TOPIC 5.1

Chemical Bonds: An Overview

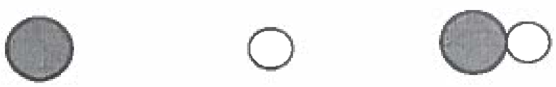
What are chemical bonds and why do they form?

Chemical change involves the movement of electrons to form new substances, and in that process bonds are broken and formed.

Cl_2 → Cl + Cl

Breaking bonds: absorbs/requires energy

Na + Cl → NaCl



Making bonds: releases energy

Why bother forming bonds at all? Recall from Unit 4 that reactivity depends on: stability
Depending on the type of elements involved, there are different types of bonds that will form.

Bond type is determined by differences in electronegativity

What's that again? ability of an atom to attract electrons in a chemical bond

ΔEN < 1.7 ΔEN > 1.7

COVALENT BONDS

- Electrons are shared between atoms because of a small difference in electronegativity (ΔEN)
- Form between…
  2 nonmetals

IONIC BONDS

- Electrons are transferred between atoms because of a large difference in electronegativity (ΔEN)
- Form between…
  metal & nonmetal

TOPIC 5.2

Chemical Formulas: Ionic Compounds

How do we know the ratios that a positive ion and negative ion will combine?

Metals & Nonmetals

The sum of **charges** in an ionic compound must = ZERO! ☆

How do we know the charge of each element in an ionic compound?

METALS lose electrons to form positively charged ions

NONMETALS gain electrons to form negatively charged ions

The charge that an element has can also be determine by checking out its: Oxidation #

9.01218 +2
Be
4
2-2

Practice

Sodium bonding with chlorine: NaCl

Na^{+1} Cl^{-1} 1:1 ratio

Magnesium bonding with chlorine: $MgCl_2$

Mg^{2+} Cl^{-1} Cl^{-1} 1:2 ratio

Polyatomic Ions – many atoms covalently bonded together; the whole structure acts as a single charged unit (ION)

☆Table E

Ammonium bonding with chlorine: NH_4Cl

NH_4^{+} Cl^{-}

Calcium bonding with hydroxide: $Ca(OH)_2$

Ca^{2+} OH^{-} OH^{-}

What type of bonding is present in these compounds?

BOTH covalent and ionic!

Transition Metals (Groups 3 – 12)

Iron bonding with chlorine: $FeCl_2$ or $FeCl_3$

Fe^{2+} Cl^{-} Cl^{-}

OR

Fe^{3+} Cl^{-} Cl^{-} Cl^{-}

55.845 +2 +3
Fe
26
2-8-14-2

Copper bonding with oxygen: Cu_2O or CuO

Cu^{+1} Cu^{+1} O^{2-}

OR

Cu^{2+} O^{2-}

TOPIC 5.3

Naming Ionic Compounds

How do we name ionic compounds?

I. Rules for Naming Ionic Compounds

Chemists have a universal way of naming compounds based on what's called the IUPAC system. That way scientists from all across the world can have a common language when writing/talking about compounds.

Here's how to talk about ionic compounds in the language of chemistry:

1) First, name the positively charged ion
   - Metal – element name (unchanged)
   - Polyatomic Ion – Table E
2) Next, write the name of the negatively charged ion
   - Nonmetal – Table S, but change the suffix to "-ide"
   - Polyatomic Ion – Table E

Most Common Nonmetals	
Hydrogen =	hydride
Nitrogen =	nitride
Carbon =	carbide
Oxygen =	oxide
Fluorine =	fluoride
Sulfur =	sulfide
Phosphorous =	phosphide

Practice – Fill in the following table by naming the compound or writing the formula based on the rules outlined above

Formula	Name
Na_3N	sodium nitride
Al_2S_3	aluminum sulfide
$Ca(OH)_2$	calcium hydroxide
LiF	Lithium fluoride Li^+ F^-
$MgCl_2$	Magnesium chloride Mg^{2+} Cl^- Cl^-
Ca_3P_2	Calcium phosphide Ca^{2+} P^{3-} P^{3-} or Ca^{2+} P^{3-} Ca^{2+} Ca^{2+}

II. Stock System – Transition Metals

Remember when we tried to write the chemical formula for when iron bonds with chlorine, but there were multiple options? How do we name it so we know exactly which option we're talking about?

Since transition metals have multiple charges, we have to use Roman numerals to refer to which charge the atom has in that particular compound.

Practice Problems – Name the following ionic compounds

$FeCl_2$ (+2, -1) iron (II) chloride

CuO (+2, -2) copper (II) oxide

$FeCl_3$ (+3, -1) iron (III) chloride

Cu_2O (+1, +1, -2) Copper (I) oxide

TOPIC 5.4

Lewis Dot Structures: Ionic Compounds

How can we model ionic compounds using Lewis Dot Structures?

Steps for Drawing Lewis Dot Diagrams of Ionic Compounds

1. Draw Lewis Dot Diagrams for each individual element in the compound

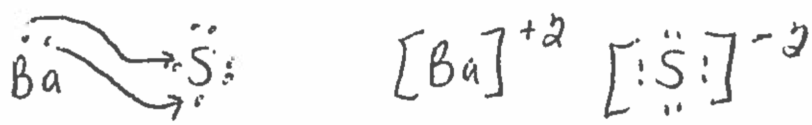
   Example: NaCl

2. Draw arrows to indicate the transfer of electrons
3. Redraw the bonded compound with appropriate dots, brackets, and charges

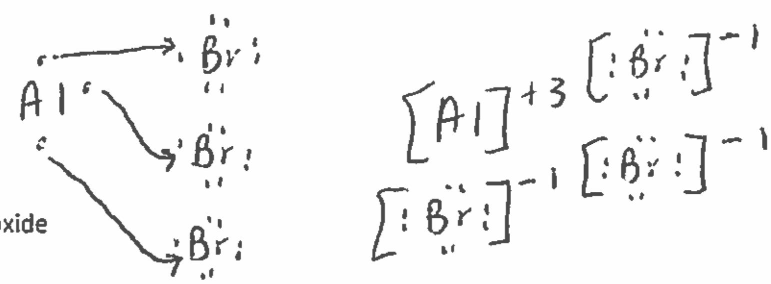
Practice Problems

KF

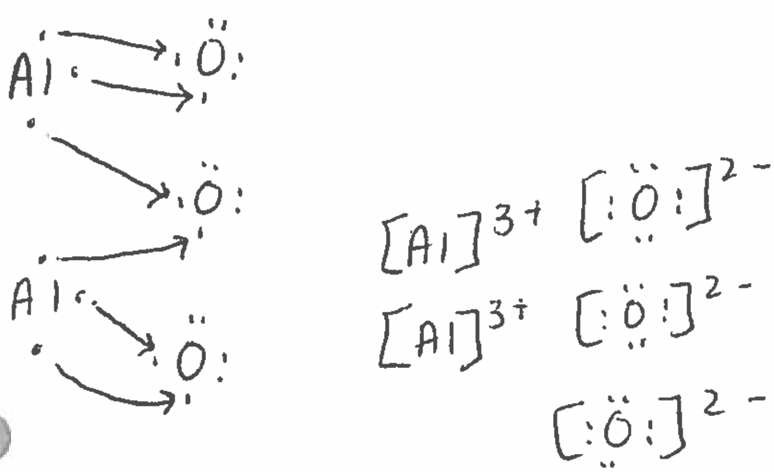
BaS



$AlBr_3$



Aluminum oxide



TOPIC 5.5

Covalent/Molecular Bonding - Diatomic Elements

Why do some elements share electrons?

I. Overview

CHEMICAL BONDS

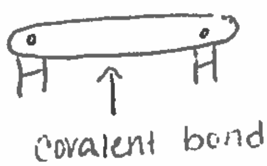
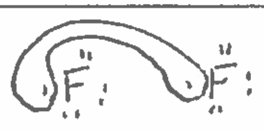
- Electrons are transferred → IONIC
- Electrons are shared → COVALENT / MOLECULAR

II. Diatomic Elements

Recall that there are 7 elements that exist naturally as pairs, or as diatomic elements. YOU MUST MEMORIZE THESE.

HOFBrINCl or "7 - up"

Why do these elements pair up in nature? STABILITY.

Diatomic Element	Lewis Dot Structure of Individual Atoms	Lewis Dot Structure of Covalent Molecule	# of bonds present	# of electrons shared
H_2		H — H 1 dashed line → single bond	1 single bond	2
O_2		O = O 2 dashed lines double bond	2 double bond	4
F_2 (Cl_2 Br_2 I_2)		:F — F:	1 single bond	2
N_2		:N ≡ N: 3 dashed lines → triple bond	3 triple bond	6

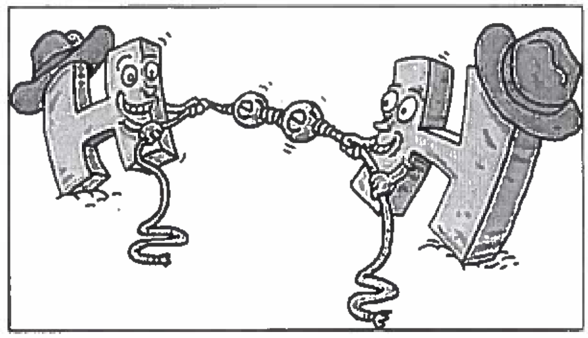
TOPIC 5.6

Polarity of Chemical Bonds

How do differences in electronegativity affect electron sharing?

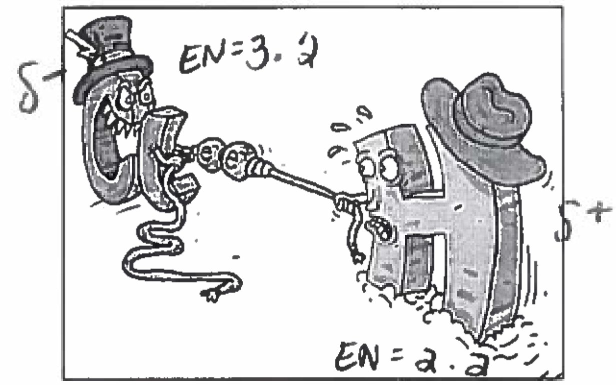
I. **Nonpolar Covalent Bonds**

- Electrons are <u>equally</u> shared between atoms
- $\Delta EN = 0$
- Examples:
  - ➢ Diatomic molecules
  - ➢ CS (2.6, 2.6)



II. **Polar Covalent Bonds**

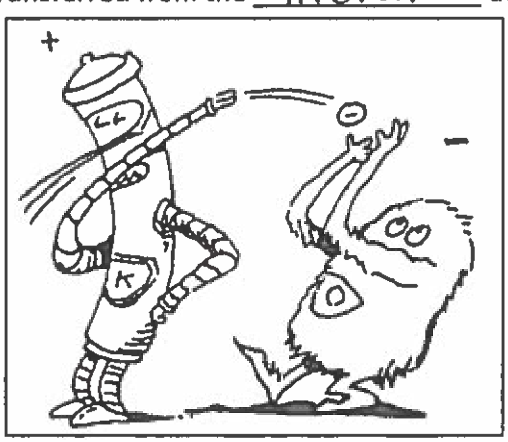
- Electrons are <u>unequally</u> shared between atoms
  - ➢ The atom with the greater EN <u>pulls</u> the electrons closer and becomes partially <u>NEGATIVE (δ^-)</u>
  - ➢ The atom with the lower EN becomes partially <u>POSITIVE (δ^+)</u>



- $\Delta EN > 0$
- Examples:
  - ➢ N and H
  - ➢ H and F

III. **Ionic Bonds**

- Ionic bonds are considered **completely polarized**
- $\Delta EN > 2.0$
- The <u>nonmetal</u> "pulls" so much harder on the electrons that they are completely transferred from the <u>metal</u> atom



TOPIC 5.7

Lewis Dot Structures: Covalent/Molecular Compounds

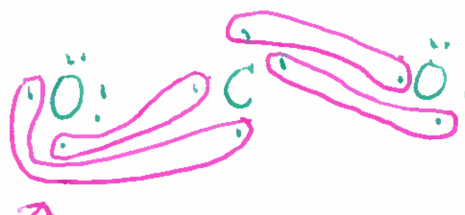
How can we model covalent compounds using Lewis dot structures?

You've already seen how to draw Lewis dot structures of covalent compounds when we looked at the 7 diatomic elements. Let's revisit the concepts/vocabulary involved:

Bond Type	Electron *Pairs* Shared	*Total* Electrons Shared	Example Diagram
SINGLE BOND	1	2	:F̤̈—F̤̈:
Multiple Covalent Bonds			
DOUBLE BOND	2	4	Ö=Ö
TRIPLE BOND	3	6	:N≡N:

So we've seen examples of each type of bond in the diatomic molecules with just 2 atoms; but what about molecules that have more than 2 atoms, like H_2O or CCl_4 or $C_{11}H_{22}O_{11}$? (just kidding about that last one, you won't have to draw a molecule that big)

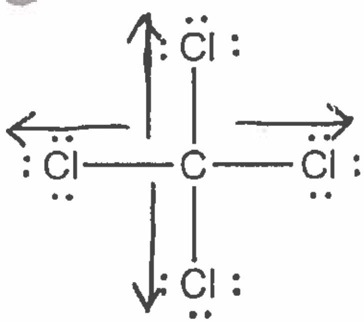
PLAN OF ATTACK (aka rules for drawing Lewis Dot Diagrams of Covalent Compounds)

Steps	Example – CO_2
1) Make sure you are dealing with a covalent compound and NOT an ionic compound	2 nonmetals ✓
2) Draw Lewis Dot Stuctures of each individual atom, arranged so that the element that can form the most attachments is in the center Carbon: 4 Nitrogen: 3 Oxygen: 2 Halogens and hydrogen: 1	
3) Circle bonding electron pairs so that all atoms have satisfied the octet rule.	
4) "Clean it up" – replace circled electrons with single lines to represent bonds	O = C = O

TOPIC 5.8

Polarity of Molecules

How does symmetry affect the polarity of molecules?



Each individual bond in the CCl_4 molecule is polar.

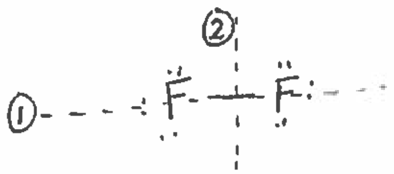
But is the <u>entire</u> molecule polar? NO

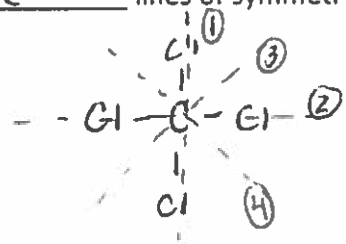
think about it like a game of tug-of-war and each team is pulling equally

HOW TO DETERMINE WHETHER A MOLECULE IS POLAR OR NONPOLAR

If the molecule is...

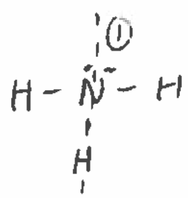
Symmetrical: Two or more lines of symmetry through molecule/compound

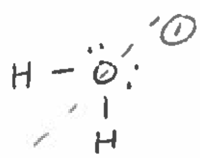




⟹ NONPOLAR

Asymmetrical: ONE or no lines of symmetry through molecule/compound

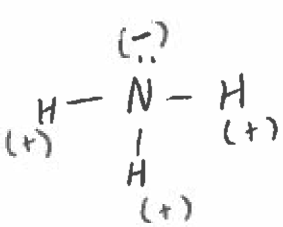




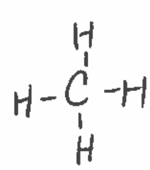
⟹ POLAR

Practice Problems – Determine if the molecule is polar or nonpolar. If polar, identify which end is positive and which is negative.

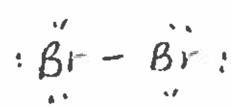
NH_3 – polar



CH_4 – nonpolar



Br_2 – nonpolar



For more practice, go back to Assign. #8 and see if you can determine which molecules are polar.

TOPIC 5.9 → **Intermolecular Forces (IMFs)**

What are the different types of forces that hold molecules together?

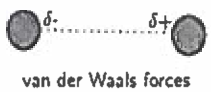
I. **Intermolecular Forces/"** Particle Attractions **"**

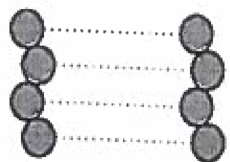
(Inter → between, molecular → molecules)

★ forces between molecules (NOT bonds)

II. **Types of IMFs**

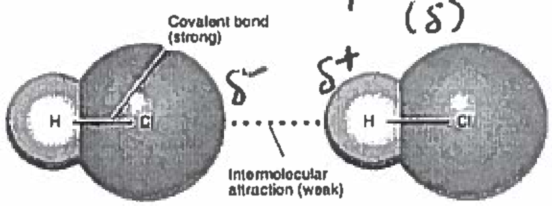
van der Waals forces — short-lived attractions between NONPOLAR molecules





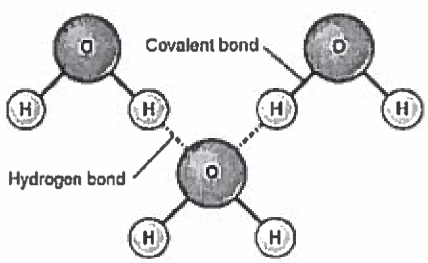
★ strongest when molecules are larger

Dipole-dipole attraction — partial (δ) +/− attraction between POLAR molecules



★ strongest when ΔEN is greater

Hydrogen bonding — attraction between δ+ (partial positive) hydrogen in one polar molecule to a F, O, or N of another molecule.



Ex: ★H_2O★, NH_3, HF

III. **Importance of IMFs**

- IMFS are determined by the Structure of a substance
- the strength of IMFs determines the properties of a substance
- Example:

The stronger the intermolecular force, the higher the melting point/boiling point

Recall Table H: Which liquid has the **strongest** intermolecular forces? ethanoic acid (highest MP)

★ **BIG IDEA: STRUCTURE DETERMINES PROPERTIES!!!** ★

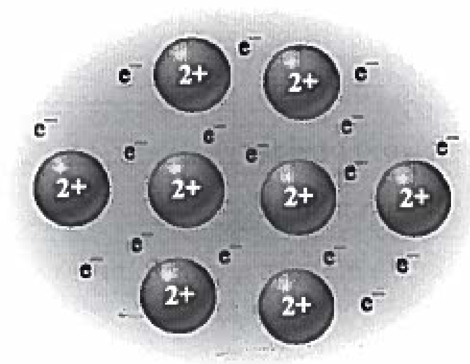
TOPIC 5.10

Metallic Bonding

How do metal atoms stick together?

You may (or may not) be wondering about bonding between metal atoms. After all, we've discussed bonds that form between a metal and a nonmetal (ionic) and bonds between nonmetals (covalent). So how does, say, a chunk of aluminum stay together? The answer lies in another type of "bond" that's not really a bond...

Metallic bond: attraction between positive metal ions and "sea" of valence electrons



This means of staying attached gives metals unique properties among groups of atoms:

- malleable: able to be flattened into sheets
  - Contrast: brittle
- ductile: able to be pulled into wires
- good conductors: able to provide a mobile path for the flow of electricity
- luster: distinct metallic shine
- HIGH melting points: result of strong attractive forces between nuclei and "delocalized' electrons

TOPIC 5.11

Bonding and Physical Properties

Do metals and nonmetals retain their properties when they form compounds?

We've stuck together atoms in many different ways this unit. How does that bonding process affect the properties of the macro-sized samples of elements or compounds that we create? The key idea to tie everything together is that

Chemical structure determines **physical properties**.

What type of bonds hold a sample of atoms together is a huge component of the sample's chemical structure; therefore, bond type is a key indicator of the physical properties a sample will have.

Bond Type	Covalent Molecules		Ionic Compounds	Metals
	Nonpolar	***Polar***		
Type of IMF	van der Waals	dipole-dipole	ion - ion	⊕ nuclei and ⊖ electrons
Melting/boiling point	low	moderate	high	high
Phase of matter at STP	s, l, g	s, l, g	solid	solid
Conductivity as a solid*	poor	poor	poor	good
Conductivity as a liquid or aqueous solution*	poor	poor	good	good
Likelihood to dissolve in water	low	high	high	low

*Conductivity is a result of MOBILE ions (or electrons) that can provide electricity with a moving path through which to flow.

Please note that these are just GENERAL trends that certainly have exceptions! This table seeks to give you a big-picture idea of how changing bond type can change the properties of a substance, but note that changing molecule size, ion charge, type of metal, etc. can also have an effect on physical properties!