

Chemistry Honors Study Guide

Test 2 S1

Test date: October 18, 2024

(All links are clickable!)

1 Forming Ionic Compounds

Definitions

- **ion:** Charged particle.
- **cation:** Positively charged particle. All cations are metals.
- **anion:** Negatively charged particle. All anions are either nonmetals or metalloids.
- **ionic compound:** Created from electrostatic attraction between cations and anions.
- **salt:** Another name for an ionic compound.

Ionic Bonding

Atoms form bonds to fill their valence shells. Depending on what is easier for an atom of each element (getting rid of/gaining electrons), they will have a positive or negative charge.

Once bonded, the charges on each ion must cancel out. Like fraction addition, we change the charges to their least common multiple (if they are not already equal).

Example: $2\text{K}^{1+} + \text{O}^{2-} \rightarrow \text{K}_2\text{O}$ **cation first!**

After the bonding occurs, the charge on the resulting particle is neutral.

2 Naming Ionic Compounds

Definitions

- **Type I cation:** Includes groups 1, 2, 13, Ag^+ , Zn^{2+} . Can only have a single charge state.
- **Type II cation:** Transition metals (not Ag or Zn), Pb, Sn. Can have multiple charge states.

Naming Compounds with Type I Cations (Monatomic Anions)

name of cation + anion ending in -ide

Example: Al_2S_3 : Aluminum sulfide Note that the number of atoms present does not influence the naming.

Naming Compounds with Type II Cations (Monatomic Anions)

cation + charge on cation + anion ending in -ide

Example: Fe_2O_3 : iron (III) oxide

In this example, the number of iron atoms does not matter; the number in parentheses is always the charge.

Naming Compounds with Polyatomic Anions

- NO_3^- : nitrate
- ClO_3^- : chlorate
- CO_3^{2-} : carbonate
- SO_4^{2-} : sulfate
- PO_4^{3-} : phosphate

Nick the Camel ate Clam Supper in Phoenix

Number of consonants: Number of oxygen atoms

Number of vowels: Number of negative charges

Relative # of Oxygen	Prefix/suffix	Equation	Name
+1	per...ate	ClO_4^-	perchlorate
-1	-ite	ClO_2^-	chlorite
0	-ate	ClO_3^-	chlorate
-2	hypo...ite	ClO^-	hypochlorite

Example: PO_5^{3-} : perphosphate; Na_3PO_5 : sodium perphosphate

(Note that the charge does not change with the number of oxygen atoms, but when the polyatomic anion bonds with a cation, the resulting compound becomes neutral.)

3 Covalent Compounds

Definitions

- **covalent compound:** A compound bonded covalently (sharing valence electrons).

Naming Binary Covalent Compounds (two elements)

if > 1, number prefix + first element + number prefix + second element + -ide

Example: N_2O_3 : dinitrogen trioxide

Prefixes:

1. **mono-**
2. **di-**
3. **tri-**
4. **tetra-**
5. **penta-**
6. **hexa-**
7. **hepta-**
8. **octa-**

4 Identification and Diagrams

Identifying Ionic vs Covalent Compounds

Ionic compounds...

- start with a **metal** or with NH_4^+ (ammonium).
- form from the electrostatic attraction of **metals with nonmetals/metalloids**.

Covalent compounds...

- are formed between **metals or metalloids**.
- involve the **sharing of electrons**.

Lattice Diagrams

Essential aspects of lattice diagrams:

- separated charges (- not touching +)
- atomic ratio
- atomic symbol
- at least 2 repetitions

Identifying Covalent Lattice Diagrams

Molecular: distinct molecules, not interconnected (such as H_2O)

Network: interconnected network of atoms (such as C (diamond))

Ionic Lattice Diagrams¹

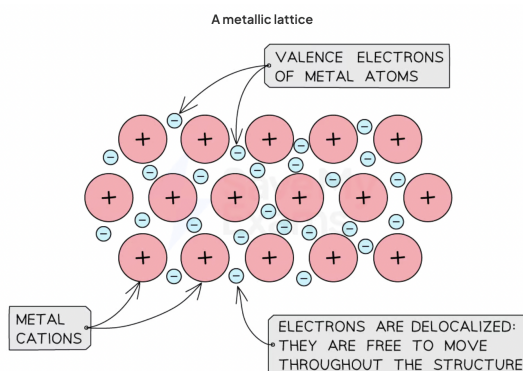
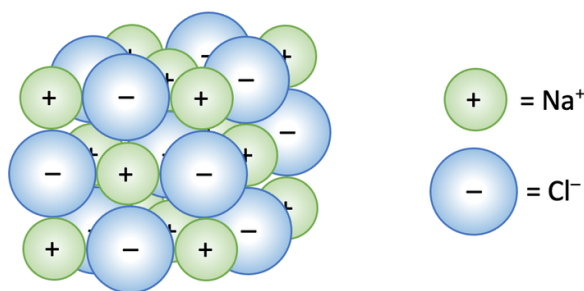
Example: [NaCl \(sodium chloride/table salt\)](#)

Metallic Lattice Diagrams²

(This is the diagram for all metallic bonding. The number of atoms, electrons, and arrow directions do not matter as long as the picture shows a cluster of atoms with moving

¹<https://www.elevise.co.uk/gac2c.html>

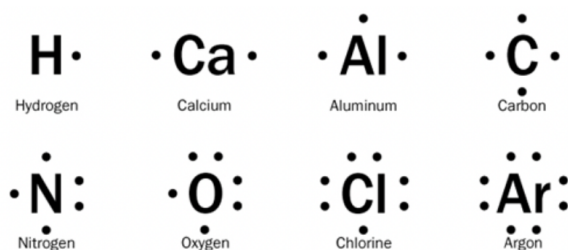
²<https://www.savemyexams.com/ap/chemistry/college-board/22/revision-notes/unit-2-molecular-and-ionic-compound-structure-and-properties/2-4-structure-of-metals-and-alloys/representing-metallic-bonding/>



electrons.)

Lewis Dot Structures

Atomic LDS³



As shown above, the number of valence electrons increases as you go across the periodic table, so the number of dots increase. (Placement is intentional – when drawing the atomic structures, electrons are separated, but in ionic/covalent structures, electrons are drawn in pairs.)

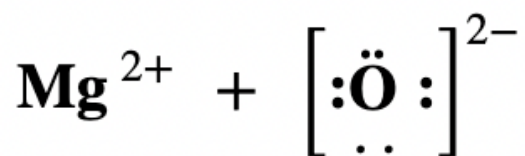
Ionic LDS⁴

- The **cation** always has an **empty valence shell** (no dots).
- The **anion** always has a **filled valence shell** with **brackets** around the atom.
- **Charges should be indicated** in the top right of each atom.
- Charges are **properly separated**.

Example: [MgO \(magnesium oxide\)](#)

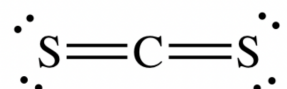
³<https://americanboard.org/Subjects/general-science/bonding-and-atomic-structure/>

⁴<https://scienceready.com.au/pages/lewis-dot-diagram>



Covalent LDS⁵

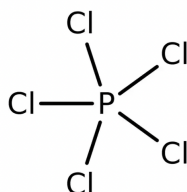
1. Determine the **number of valence electrons**.
2. Determine the **central atom** if there are more than 2 atoms. **The central atom (with the exception of hydrogen) should be the one with the lowest electronegativity.**
3. **Draw the LDS** so that each atom (note exceptions below) has a full valence shell.
4. Replace shared electrons with lines. (**2 electrons = 1 bond = 1 line**).



Example: [CS₂ \(carbon disulfide\)](#)

Exceptions⁶

- **Hydrogen** forms a **full duet** (2 valence electrons).
- **Boron** forms a **full sextet** (6 valence electrons).
- Elements in the **third period or below** may have an **expanded octet** (more than 8 electrons in their valence shell).



Example: [PCl₅ \(phosphorus pentachloride\)](#)

Polyatomic Ions

Polyatomic ions are *covalent structures within an ionic structure*. Draw the Lewis dot structure of a polyatomic ion as a covalent compound, treating it as a single unit, then add the element it bonds ionically with to the diagram. As usual, include brackets when necessary.

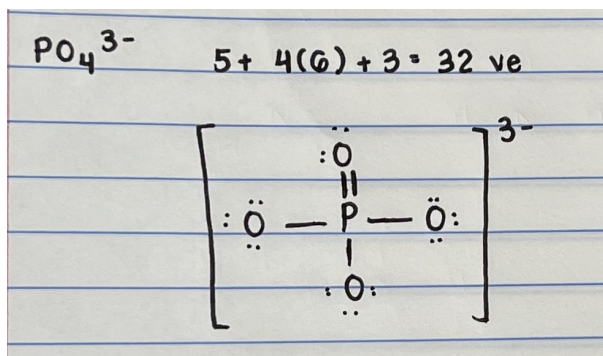
⁵<https://byjus.com/chemistry/carbon-disulfide/>

⁶<https://www.fishersci.ca/shop/products/phosphorus-v-chloride-98-thermo-scientific/p-3257053>

When making LDS with polyatomic ions that have expanded octets, the **formal charge on the central atom must equal 0**. The formula to find formal charge of an atom is as follows:

$$\text{formal charge} = \# \text{ of valence electrons} - \# \text{ of lone pair electrons} - \# \text{ of bonds}$$

Note that the number of lone pair electrons refers to the number of individual electrons, not the number of lone pairs. (Formal charge does not need to be indicated unless the problem states otherwise.)



Example: PO_4^{3-} (phosphate)

5 Periodic Table Trends

Definitions

- **effective nuclear charge:** The force felt by an electron from the nucleus.
- **electronegativity:** The tendency of an atom to attract electrons.
- **ionization energy:** The energy it takes to remove an electron from an atom.
- **atomic radius:** The size of the atom, or, more precisely, the distance from an atom's nucleus to its outermost orbital.

Coulomb's Law

$$F = k \frac{q_1 q_2}{r^2}$$

where:

- F = force felt by electron
- k = Coulomb's constant (not important for this test)
- q_1 = charge on valence electron
- q_2 = charge on nucleus
- r = distance (nucleus to electron)

Electronegativity and Ionization Energy

Electronegativity and ionization energy **increase** as you go **higher** and to the **right** in the periodic table.

Electronegativity increases near the top of the periodic table because the distance from the potential electron and the nucleus is smaller (as described by Coulomb's Law above.) Near the right, electronegativity increases because the atoms grow increasingly close to filling their valence shells, needing to attract extra electrons.

Ionization energy increases near the top because it is harder to remove an electron close to the nucleus (again, as described by Coulomb's Law). Near the right, ionization energy increases because it becomes increasingly difficult to remove electrons from atoms that are close to filling their valence shells.

Atomic Radii

The atomic radius **increases** as you go **lower** and to the **left** in the periodic table. This is because as the number of protons increases (as you go to the right), it will pull the electrons closer, resulting in a smaller atomic radius, and as the number of orbitals increases (as you go lower), the atomic radius increases likewise.