

Balancing equations

Law of conservation of mass

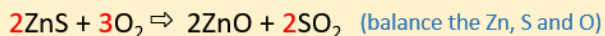
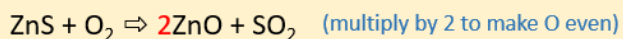
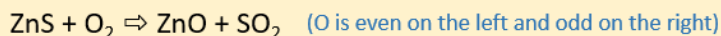
The law of conservation of mass states that, during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants.

Balancing chemical equations

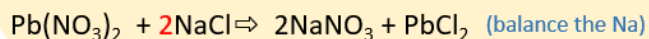
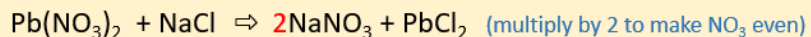
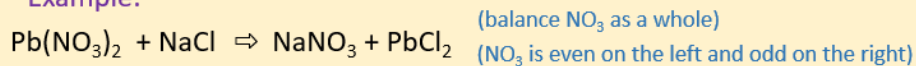
Hints to Balancing Chemical Equations

- Adjust the coefficients not the subscripts.
- Even/Odd – Multiply by 2 to make all even.
- Balance polyatomic ions as a whole.
- Check the balanced equation.

Example:



Example:



Ionic Bonding

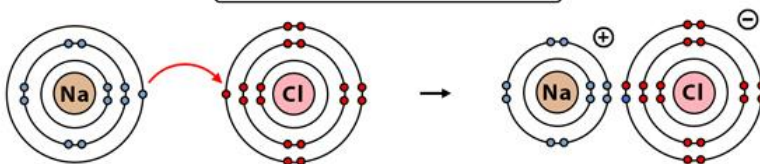
Formation

Positive and negative ions form when a metal reacts with a non-metal, by transferring electrons. The oppositely charged ions are strongly attracted to each other, forming ionic bonds. The attraction is based upon the octet rule, where each atom aims to attain stability through having a fuller valence electron shell

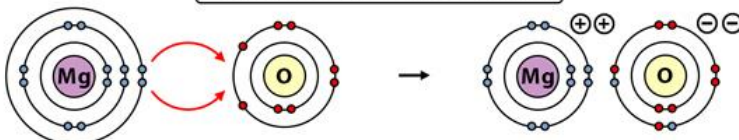
Illustrations

Ionic Bond Examples

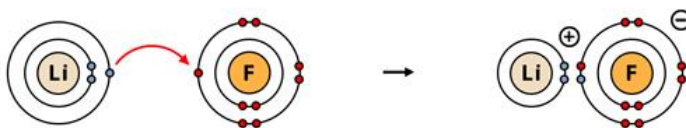
1. Sodium chloride (NaCl)



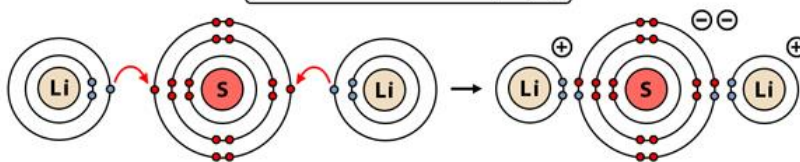
2. Magnesium oxide (MgO)



3. Lithium fluoride (LiF)



4. Lithium sulfide (Li₂S)



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Formula of ionic compounds

The formula for an ionic compound must contain the same number of positive and negative charges so that the charges are balanced and it is neutral overall. Here are some examples.

Example 1

Sodium chloride contains Na⁺ and Cl⁻ ions:

- this is one positive charge and one negative charge
- the charges are balanced
- so the formula is **NaCl**

Example 2

Magnesium oxide contains Mg²⁺ and O²⁻ ions:

- this is two positive charges and two negative charges
- the charges are balanced
- so the formula is **MgO**

Example 3

Aluminium oxide contains Al³⁺ and O²⁻ ions:

- this is three positive charges and two negative charges
- to balance, we need two Al^{3+} ions and three O^{2-} ions
- so the formula is Al_2O_3

Covalent bonding

Formation

A covalent bond forms when two non-metal atoms share a pair of electrons. The electrons involved are in the outer shells of the atoms. An atom that shares one or more of its electrons will complete its outer shell.


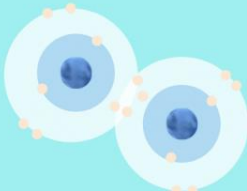
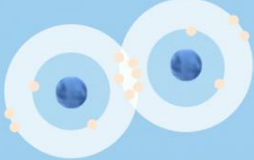
Single, double and triple bonds

a single covalent bond is when two atoms share a single pair of electrons. Represented by a single line (–).

a double covalent bond is when two atoms share two pairs of electrons. Represented by a double line (=).

a triple covalent bond is when two atoms share three pairs of electrons. Represented by a triple line (\equiv).

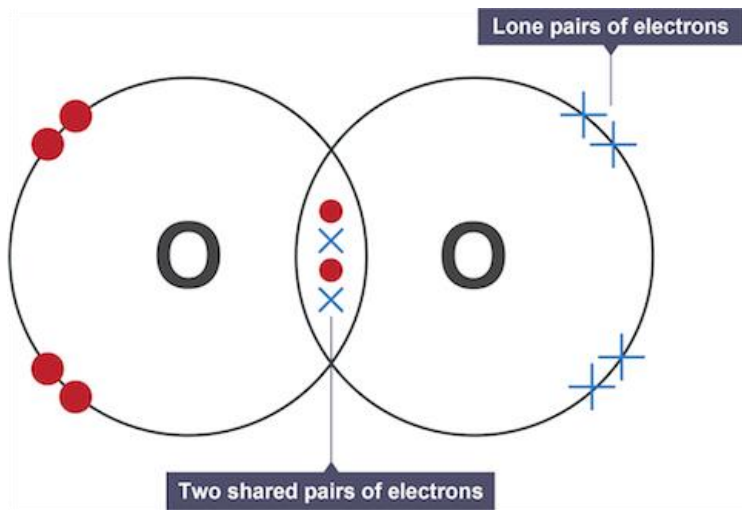
Single, Double, and Triple Bonds

H-H	O=O	N\equivN
		
<ul style="list-style-type: none"> • Share 1 pair valence electrons • Long bond length • Weakest • Lowest reactivity 	<ul style="list-style-type: none"> • Share 2 pairs valence electrons • Medium bond length • Intermediate strength • Medium reactivity 	<ul style="list-style-type: none"> • Share 3 pairs valence electrons • Short bond length • Strongest • Highest reactivity

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Illustration

Covalent bonds can be illustrated using dot and cross diagrams, in which shared and lone electrons are drawn.



Examples:

Name	Formula	Dot and cross diagram	Structural formula	Number of covalent bonds	Number of lone pairs
Hydrogen	H ₂		H — H	1	0
Chlorine	Cl ₂		Cl — Cl	1	6
Hydrogen Chloride	HCl		H — Cl	1	3
Water	H ₂ O			2	2
Ammonia	NH ₃			3	1
Methane	CH ₄			4	0

Carbon Allotropes

Allotropes are different forms of the same element in the same state:

Diamond, graphite, graphene and fullerenes (substances that include nanotubes and 'buckyballs', such as buckminsterfullerene) are four allotropes of pure carbon.

In all four allotropes, the carbon atoms are joined by strong covalent bonds, but in such different arrangements that the properties of the allotropes are very different.

Graphite

Graphite has a giant covalent structure in which:

- each carbon atom forms three covalent bonds with other carbon atoms
- the carbon atoms form layers of hexagonal rings
- there are weak forces of attraction between the layers
- there is one, non-bonded – or delocalized – electron for each atom

Graphite's properties include:

- high melting and boiling points. Graphite's many covalent bonds are strong and substantial energy is needed to break them.
- good electrical conductivity. Each carbon atom has an unbonded electron. The unbonded electrons are delocalised electrons that are free to move and carry charge.
- softness. The weak forces between graphite's layers allow them to slide.

Graphite is used as a lubricant and in pencils.

Graphene

Graphene is a single-atom thick layer of graphite with strong covalent bonds between each carbon atom. The atoms are arranged in hexagons. Its properties include:

- high melting and boiling points. Graphene's many covalent bonds are strong and substantial energy is needed to break them.
- good electrical conductivity. Each carbon atom has an unbonded electron. The unbonded electrons are delocalised electrons that are free to move and carry charge.
- very strong. Graphene's strong covalent bonds makes it 100 times stronger than steel. It is also the thinnest material possible – one atom thick – and very lightweight and transparent.

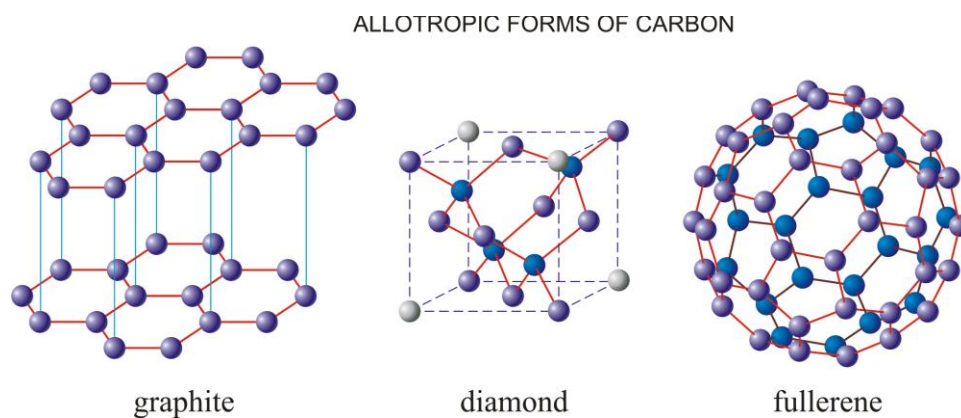
Diamond

Diamond is a giant covalent structure in which each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral, three-dimensional structure. Diamond's properties include:

- high melting and boiling points. Diamond's many covalent bonds are strong and substantial energy is needed to break them.
- does not conduct electricity. Diamond has no free ions or delocalised electrons to move and carry the charge.
- hardness. Diamond's three-dimensional tetrahedral structure with strong covalent bonds makes it very hard.

Bucky Balls (c60 fullerene)

Buckyballs are spheres or squashed spheres of carbon atoms. They are made up of large molecules so are not classed as giant covalent networks. Weak intermolecular forces exist between buckyballs. These need little energy to overcome, so substances consisting of buckyballs are slippery and have lower melting points than graphite or diamond



Simple and giant covalent structures

Simple Covalent Structure:

- In a simple covalent structure, a small number of atoms are bonded together by strong covalent bonds.
- Examples of simple covalent structures include molecules like water (H₂O), methane (CH₄), and carbon dioxide (CO₂).
- Properties of simple covalent structures:
 - Relatively low melting and boiling points: The intermolecular forces between simple covalent molecules are weaker compared to the strong intramolecular covalent bonds, resulting in lower melting and boiling points.
 - Softness: Simple covalent structures are often soft and malleable because the intermolecular forces are not as strong.

- Some simple covalent structures can exhibit electrical conductivity: Certain simple covalent structures, such as graphite, have delocalized electrons that can move freely, allowing them to conduct electricity.

Giant Covalent Structure:

- In a giant covalent structure, a large number of atoms are bonded together by strong covalent bonds, forming an extensive network.
- Examples of giant covalent structures include diamond, graphite, silicon dioxide (silica), and graphene.
- Properties of giant covalent structures:
 - High melting and boiling points: The extensive network of strong covalent bonds requires a significant amount of energy to break, resulting in high melting and boiling points.
 - Hardness and rigidity: The strong covalent bonds in giant covalent structures contribute to their hardness and rigidity.
 - Lack of electrical conductivity in most cases: Giant covalent structures generally do not conduct electricity due to the absence of free-moving electrons. However, there are exceptions like graphite, where delocalized electrons allow for electrical conductivity.
 - Varied allotropes: Different arrangements of atoms in giant covalent structures can give rise to different allotropes, such as diamond and graphite in the case of carbon.

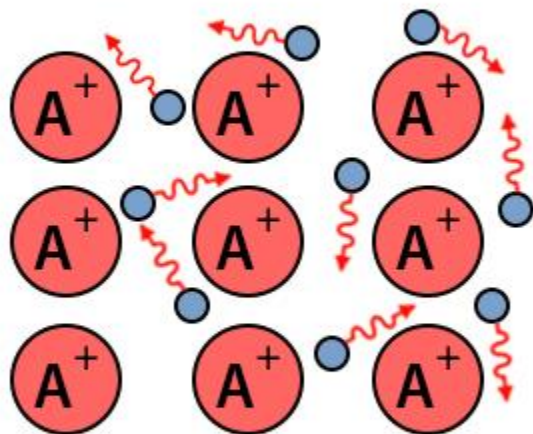
Metallic bonding

Metallic bonding is the electrostatic attraction between metal ions and a sea of delocalized electrons, where the electrons move freely throughout a lattice structure of metal atoms

Illustration

Metallic bonds can be represented through a 'sea of electrons' model:

Metallic Bond



Electrostatic attraction between
the electrons (●) and metal ions (A⁺)

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Links between properties and structure

- **Electrical conductivity:** Metallic bonding allows for the presence of a sea of delocalized electrons that can move freely throughout the metal lattice. This mobility of electrons enables metals to conduct electricity efficiently.
- **Thermal conductivity:** The presence of delocalized electrons in metallic bonding also contributes to high thermal conductivity in metals. The mobile electrons can transfer thermal energy rapidly through the metal lattice.
- **Malleability and ductility:** Metallic bonding's nature allows for the layers of metal atoms to slide past one another without breaking the bond. This property enables metals to be malleable (able to be hammered into different shapes) and ductile (able to be drawn into wires).
- **Luster:** The delocalized electrons in metallic bonding interact with light, resulting in the characteristic shine or luster observed in metals.
- **High melting and boiling points:** Metallic bonds are strong and require a significant amount of energy to break. This property gives metals high melting and boiling points, allowing them to exist as solids at room temperature (with exceptions like mercury).
- **Thermal expansion:** The presence of delocalized electrons affects how metals expand and contract with changes in temperature. The movement of these electrons contributes to the expansion of the metal lattice when heated.

Skills

Comparing different properties of different types of bonding

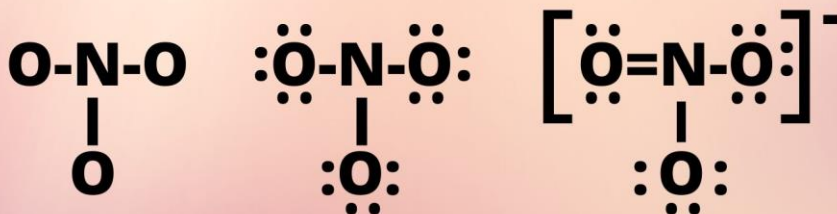
1. To compare the properties of substances resulting from different types of bonding, you can focus on specific characteristics and evaluate how they differ across the different types of bonding. Here's a step-by-step approach to comparing the properties:
2. Identify the type of bonding: Determine whether the substance exhibits ionic, covalent, or metallic bonding.
3. Select relevant properties: Choose properties that are commonly associated with the type of bonding under consideration. For example, for ionic bonding, you might consider properties such as melting point, solubility, or electrical conductivity.
4. Compare the properties: For each selected property, analyze how it differs across the different types of bonding. Consider factors such as the strength of the bonding, the nature of the interaction between particles, or the presence of mobile charge carriers.
5. Provide specific examples: Use specific substances or compounds as examples to illustrate the differences in properties resulting from different types of bonding. For instance, you could compare sodium chloride (ionic bonding), carbon dioxide (covalent bonding), and copper (metallic bonding) when discussing properties like melting point or electrical conductivity.

Lewis structure & dot and cross diagrams

Lewis

How to Draw a Lewis Structure

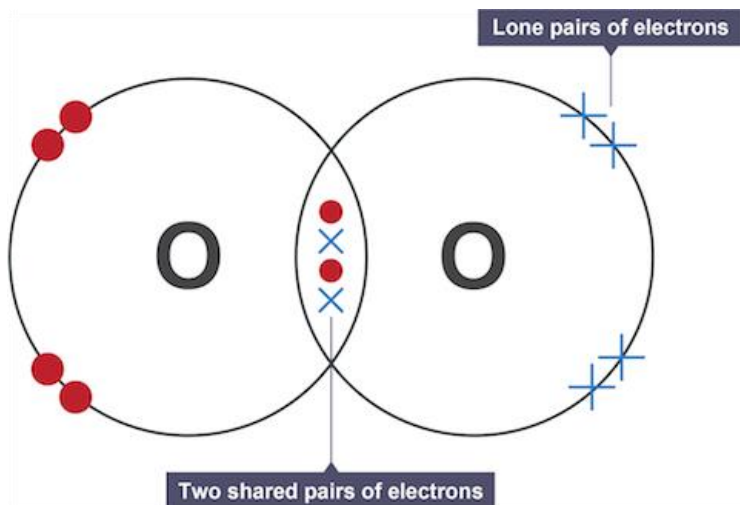
1. Find total valence electrons.
2. Draw skeleton structure using single bonds.
3. Assign remaining valence electrons.
4. Make multiple bonds if all octets aren't filled.



Add bracket and charge last.

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Dot and cross



Classification of bonding (intermolecular and intramolecular forces of attraction)

Intramolecular Forces of Attraction:

Covalent Bonding:

Classification: Covalent bonding occurs when atoms share electrons to form covalent bonds.

Identification: Look for the presence of shared electron pairs between atoms within a molecule.

Covalent bonds are typically formed between non-metal atoms or between non-metal and metalloid atoms.

Examples: H₂O (water), CH₄ (methane), CO₂ (carbon dioxide)

Ionic Bonding:

Classification: Ionic bonding occurs between a metal and a non-metal when electrons are transferred from the metal to the non-metal, resulting in the formation of cations and anions.

Identification: Look for the presence of charged ions within a compound. Ionic compounds are composed of positively charged cations and negatively charged anions.

Examples: NaCl (sodium chloride), MgO (magnesium oxide), CaCO₃ (calcium carbonate)

Metallic Bonding:

Classification: Metallic bonding occurs within metals, where delocalized electrons are shared between metal ions.

Identification: Look for the presence of a lattice structure made up of metal ions surrounded by a "sea" of delocalized electrons. Metallic bonding provides metals with their characteristic properties, such as high electrical conductivity and malleability.

Examples: Cu (copper), Fe (iron), Al (aluminum)

Intermolecular Forces of Attraction:

London Dispersion Forces (van der Waals Forces):

Classification: London dispersion forces are the weakest type of intermolecular force and occur between all molecules, regardless of polarity.

Identification: Look for temporary fluctuations in electron distribution, which result in temporary dipoles. These forces are present in all substances to some extent.

Examples: Non-polar molecules such as O₂ (oxygen), CH₄ (methane), and noble gases (e.g., Ne, Ar, Kr)

Dipole-Dipole Interactions:

Classification: Dipole-dipole interactions occur between polar molecules.

Identification: Look for a significant difference in electronegativity between atoms within a molecule, resulting in a permanent dipole moment. The positive end of one molecule is attracted to the negative end of another molecule.

Examples: HCl (hydrogen chloride), H₂O (water), NH₃ (ammonia)

Hydrogen Bonding:

Classification: Hydrogen bonding is a specific type of dipole-dipole interaction that occurs when hydrogen is bonded to highly electronegative atoms (N, O, F).

Identification: Look for hydrogen atoms bonded to nitrogen, oxygen, or fluorine. Hydrogen bonding is stronger than regular dipole-dipole interactions and leads to unique properties.

Examples: H₂O (water), NH₃ (ammonia), HF (hydrogen fluoride)

To determine the type of force present, analyze the nature of the atoms involved, their electronegativities, and the arrangement of electrons within the molecules or compounds. Understanding these classifications and their characteristics will help you differentiate between intramolecular and intermolecular forces of attraction in various substances.