate	ClO4-	Perchlorate
SO32-	Sulfite	ClO3-	Chlorate
SO42-	Sulfate	ClO2-	Chlorite
HSO4-	Bisulfate	ClO-	Hypochlorite
OH-	Hydroxide	C2H3O2-	Acetate
PO43-	Phosphate	Cr2O72-	Dichromate
HPO42-	Hydrogen phosphate	CrO4-	Chromate
H2PO4-	Dihydrogen phosphate	CN-	Cyanide

Table 4. The Nomenclature of different polyatomic ions

Constructing Lewis Structures

In constructing Lewis Structures, generally follow the steps:

1. Find how many valence electrons (N) are in the molecule, which must be shown on the Lewis Structure using the periodic table. Find the charge, add an electron for every negative charge, and remove an electron for every positive charge.
2. Draw out the molecule's single bonds and initial framework, called the skeleton.
3. Complete the octets around the non-central atoms, i.e., the terminal atoms, by using the lone pairs of electrons.
4. Compare the number of electrons currently depicted to the number needed (N) in the central atom and add electrons to it until the octet is complete.
5. If there are extra lone-pair electrons and the octet rule is not filled for the central atom, use the extra electrons to form double or triple bonds around the central atom.
6. Check the formal charge of each atom (Formal Charge explained below).

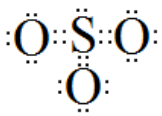
Consider the following guidelines to avoid common mistakes:

- The dots surrounding the chemical symbol are the valence electrons, and each dash represents one covalent bond (consisting of two [2] valence electrons).
- Hydrogen is always terminal in the structure.
- The atom with the lowest ionization energy is typically the central atom in the structure.
- The octet rule means there are eight (8) valence electrons around the atoms, but the maximum is 2 electrons for hydrogen.

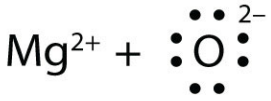
Lewis structures of some ionic and non-ionic compounds:

NaCl – Sodium Chloride (Ionic)

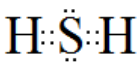
SO_3 – Sodium Trioxide (Covalent)



MgO – Magnesium Oxide (Ionic)



H_2S – Hydrogen Sulfide (Covalent)

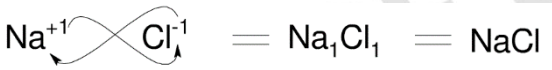


Criss Cross Method for Predicting Formulas

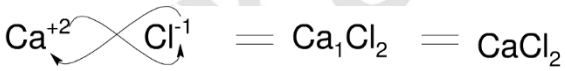
Ions (cations and anions) attract one another to form neutral compounds. Employing the criss-cross method to predict the compound ions ratio can predict the chemical formulas.

The ions sodium (Na^{+1}) and chlorine (Cl^{-1}) combine in a 1:1 ratio. One (1) Na ion for every one (1) Cl ion. The compound that results is NaCl (sodium chloride).

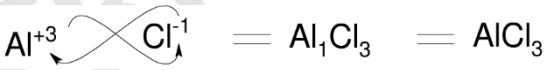
In the criss-cross method, you take the number for the charge on an ion (ignore the positive and negative) and make it the other atoms subscript.



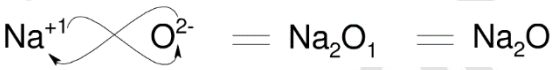
Calcium chloride ($CaCl_2$)



Aluminum chloride ($AlCl_3$)



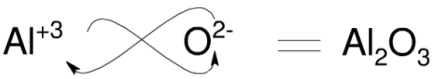
Sodium oxide



Magnesium oxide

In a case like this, where you have two (2) ions that have the same number of an oxidation state, find the least common multiple of the numbers. Always reduce the compound to its simplest form when dealing with ionic compounds.

Aluminum oxide



Check if this method has generated a compound with a net 0 charge. To do this, all we need to do is add up all the positive charges and add that to the sum of all the negative charges.

$$2Al^{+3} = (+3) + (+3) = +6$$
$$3\,O^{2-} = (-2) + (-2) + (-2) = -6$$
$$(+6) + (-6) = 0\ neutral\ compound$$

Therefore, the compound's formula is correct.

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