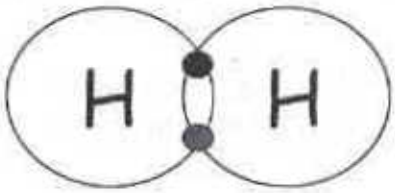
There are TWO types of covalent bonds:

A) Non-Polar Covalent Bond – two atoms interact to **equally share** electrons (i.e. H_2)

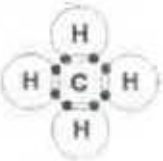


B) Polar Covalent Bond –two atoms **unequally share** electrons. Electrons are pulled more closely to the more **electronegative atom.** This means one "side" of the molecule will be more negative, and one side will be more positive...since the molecule has partially negative and positive sides, this allows it to interact with other molecules! More on this concept will be covered in our first unit!

Rapid Review! Try these questions:

1. If an ion were to gain an electron, is it positively charged, or negatively charged? Negatively charged
2. Using a periodic table, examine an atom of chlorine. How many electrons does chlorine have in its valence shell? What will typically happen to an atom of chlorine? 17 Electrons,
3. Explain what is meant by the notation: Cl^{1-} Chlorine with an extra electron (18 electrons)
4. A carbon atom has 4 electrons in its outer energy level. It thus needs 4 more electrons to become stable.
5. How many covalent bonds are there in methane (CH_4)? (see diagram on right)
6. A covalent bond can be represented by using a short line rather than dots. Redraw CH_4 (from question 7) using the short line notation (each line represents a pair of electrons)
7. Consider the following arrangement of carbon atoms: -C – C – C -
   (a) How many hydrogen atoms can be attached to carbon #1 (leftmost carbon)? _____
   (b) How many hydrogen atoms can be attached to carbon #2 (centre carbon)? _____
8. Consider the following arrangement of carbon atoms C = C. There are two bonds (a double bond) between the carbon atoms. Each bond (a single bond) represents one shared pair of electrons. How many electrons are there in a double bond?
9. Draw a triple bond between two carbon atoms. How many electrons are there in a triple bond?
10. A bond is called nonpolar covalent when electrons are equally shared (therefore the molecule is nonpolar and has no negative end or no positive end). A polar covalent bond occurs when electrons are shared unequally (that is, one atom has a stronger attraction for the electrons) which makes one side of the molecule slightly negative and the other side of the molecule slightly positive. There is a table of electronegativities that you can use to look up values of how strongly different atoms attract electrons.

    In the compound HCl, Cl is more electronegative than H. This means that Cl attracts electrons to it more than H does. Is this a nonpolar or polar compound? Which "side" of this molecule (i.e. which atom) would be more negative? Which side would be more positive?



SBI4U

Getting Ready for Grade 12 Biology: Review Package

Important Chemistry Concepts

There are a few important concepts from SCH3U0 (don't worry if you haven't taken it. We will get you caught up!) which we will be looking at in our biochemistry unit. Try to work through some of the questions in this package to give yourself an idea of where you stand on the basics of biochemistry.

Part 1 – Intramolecular Bonding

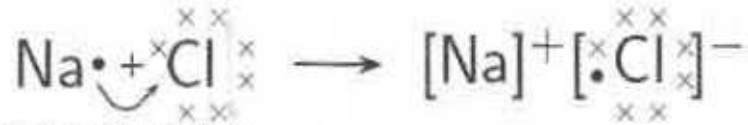
Intramolecular = the bonds within a molecule (i.e. the forces holding a molecule together)
There are two main types of intramolecular bonds that you have studied in the past: **ionic** and **molecular (covalent).** Recall that covalent compounds can be further classified as polar covalent or non-polar covalent (more on this below).

Why do atoms/elements/ions form bonds?

> Recall: Molecules form bonds to fulfill the **OCTET RULE**. Atoms bond to completely FILL or EMPTY their valence shell. When this happens, they achieve an electron configuration similar to Noble Gases (i.e. they achieve stability by having a FULL valence shell)

Recap of IONIC Bonding:

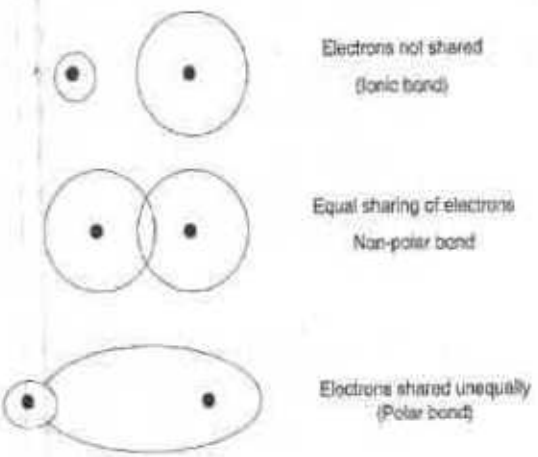
- Metal + non-metal
- Metal loses an electron, non-metal gains an electron
- Opposites attract!
- Properties of ionic compounds:
  - They form crystalline solids
  - They have high melting & boiling points
  - They are hard & brittle
  - They are soluble (break apart in water)
  - They conduct electricity when dissolved in water



electron transer from sodium to chlorine

Recap of Covalent (Molecular) Bonding:

- 2 or more non-metals
- Sharing of electrons (not transferring)
- Can form single bonds (1 pair of electrons), double bonds (2 pairs of electrons), or triple bonds (3 pairs of electrons)
- Electrons participating in the bond are called the bonding pair
- Electrons not participating in the bond are called the **lone pair**
- Properties of covalent/molecular compounds
  - Soft & squishy compared to ionic
  - They have lower melting & boiling points
  - They are insoluble
  - They don't conduct electricity when dissolved in water



Part 2 – Intermolecular Forces

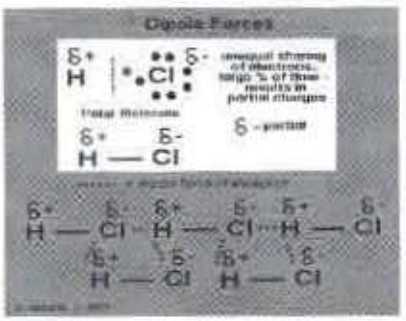
Intermolecular = the bonds/forces occurring BETWEEN molecules

Three Types:

1. Dipole-Dipole
2. Hydrogen Bonding (strongest)
3. London Disperson Force (LDF) aka Van Der Waals Forces (weakest)

1. Dipole-Dipole Forces
   - Holds polar molecules together
   - δ + (delta positive) or δ – (delta negative). The partial +ve charge on one molecule is electrostatically attracted to the partial –ve charge on a neighbour molecule

2. Hydrogen Bonds (a type of dipole-dipole)
   - Force between H and O, N or F
   - Strongest intermolecular force (H nucleus is small and positively charged and F, O, and N are very electronegative so there is a strong attraction); but easily formed and easily broken
   - The Force that holds water molecules together

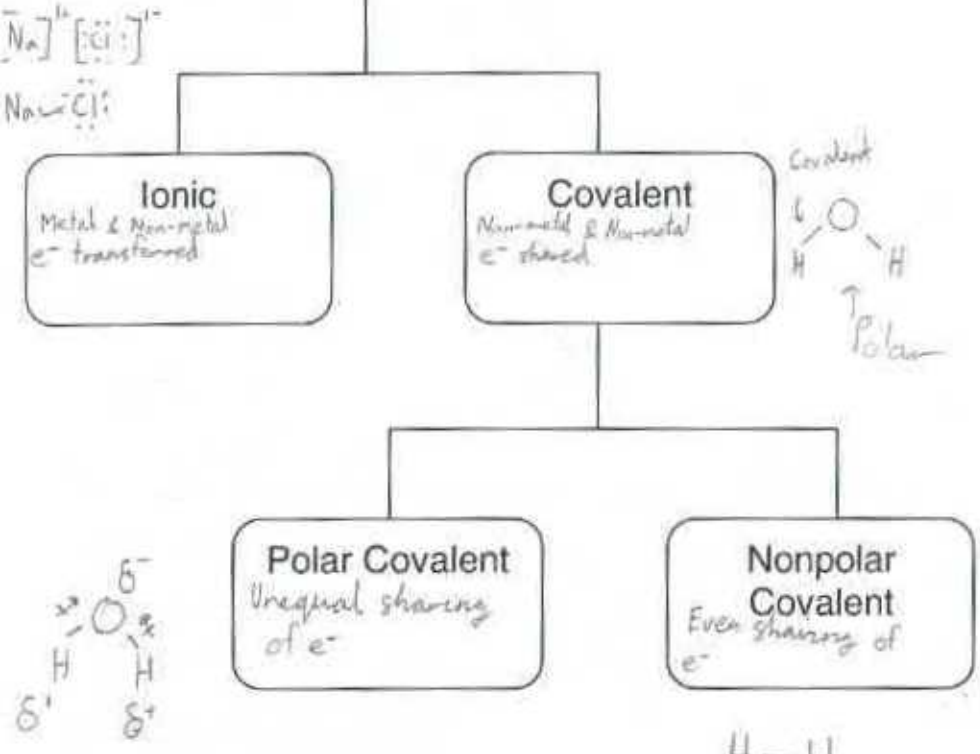


3. London Dispersion Forces (also called Van Der Waals Forces)
   - Weakest bonds; exist between all atoms and molecules
   - More significant in NON-POLAR molecules
   - The only type of intermolecular force operating between non-polar molecules (ex. between two H)
   - Equal shading shows no electrical distortion
   - Electrons are moving so at any given time, one side is slightly more negative than the other
   - An instant later...the molecule on the right is attracted to the molecule on the left because of the slight positive charge

Chemical Bonds Summary

All Chemical Interactions

- Intramolecular Bonds
  Bonds within a single molecule
  - Ionic
    Metal & Non-metal
    e⁻ transferred
  - Covalent
    Non-metal & Non-metal
    e⁻ shared
    - Polar Covalent
      Unequal sharing of e⁻
    - Nonpolar Covalent
      Even sharing of e⁻
- Intermolecular Forces aka **van der Waals Forces**
  - Dipole-Dipole
    - Hydrogen Bonds
  - London Dispersion Forces (LDF)
    Weak beta LDF



Not a macromolecule but essential to life

Biochemistry Fundamentals – Chemistry of Life

Types of Biological Molecules:

Water (inorganic)	Proteins (organic)	Lipids (organic)	Carbohydrates (organic)	Nucleic Acids (organic)

Organic Compounds: Has Carbon & Hydrogen + other atoms (Oxygen, Nitrogen, Phosphorus, sulfur, etc.)

Inorganic Compounds: Does not contain Hydrogen and/or Carbon

IMPORTANT – the function of a molecule is determined by its unique structure. The unique structure is determined by:

- Bonds between atoms
  - The bonds between organic molecules often act as energy reservoirs. Especially for C-H bonds
- The overall shape of the molecule.

Isomers: Same chemical formula but differing position/placement of atoms

CHEMICAL BONDS

Chemical bonds are linkages between atoms within a molecule. The bonds act like a chemical glue that hold the atoms together.

Bonding Capacity: The number of bonds an atom can form with other atoms. You can determine bonding capacity of an element by taking the maximum number of electrons an orbital shell can hold and subtracting the number of valence electrons.

E.g. Oxygen has 6 valence electrons in a shell that can hold 8

8 - 6 = 2

Therefore, oxygen's bonding capacity is 2

Determine the bonding capacity of the following elements:

Carbon	Sulphur	Nitrogen	Hydrogen	Phosphorus
8-4=4	8-6=2	8-5=3	2-1=1	8-5=3

Research drawing & understanding α & β glucose molecule