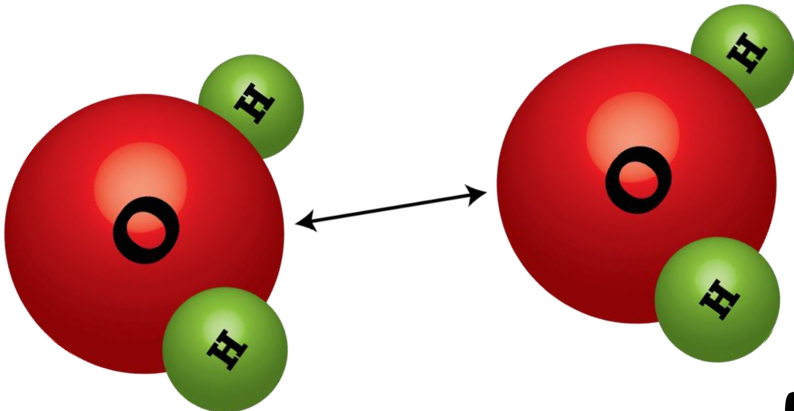
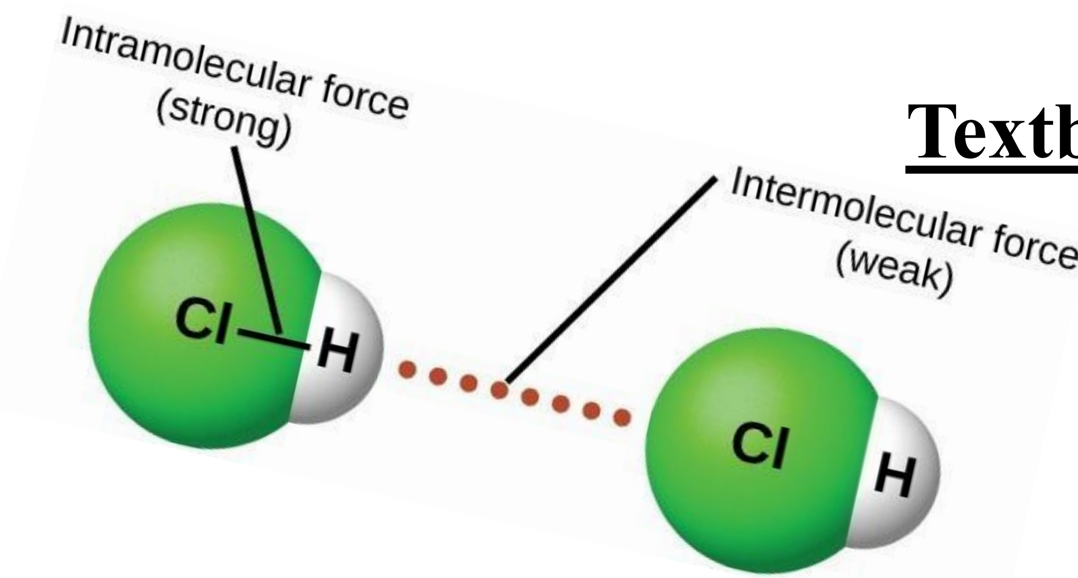


[4.1] Intermolecular and Intramolecular Forces



<u>Textbook</u>: p. 171 – 182

Chemical bonding

Chemical Bonding

- **Chemical bond** = "chemical glue", **attraction (electrostatic force)** between atoms, ions, molecules to **achieve more stability.**

When the **outer shell** is full it is called a **STABLE OCTET**.

Atoms want to have **full** outer shell. So they will bond with other atoms to get ***full outer shell***.

- There is **2 types** by location:
  - ***Intra***molecular – ***inside*** of the molecule
  - ***Inter***molecular – between 2 or more molecules (***outside***)

Types of **bonding** by electron arrangement

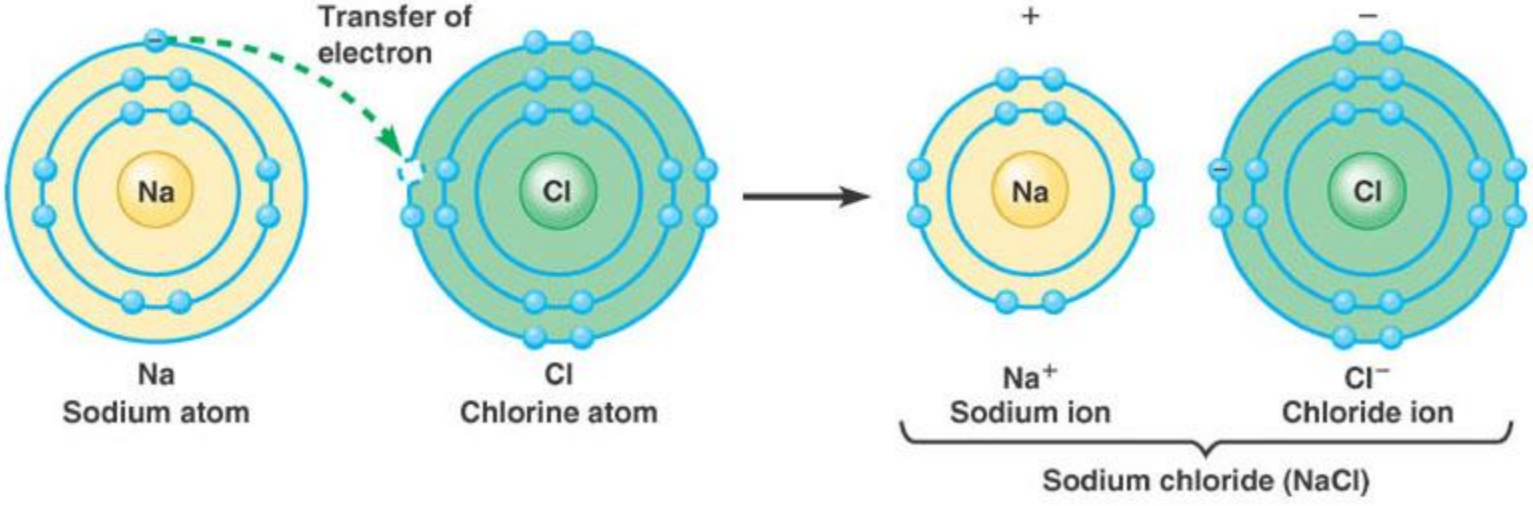
- **Ionic** bonding (total transfer of **e-** from one atom to another)

- **Covalent** bonding – sharing of **e-** between two atoms

- **Metallic** bonding –all **e-** are shared among all atoms

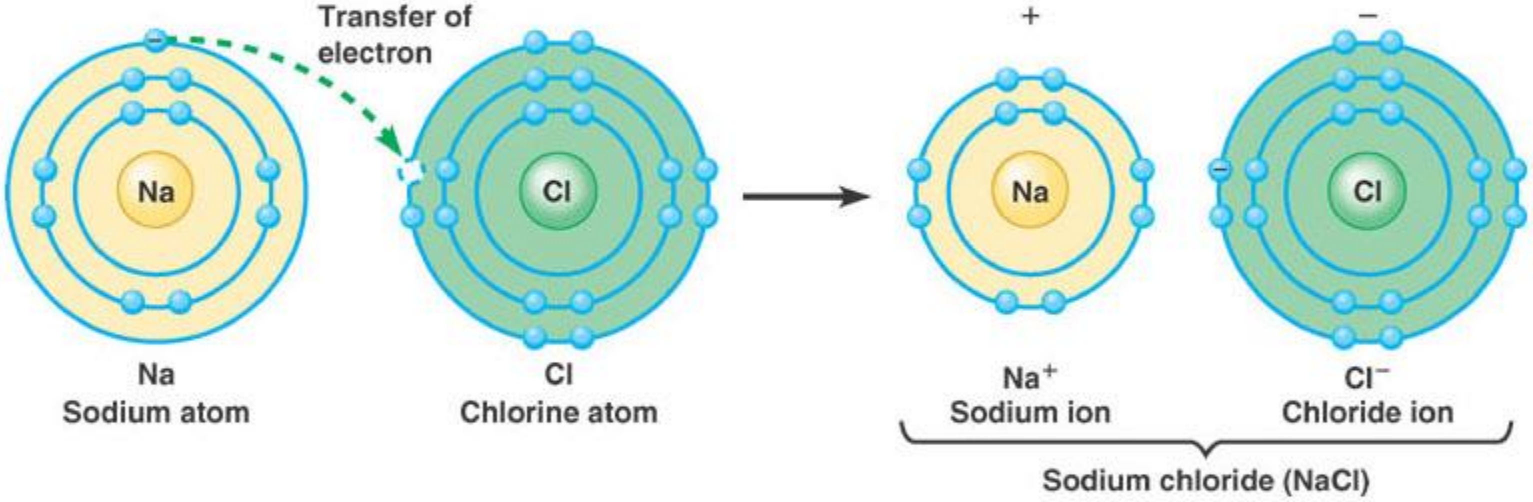
Intramolecular Forces

- Bonds between atoms in a molecule
- 2 Types:
  - Ionic (Metal + Nonmetal)
  - Covalent (Nonmetal + Nonmetal)



Ionic Bonds

- One atom gives an electron to another.
- They now have charges
  - Lost electron = positive ion
  - Gained electron = negative ion
- They are ATTRACTED to each other.

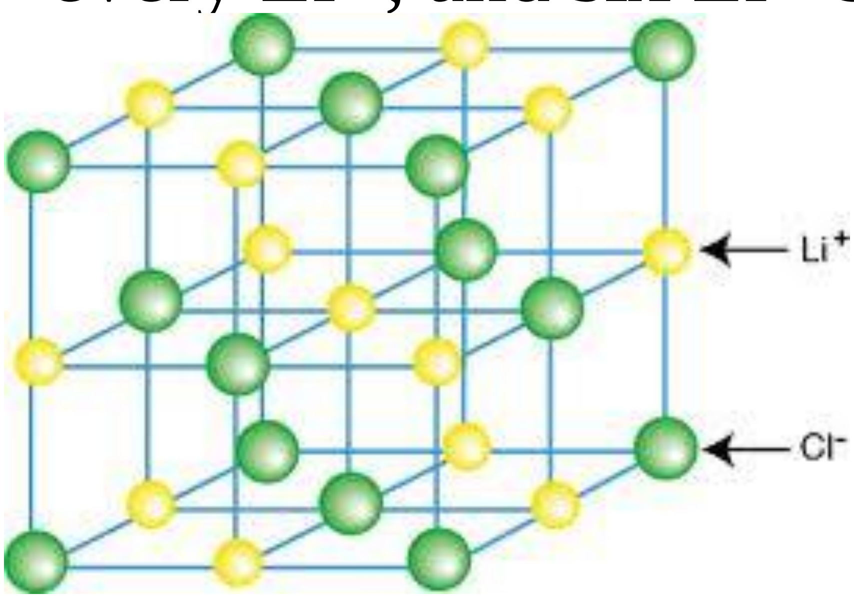


Ionic Bonds

Opposite charges attract in all directions, form a **crystal lattice**

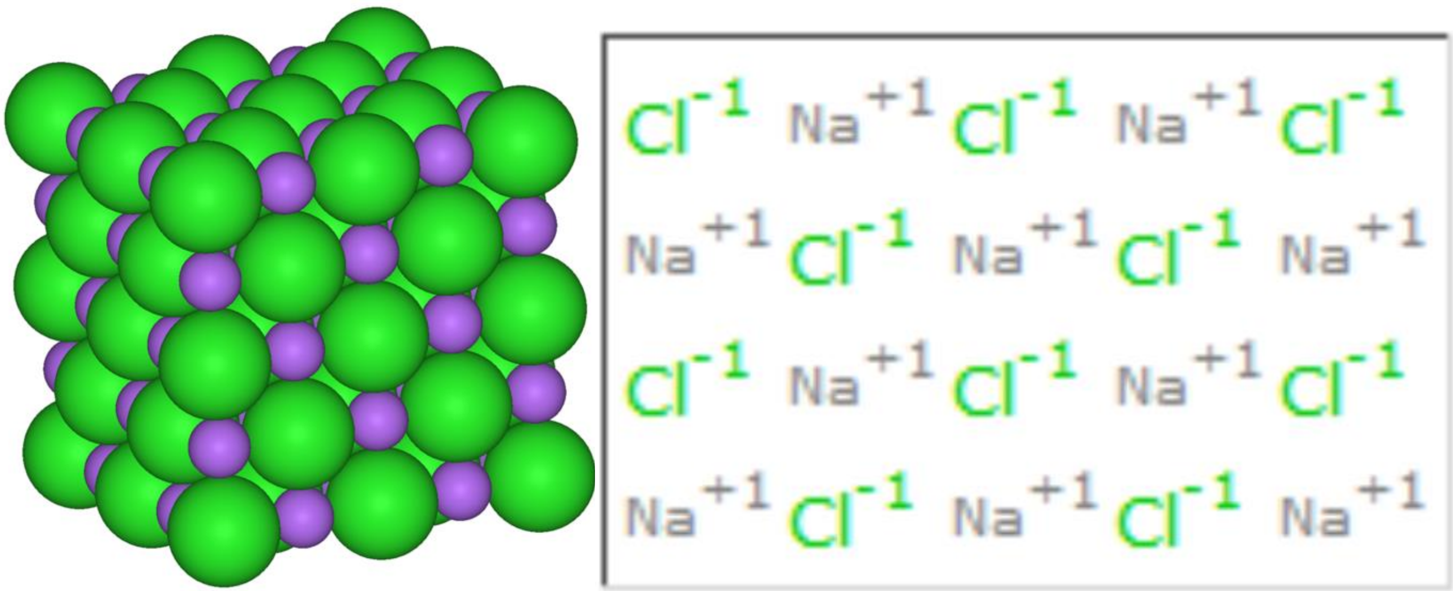
Ex:

LiCl exist as a cube, with **six** Cl- surrounding every Li+, and **six** Li+ surrounding every Cl-



Ionic Compounds are usually **solid** at room temperature.

Crystal lattice



NaCl crystal schematic

Covalent Bonds

- Two atoms **share** their electrons

Polar vs Non-Polar Covalent Bonds

Non-polar covalent bond

- a covalent bond where electronegativity difference between atoms is effectively 0 (zero). **Actually 0.0 - 0.4**

- Means an **equal** sharing of electrons happens.

Polar vs Non-Polar Covalent Bonds

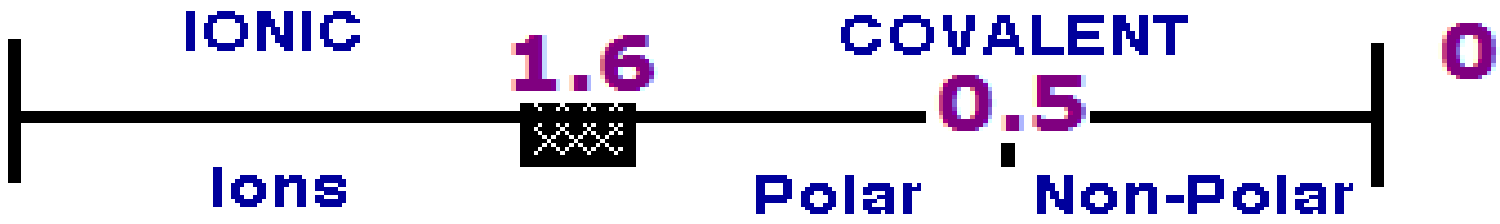
<u>Polar covalent bond</u>

- a covalent bond where electronegativity difference between atoms is more than 0 (zero). ***Actually 0.5 - 1.6***
- **<u>Unequal</u>** sharing of electrons
- What about ***more than 1.6***? **(1.7 – 2.0)**
  - It's **<u>ionic</u>**. Involves ***transfer of electrons***

Covalent Bonds

Polar vs.Non-polar

Difference in Electronegativity

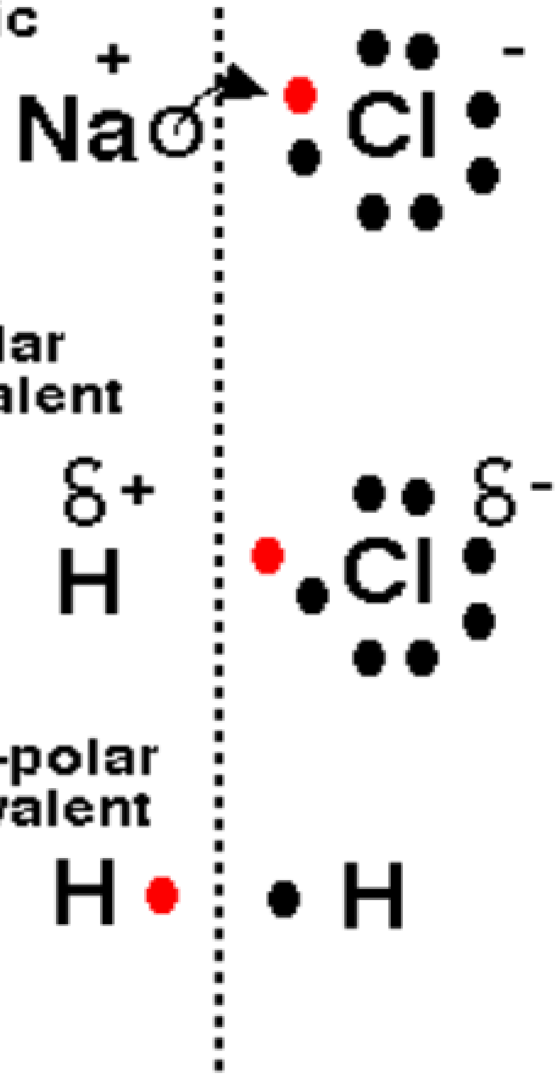


Complete tranfer of electrons	Unequal sharing of electrons	Equal sharing of electrons
Full Ionic Charges	Partial Ionic Charges	No Charges
$Na^{+}Cl^{-}$	$H^{\delta+}\ Cl^{\delta-}$	H - H
metal + non-metal	two different non-metals	two identical non-metals

Electronegativity of elements (dimensionless, Pauling scale **0.7-3.98**)

1A	2A											3A	4A	5A	6A	7A
2.1 H																
1.0 Li	1.5 Be											2.0 B	2.5 C	3.0 N	3.5 O	4.0 F
0.9 Na	1.2 Mg											1.5 Al	1.8 Si	2.1 P	2.5 S	3.0 Cl
0.8 K	1.0 Ca	1.3 Sc	1.5 Ti	1.6 V	1.6 Cr	1.5 Mn	1.8 Fe	1.8 Co	1.8 Ni	1.9 Cu	1.6 Zn	1.6 Ga	1.8 Ge	2.0 As	2.4 Se	2.8 Br
0.8 Rb	1.0 Sr	1.2 Y	1.4 Zr	1.6 Nb	1.8 Mo	1.9 Tc	2.2 Ru	2.2 Rh	2.2 Pd	1.9 Ag	1.7 Cd	1.7 In	1.8 Sn	1.9 Sb	2.1 Te	2.5 I
0.7 Cs	0.9 Ba	1.1–1.2 La–Lu	1.3 Hf	1.5 Ta	1.7 W	1.9 Re	2.2 Os	2.2 Ir	2.2 Pt	2.4 Au	1.9 Hg	1.8 Tl	1.8 Pb	1.9 Bi	2.0 Po	2.2 At
0.7 Fr	0.9 Ra	1.1–1.7 Ac–Lr														

Comparison of Bonding

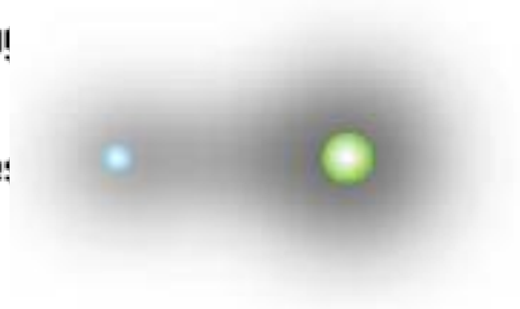


Ionic

complete transfer of electrons - full ionic charges

Polar Covalent

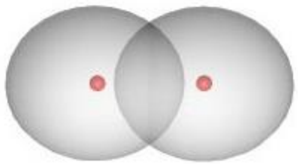
unequal sharing of electrons, results in partial charges



δ = partial

Non-polar Covalent

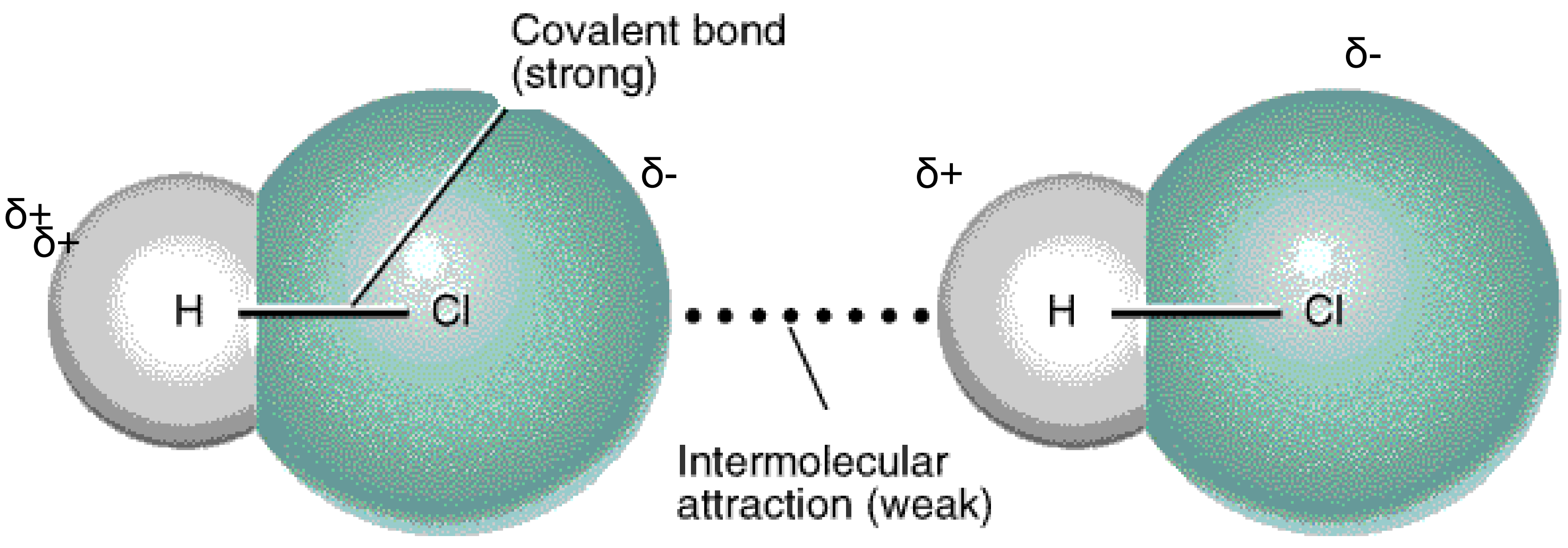
equal sharing of electrons, results in no charges



Intermolecular Forces

- Bonds **between** molecules.

- Not as strong as bonds between atoms inside the molecule.

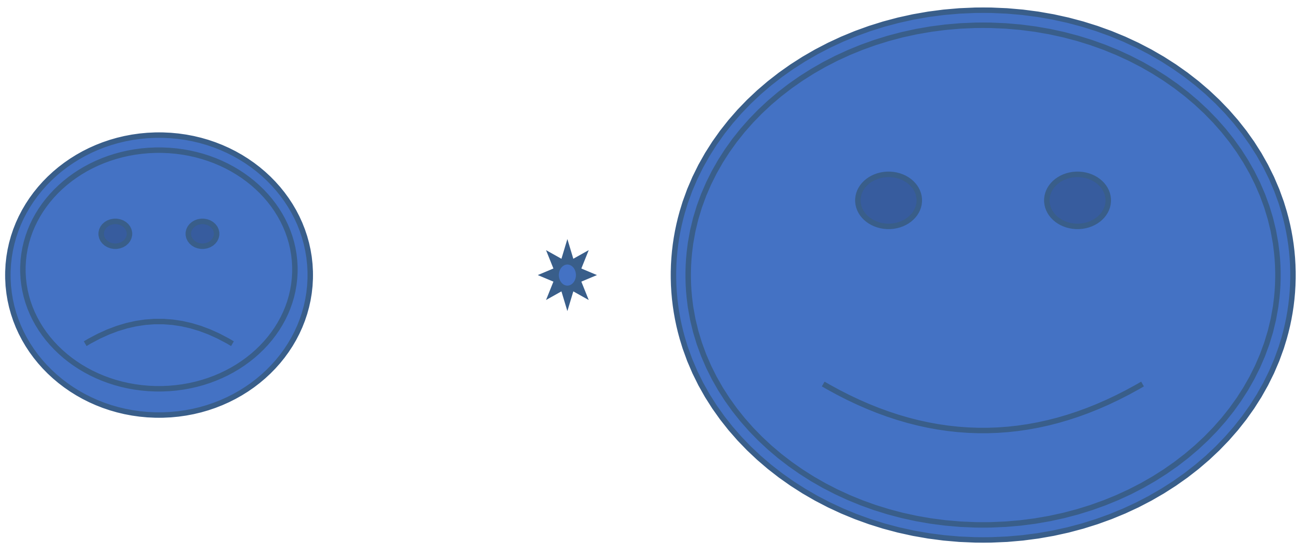


Intermolecular Forces

- Many types, we will learn three
  1. **Dipole-dipole forces**
  2. **Hydrogen bonds**
  3. **London forces**

Dipole-Dipole

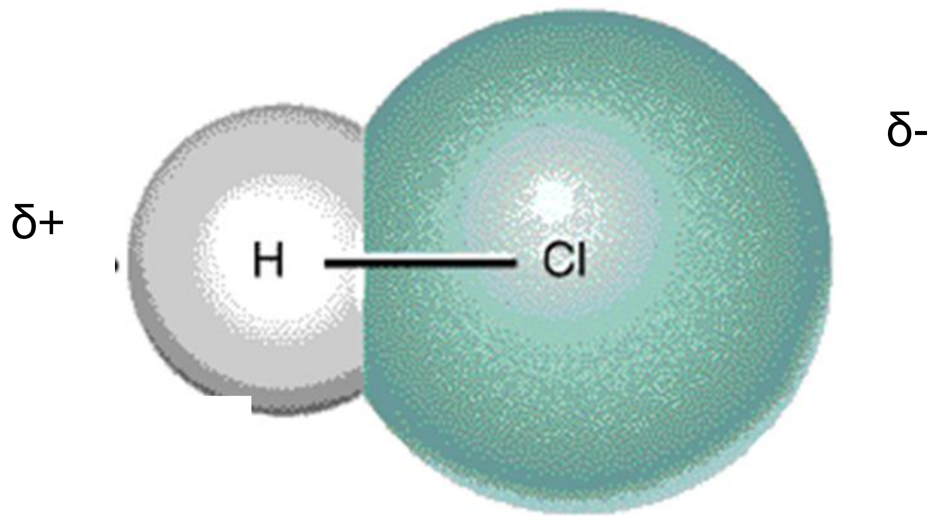
- Results from **<u>uneven sharing</u>** of electrons.
- Some atoms don't share well…
- Maybe one pulls on the electron stronger than the other.




©StephanieDRoeser
www.johnlund.com

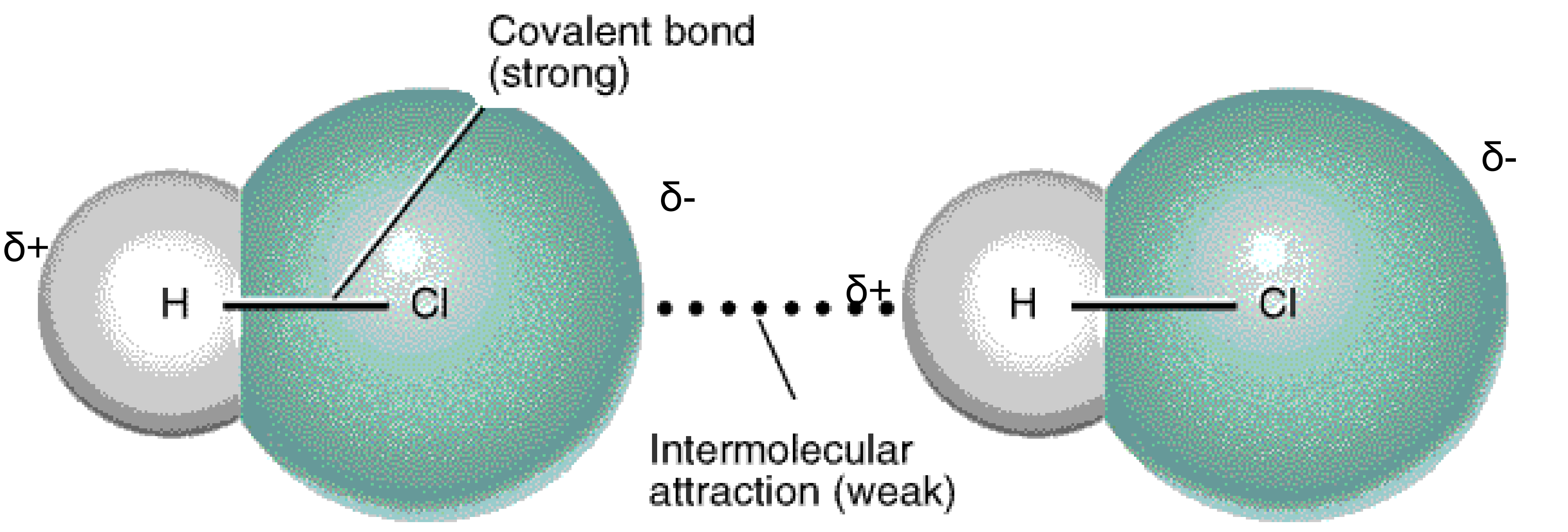
Dipole-Dipole

- Uneven sharing makes one side a little positive and one side a little negative.



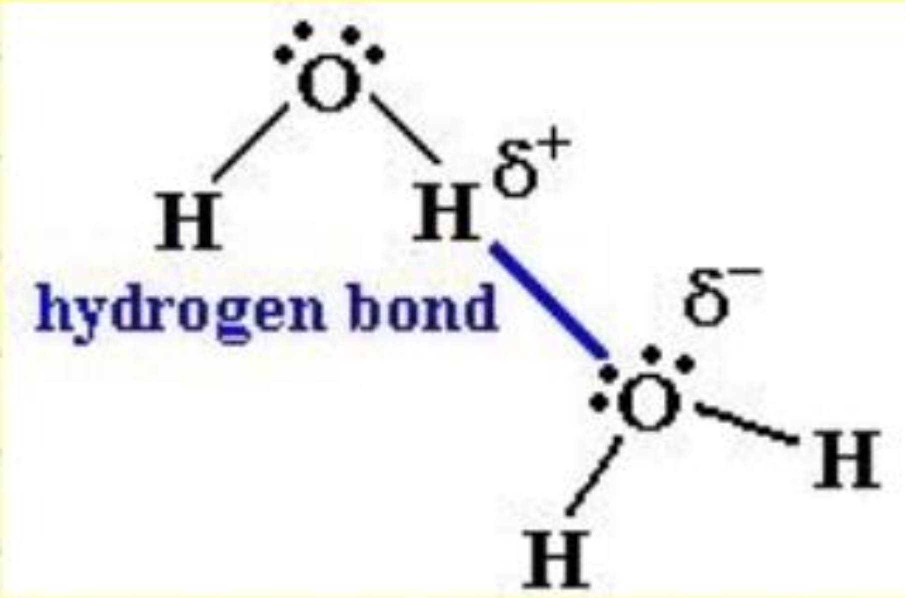
Dipole-Dipole

- The negative dipole of one molecule is attracted to the positive dipole of the other molecule. Cl is slightly negative, H - positive



Hydrogen Bonds

- Special case of dipole forces.
- Attractive force between the H attached to an electronegative atom (O, N, F) of *one* molecule and an electronegative atom of a *different* molecule.



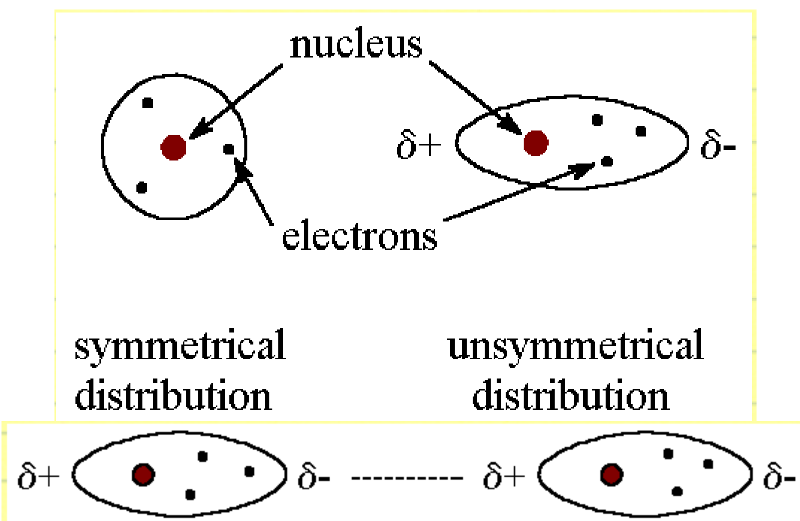
Hydrogen Bonds

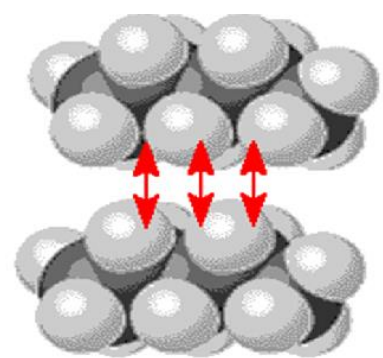
London Dispersion Force

- **Temporary dipoles** caused by the movement of electrons around the nucleus.
- Sometimes the electrons are more to one side, or are **uneven.**
- Makes a **temporary dipole**, and <u>**neighbours**</u> are attracted.

London Dispersion Force

- The <u>**weakest**</u> intermolecular force.
- Cause nonpolar substances to **condense** to liquids and to **freeze** into solids when the temperature is lowered sufficiently.
- Bigger molecule = more **e-** = **stronger London forces.**





How strong?

Covalent bonds > Hydrogen bonding > Dipole-dipole interactions > London forces

400 kcal >	12-16 kcal >	2-0.5 kcal >	less than 1 kcal

From this we can see that normal covalent bonds are almost 40 times the strength of hydrogen bonds.

Covalent bonds are almost 200 times the strength of dipole-dipole forces, and more than 400 times the size of London forces.

How strong?

Comparing Forces:

covalent bonds ≥ ionic bonds >>hydrogen bonds> other dipole-diplole forces> London forces.

Melting point & bonding? How they relate?

- ***stronger chemical bond*** of one atom to another atom – the <u>**more energy**</u> is necessary to **break it**.

- Thus, ***higher*** <u>melting</u> and <u>boiling</u> <u>points</u>.

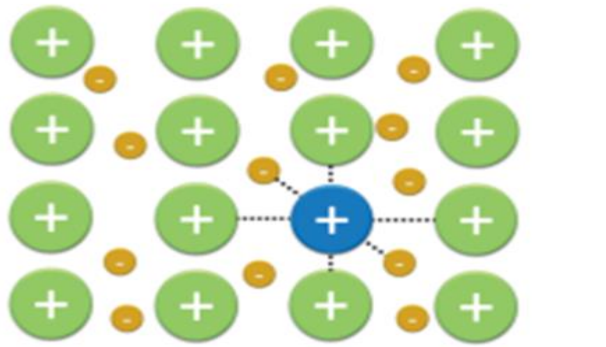
Metallic bonding

- bonds that hold together **pure metals** (e.g. a piece of pure gold, iron, copper, etc.)
- They are usually **similar** value to slightly *lower* than covalent bonds.
- Thus metals have ***similar*** melting and boiling points than non-metals.

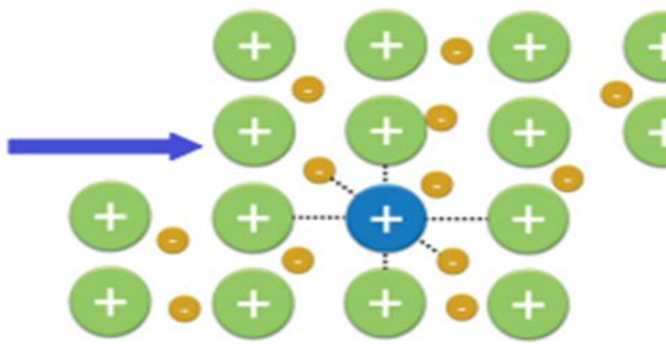
Metallic bonding

METALLIC BONDS

- Model of a sea of electrons
  - Atomic nucleus surrounded from a sea of e^-.
  - Metallic shine .
  - Workability.



Applied force



Applied force

Deformation

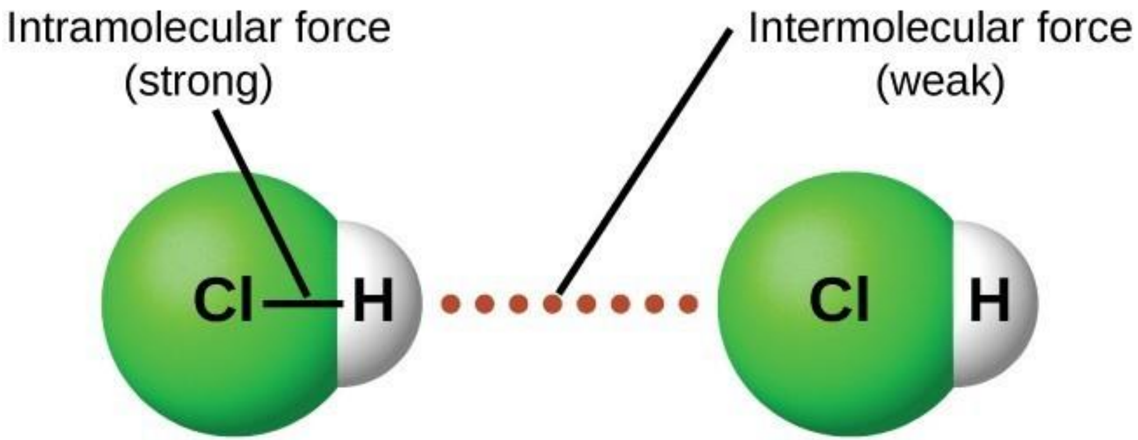
Metallic Bonds Sample Question

- Explain why the **melting point** of metals increases **across a period** and decreases **down** a column with respect to the changes of atomic radius.

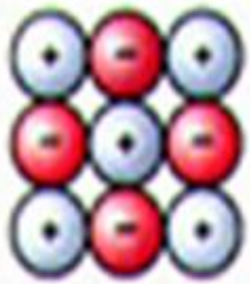
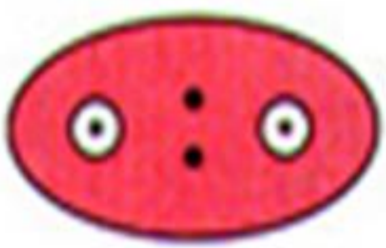
Because of smaller radius, electrons are pulled-up closer and

metallic structure became more compact, need higher force to separate particles.

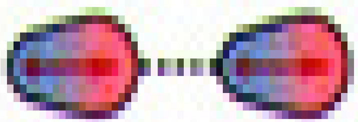
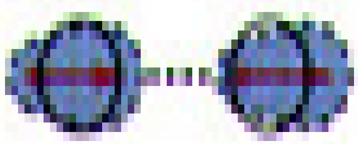
Summary


Intramolecular force
(strong)
Intermolecular force
(weak)
Cl
H
Cl
H

Comparison of the Energies Associated with Bonding (Intramolecular) Forces and Intermolecular Forces

FORCE	MODEL	BASIS OF ATTRACTION	ENERGY (kJ/mol)	EXAMPLE
Intramolecular				
Ionic		Cation–anion	400-4000	NaCl
Covalent		Nuclei–shared e^- pair	150-1100	H–H

Intermolecular forces

H-Bonding	$\overset{\delta^-}{-A}-\overset{\delta^+}{H}\cdots\cdots\overset{\delta^-}{:B}-$	Polar bond to H–dipole charge (high EN of N, O, F)	10-40	$H-O-H\cdots O(H)-H$
Dipole-Dipole		Dipole charges	5-25	$H-Cl\cdots H-Cl$
Dispersion (London)		cloud Polarizable e^- clouds	0.05-40	$F-F\cdots F-F$

Practice

- **Ionic Bonds** - Draw the **Lewis structures** for each atom, then show the transfer of electrons and charge for each ion. Write the **chemical formula** for each compound.

  (1) Mg + Br

  (2) Pb + S

  (3) Al + Cl

- **Covalent Bonds-** Draw the **Lewis structures** for each atom, then draw circles to *show the electrons that are shared*. Write the **chemical formula** for each compound.

  (1) H + Cl

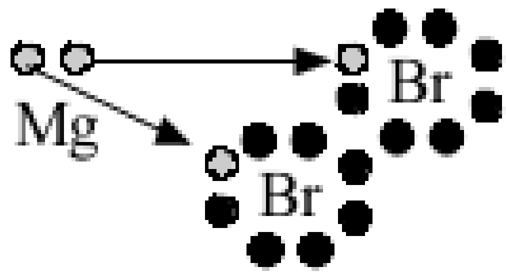
  (2) C + Cl

  (3) Si + O

Practice answers

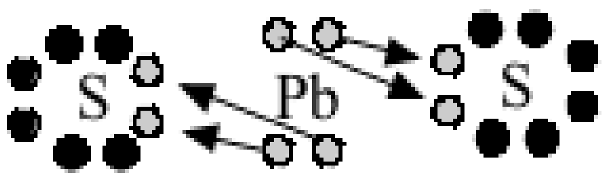
Ionic Bonds

(1) Mg + Br



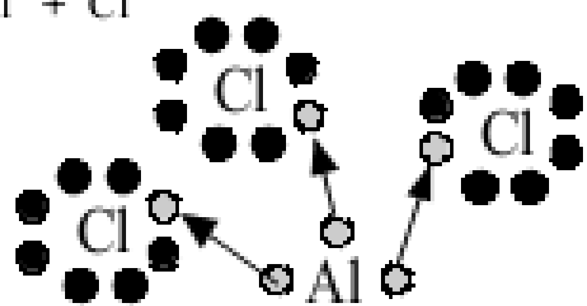
$$Mg^{2+}Br^{1-}_{2} \rightarrow MgBr_{2}$$

(2) Pb + S



$$Pb^{4+}S^{2-}_{2} \longrightarrow PbS_{2}$$

(3) Al + Cl



$$Al^{3+}Cl^{1-}_{3} \longrightarrow AlCl_{3}$$

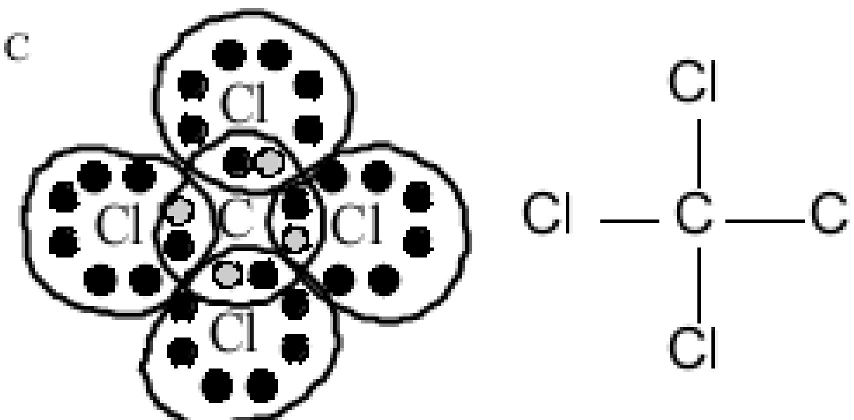
Practice answers

Covalent Bonds

(1) H + Cl

$H-Cl \longrightarrow HCl$

(2) C + C



$\longrightarrow CCl_4$

(3) Si + O

$O=Si=O \longrightarrow SiO_2$

Homework

- **Assignment questions #79-84 page 182-183**
- **Tomorrow 79-80**