Chemical Bonding

***What is a chemical bond?***

bond = something that binds, attaches or restrains

chemical bond = the force that binds atoms to each other

***Why do atoms bond?***

to achieve stability (= noble gas configuration: …s2p6)

* **“Octet Rule”**

“Representative (Main-block) elements form bonds by rearranging electrons so that each atom has a stable octet in its outermost energy level.”

(*Exception:* Duet Rule – H will react to have a full 1s energy level with 2 electrons)

Valence electrons = electrons involved in bonding, those in the outermost energy level of an atom (main-block; others may involve d energy as well)

* Electron Dot Diagrams (Lewis Structure)

Using X to mean any element, draw Lewis Structures for elements in these groups:

Group 1 Group 2 Group 13 Group 14 Group 15 Group 16 Group 17 Group 18

 X X X X X X X X

Note: The element’s symbol represents the element’s core electrons, while dots represent the valence electrons:

 e.g. Na 

 1s22s22p6 3s1

*Application of Octet Rule to Ionic Compounds*:

Metals **lose** electrons to become **cations**.

e.g. Na 🡪 Na+ + 1e-

Nonmetals **gain** electrons to become **anions**.

e.g. Cl + 1 e- 🡪 Cl-

Whiteboard work:

formation of NaCl – electron dot diagram and orbital notation



[ ]

then magnesium oxide, calcium fluoride, aluminum oxide, ammonium chloride

Types of Chemical Bonds

ionic, metallic, covalent, covalent network

***Type of bonding 🡪 properties of the chemical compound***

**Ionic compounds (salts)**

Salts are made of charged particles formed by donating (and receiving) electrons.

Note: Salts are neutral compounds, (+) = (-)

* held together by electrostatic forces of attraction between (+) and (-) charges in a crystal lattice
* crystalline at room temperature (~ 24oC): repeating 3-D structure in crystal lattice = unit cell (e.g. NaCl = cubic)
* strength of the ionic bond is amplified throughout the crystal structure
* hard
* brittle
* high m.p., high b.p. (due to compounded attraction of opposite charges in crystal lattice)
* soluble in water
* conduct electricity when molten or when dissolved in water (breaks down the crystal lattice, so charges can move)
* do **not** conduct electricity when solid

**Energy considerations:**

* Forming a crystal lattice 🡪 loss of a lot of PE 🡪  stability
* “lattice energy” = energy released when a crystal containing 1 mole of an ionic compound is formed from gaseous ions.

**Bonding terms**

*Chemical formula:*

shows relative numbers of atoms of each kind in a chemical compound by using atomic symbols and numerical subscripts. e.g. C2H4

*Empirical formula:*

smallest whole number ratio of atoms. e.g. CH2

(subset of chemical formula)

*Formula unit:*

smallest whole number ratio for ionic compounds

(= empirical formula, except for Hg22+). e.g. NaCl

*Molecular formula:*

chemical formula of one molecule. e.g. C2H4

*Structural formula:*  H H

hows how atoms are bonded together in a molecule. e.g. C=C

 H H

**Naming Monatomic Ions**

nomenclature = “naming”

from Latin *nomen* = “name”, *calare* = “to call”

monatomic ion vs. polyatomic ion

 1 atom + or - many atoms + or –

Use **atomic symbol** plus the **charge** (as **superscript**).

1. monatomic cation:

 Na**+**  sodium ion element’s name + “ion”

2. monatomic **anion**:

 Anions provide the **second** part of the name of an ionic compound.

 Cl- chlor**ide** **ion** use suffix “**ide**” + “**ion**”

#  S2-

 N**3-**

**How do you figure out the charge on an ion?**

1. Apply the **Octet Rule** (Groups 1, 2, 16, 17, many elements in Groups 13, 14, 15).
2. Memorize ion list….

# Naming cations with > 1 possible charge

Use Roman numerals:

e.g. copper**(I) ion** = Cu**+**

 copper**(II) ion** = Cu**2+**

(***Note***: Arabic numbers are used to denote **# of atoms** in a compound; are used as **subscripts**.)

**Naming Polyatomic Anions**

Polyatomic anions are usually oxyanions, and contain oxygen atoms covalently bonded to a central nonmetal or metal ion. In order to fulfill the Octet Rule, these compounds gain one or more electrons, giving them an overall charge.

“ite” NO2- nitr**ite** S**O3**2- sulf**ite**

more oxygen 🡪 “**ate**” N**O3**- nitrate S**O4**2- sulf**ate**

 Cl**O**- **hypo**chlor**ite** ***least*** oxygen

 Cl**O2**- chlor**ite**

 Cl**O3**- chlor**ate**

 Cl**O4**- ***per***chlor***ate*** ***most*** oxygen

 Mn**O4**- ***per***mangan***ate***

If the anion contains **hydrogen**, prefix name with “hydrogen”

**H**C**O3**- **hydrogen** carbon**ate**

**H**S**O4**- **hydrogen** sulf**ate**

**H**S- **hydrogen** sulf**ide**

**Binary Ionic Compounds**

cation + anion

Na+ Cl-

NaCl 🡪 sodium chloride

1. Write ions Pb4+ O2-

1. Balance charges by finding least common multiple of the ions’

charges. 4+ 2 x 2-

1. Write formula using the lowest whole number ratio of the compound’s atoms.

= “empirical formula” (ionic and molecular compounds)

= “formula unit” (ionic compounds only)

 **PbO2 lead (IV) oxide**

***Practice***:

calcium chloride iron (II) oxide iron (III) sulfide

aluminum nitride strontium phosphide rubidium iodide

Ionic Compounds Containing Polyatomic Ions

many atoms + or –

Note: There are only a couple of polyatomic cations, but many polyatomic anions.

Use **atomic symbols** plus the **charge** (as **superscript**).

1. polyatomic cation:

 NH4**+**  ammonium ion

2. polyatomic **anion**:

 OH- hydroxide ion

CO32- carbonate ion

Note: Charge refers to ***whole group*** of atoms, not just the last one.

**How do you write polyatomic ion formulae?**

1. Write ions NH4+ CO32-

1. Balance charges by finding least common multiple of the ions’

charges. 2 x 1+ 2-

1. Write formula using the lowest whole number ratio of the compound’s ions. Show > 1 polyatomic ion with parentheses. (**NH4)2**CO3 **ammonium carbonate**

***Practice***:

aluminum sulfate magnesium hydroxide copper (II) acetate

(See flowchart in textbook, p. 227)

**Metallic Bonds (metals)**

* held together by the attraction of free-floating valence electrons (-) for the (+)-charged metal ions in a lattice structure



1. What is a regular, repeating three-dimensional arrangement of atoms called?

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. Do the separate electrons that are shown belong exclusively to a single atom?

 What word is used to describe such electrons? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Are the electrons shown the only ones actually present? Explain.

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. Why are the central atoms shown as positively charged?

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

5. How does the number of separate electrons shown for the group 1A metal atoms compare to the number

 of atoms? Explain why in terms of valence electrons.

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. How does the number of separate electrons shown for the group 2A metal atoms compare to the number of atoms? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. What holds the metal atoms together in such an arrangement?

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. What term is used to describe this model of metallic bonding? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
2. How well do metals tend to conduct electricity? How does the model of metallic bonding account for that property? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Do metals tend to be brittle, or are they malleable and ductile? How does the model of metallic bonding account for that property? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# Characteristics of metals

* high m.p. and b.p.
* good conductors of electricity and heat in the solid state
* malleable, ductile
* shiny, reflective

**Alloys**

* mixture of elements that has metallic properties (solid solutions)
	+ substitutional alloys – atoms of similar sizes,

e.g. brass (Cu + Zn), bronze (Cu + Sn + Zn),

pewter (Sn + Sb + Pb)

* + interstitial – much smaller atoms fill spaces between larger atoms, e.g. carbon steel

**Covalently-bonded compounds (molecules)**

* held together by shared electrons (= covalent bond) e.g. H2



Both electrons spend some time around each nucleus, but majority of time in the middle.

greater (-) charge in middle, so nuclei stay attracted.

\*Memorize: NO F Cl BrI H

all exist diatomically in nature (more stable): N2 O2 F2 Cl2 Br2 I2 H2

7- rule: Go to Element #7, travel across to Group 7A, then down in the shape of a 7.

Hairogens: H2 (H), N2 and O2 (air), F2, Cl2, Br2, I2 (halogens)

**Characteristics of covalently-bonded compounds (molecules)**

* relatively low m.p., b.p.
* do not conduct electricity under any circumstances
* generally not soluble in water, but soluble in alcohol

Covalent Network Solids

* form covalent bonds in all directions 🡪 continuous network of strong covalent bonds 🡪 no individual molecules
* extremely hard, very high m.p. and b.p.
* nonvolatile, insoluble in all solvents
* brittle, nonconductors of heat and electricity
* diamond (C), quartz (SiO2)

***Compare structures of networks/lattices:***



1. covalent network (quartz, SiO2)
2. salt (NaCl)
3. metal (e.g. Cu)

**Formation of a Covalent Bond**

• A covalent bond forms when the orbitals on two atoms **overlap**.

 There are two electrons (usually one from each atom) of opposite spin in the orbital overlap.

• As two nuclei approach each other their atomic orbitals overlap.

 • As the amount of overlap ****, the energy of the interaction ****.

 • At some distance the ***minimum*** energy is reached = **bonding distance** (or **bond length**).

 • As the two atoms get closer, their nuclei begin to repel and the energy ****.

 • At the bonding distance, the attractive forces between nuclei and electrons just balance the repulsive forces (nucleus-nucleus, electron-electron).



Energy considerations

***What does “stable” mean?*** Changes that lower potential energy are favored.

**In covalent bonds**:

Shared electrons 🡪 loss of PE 🡪 stability

Bond energy = energy required to break a chemical bond and form neutral atoms

# Relationship between bond length and bond energy in molecules

Bond length = average distance between two bonded atoms (distance of minimum potential energy)

As Ebond Lengthbond because the closer the atoms are, the more attraction between nuclei and electron clouds.  harder to separate.

Naming Molecular Compounds

See p. 248 in textbook for list of prefixes to memorize

**A.** **Prefixes, roots, suffixes**

1. Begin with element with lowest electronegativity. **nitrogen**
2. Add appropriate prefix (unless it is mono-). **dinitrogen**
3. End second element with "**ide**" (as for ionic compounds…). ox**ide**
4. Use appropriate prefix for second element. **tetr**ox**ide**
* **dinitrogen tetroxide**
* **N2O4**

**Practice:**

**CCl4**

**CO**

**CO2**

**As2S3**

**P2O5**

**P4O10**

**Compounds That Become Acids When Dissolved in Water**

General Formula: HX H+ X-

monatomic

or

polyatomic

anion

**Three Rules:**

1. When X ends in “**ide**” (e.g. chlor**ide**, cyan**ide**)

 🡪 “**hydro**\_\_\_\_\_\_\_ **ic** acid”

 e.g. **hydro**chlor**ic** acid, **hydro**cyan**ic** acid

2. When X ends in “**ite**” (e.g. chlor**ite**, sulf**ite**)

 🡪 “\_\_\_\_\_\_**ous** acid”

 e.g. chlor**ous** acid

 sulfur**ous** acid

3. When X ends in “**ate**” (e.g. chlor**ate**, sulf**ate**)

 🡪 “\_\_\_\_\_\_ **ic** acid”

 e.g. chlor**ic** acid

 sulfur**ic** acid

**Your turn:** HBr

 HNO2

 HNO3

**Writing Lewis Structures (Electron Dot Diagrams) for**

**Molecules and Polyatomic Ions**

1. Count up the total # valence electrons for all atoms in molecule or ion.

*Molecule*: sum the # electrons in all atoms

*Ion*: sum the # of valence electrons + charge on ion

1. Write down the symbols of atoms to show how joined.
2. If carbon, usually in center

OR

1. most electropositive atom in center

OR

1. nonmetal in center (not H, O)
2. Trial drawing with dots (2 for H, 8 for others)

If pair left over, make = or 

* Maybe an exception: more or less than 8 in valence

**Practice:**

CH3Br ClO- C2H6

BrI SO42- C2H4

H2S H2PO4- C2H2

PH3 NH4+ C6H6

**Molecular Geometry**

**Basic Molecular Shapes**



We count only the atoms when describing the shape of a molecule:

 



(AP Chem)



(Honors &

AP Chem)

(Honors &

AP Chem)

(AP Chem)

(AP Chem)

(AP Chem)

 **What determines the type of chemical bond?**

Differences in **electronegativities** between bonding atoms

The greater the difference 🡪 more ionic (on a continuum, where a difference of 1.7 is approximately 50:50 ionic:covalent)

 **polar** **nonpolar covalent**

4.0 **ionic** 2.0-1.7 **covalent**  0.4 0.0

e.g. electronegativity of **Cl** - electronegativity of **Na**

 3.0 - 0.9 = 2.1 ionic

 electronegativity of **Cl** - electronegativity of **Al**

 3.0 - 1.5 = 1.5 polar covalent

 electronegativity of **Cl** - electronegativity of **Br**

 3.0 - 2.8 = 0.2 nonpolar covalent

**ionic** bonds - the least electronegative atom donates 1 or more electrons to the more electronegative atom

**covalent** bonds - electrons from both atoms are shared

 **nonpolar** - electrons are shared equally by both atoms

 **polar** - electrons are not shared equally - the more electronegative atom attracts the shared

electrons more strongly

Predicting Polarity in Molecules

1. Draw electron dot diagram (Lewis structure) of the molecule.
2. Compare the electronegativities of *each* of the bonded atom pairs.
3. Note whether *each* bond is polar (P) or nonpolar (NP).
4. Draw an arrow towards the most electronegative atom for each bond.
5. Is molecule symmetrical? (if so, molecule is NP)

Examples:

CH4 CCl4 CH3Cl

NH3 CO2 H2O





**Effect of bond polarity on molecule polarity**



Intermolecular forces (van der Waals forces)

1. **Dispersion or “London” forces** (nonpolar molecules)

(named after Fritz London, 1900-1954)

1. **weakest** intermolecular force
2. results from the constant motion of electrons 🡪 uneven distribution of electrons at any particular moment:

**“temporary dipole”** which may 🡪 dipole in nearby molecule.

1. acts on **all** molecules all the time
2. **only** intermolecular force acting among **noble gas** atoms and **nonpolar** molecules
3. ****with **number of electrons**: note m.p., b.p.

e.g. halogens

F2, Cl2 gases at room T

Br2 liquid at room T (larger than F2 and Cl2)

I2 solid at room T (largest)

1. **Dipole force** (polar molecules)

a. the attraction between two **polar** molecules:

**(-)** end of one polar molecule attracts the **(+)** end

of another polar molecule

1. **more** polar 🡪 **stronger** dipole force
2. **closer together** 🡪 **stronger** dipole force

3. **Hydrogen bonding**

1. always involves **H**

usually involves **O, F** or **N (small, high electronegativity)**

1. strongest intermolecular force

# How strong? 5% of the strength of a covalent bond

1. 🡪 **higher b.p.** and **higher viscosity**
2. e.g. H2O

**Comparing Bond Types**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Properties | **Ionic: salts** | **Metallic: metals** | **Covalent: molecules** | **Covalent network** |
| Description of bond |  |  |  |  |
| General appearance |  |  |  |  |
| Malleable vs brittle |  |  |  |  |
| Conduct electricity? |  |  |  |  |
| under what circumstances? |  |  |  |  |
| Soluble in water or organic solvents? |  |  |  |  |
| m.p., b.p. |  |  |  |  |
| State of matter at room temp? |  |  |  |  |

***Comparing ALL bond types:***

Which is **stronger**?

**covalent network > metallic** > **ionic** > **covalent** (molecules) > **H bond** > **dipole** > **dispersion**

e.g. Compare melting points:

**SiO2 > Fe** > **NaCl** > **C12H22O11** > **H2O** > **HCl** > **H2**

**sand > iron** > **salt** > **sugar** > **ice** > **hydrogen chloride > hydrogen gas**