Chemical Bonding

Structure Determines Properties!

- A cardinal principle of chemistry is that the macroscopic observed properties of a material are related to its microscopic structure.
  - And visa versa
- The microscopic structure entails depends on:
  - the kinds of atoms
  - the manner in which they are attached
  - their relationship to other molecules, like and dislike
  - the shape of the molecule

Chemical Bonds :*Forces that hold atoms together*

- **1-Ionic bonds** are the forces of attraction between ions.
  - ions formed by electron transfer(Na^{+1},,Cl^{-1})
  - electrostatic forces
- **2-Covalent bonds** are the forces of attraction between two atoms which are sharing electrons.

1-Ionic Bonds

- Results from reaction between Metal and Nonmetal
- Metal loses electrons to form cation, Nonmetal gains electrons to form anion
- Ionic bond is the attraction between a positive ion and negative ion
- Larger Charge (+2,+3..-2,-3..)= Stronger Attraction
- Smaller Ion(Li^{+1}) = Stronger Attraction
- ***No bond is 100% ionic!!***
- Electrostatic attraction nondirectional
  - no direct anion-cation pair, **No ionic molecule**
    - chemical formula is empirical formula, simply giving the ratio of ions based on charge balance ,eg.(Li_2O)
- Ions arranged in a pattern called a **crystal lattice**
    - Represent maximizes attractions between + and - ions

2-Covalent Bonds

- Typical of molecular substances
- Atoms bond together to form molecules
  - strong attraction
- **Sharing pairs of electrons**
- Molecules attracted to each other weakly
- Often found between nonmetal atoms

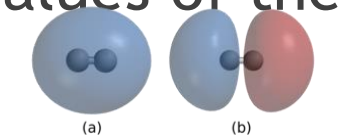
3-Bond Polarity

- Covalent bonding between unlike atoms results in unequal sharing of the electrons
  - One end of the bond has larger electron density than the other
- The result is **bond polarity**
  - The end with the larger electron density gets a partial negative charge
  - The end that is electron deficient gets a partial positive charge



Dipole Moment

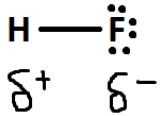
- Bond polarity results in an unequal electron distribution, resulting in areas of partial positive and partial negative charge
- the difference of the electronegativity values of the molecules.



- 0.0-0.3, then the molecule has a non-polar bond.

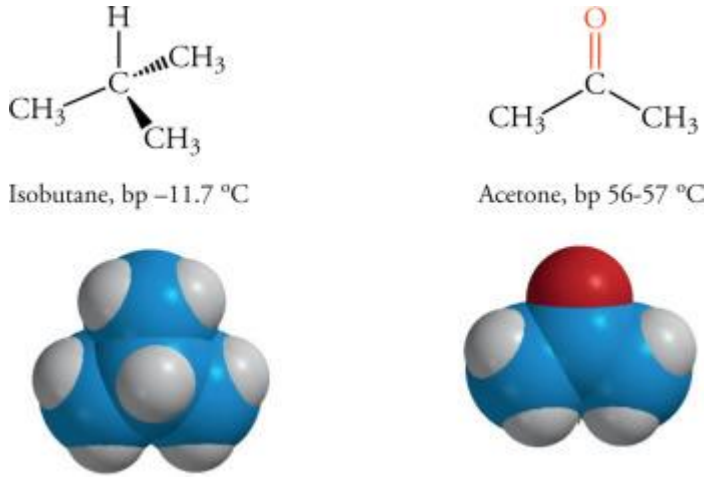
  0.3-1.7, then the molecule has a polar bond.

  1.7 or more, then the molecule has an ionic bond



- Any molecule that has a center of positive charge and a center of negative charge in different points is said to have a **dipole moment**

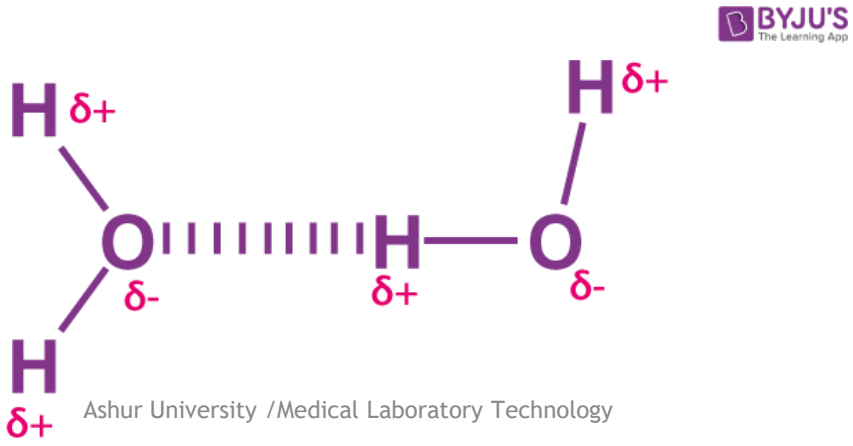
- If a molecule has more than one polar covalent bond, the areas of partial negative and positive charge for each bond will partially add to or cancel out each other.



- The end result will be a molecule with one center of positive charge and one center of negative charge
- The dipole moment effects the attractive forces between molecules and therefore the physical properties of the substance

4-Hydrogen Bonding

Compared to ionic and covalent bonding, Hydrogen bonding is a weaker form of chemical bonding. It is a type of polar covalent bonding between oxygen and hydrogen, wherein the hydrogen develops a partial positive charge. This implies that the electrons are pulled closer to the more electronegative oxygen atom.



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Electronegativity

- Measure of the ability of an atom to attract shared electrons
  - Larger electronegativity means atom attracts more strongly
  - Values 0.7 to 4.0
- Increases across period (left to right) on Periodic Table
- Decreases down group (top to bottom) on Periodic Table
- Larger difference in electronegativities means more polar bond
  - negative end directed towards more electronegative atom

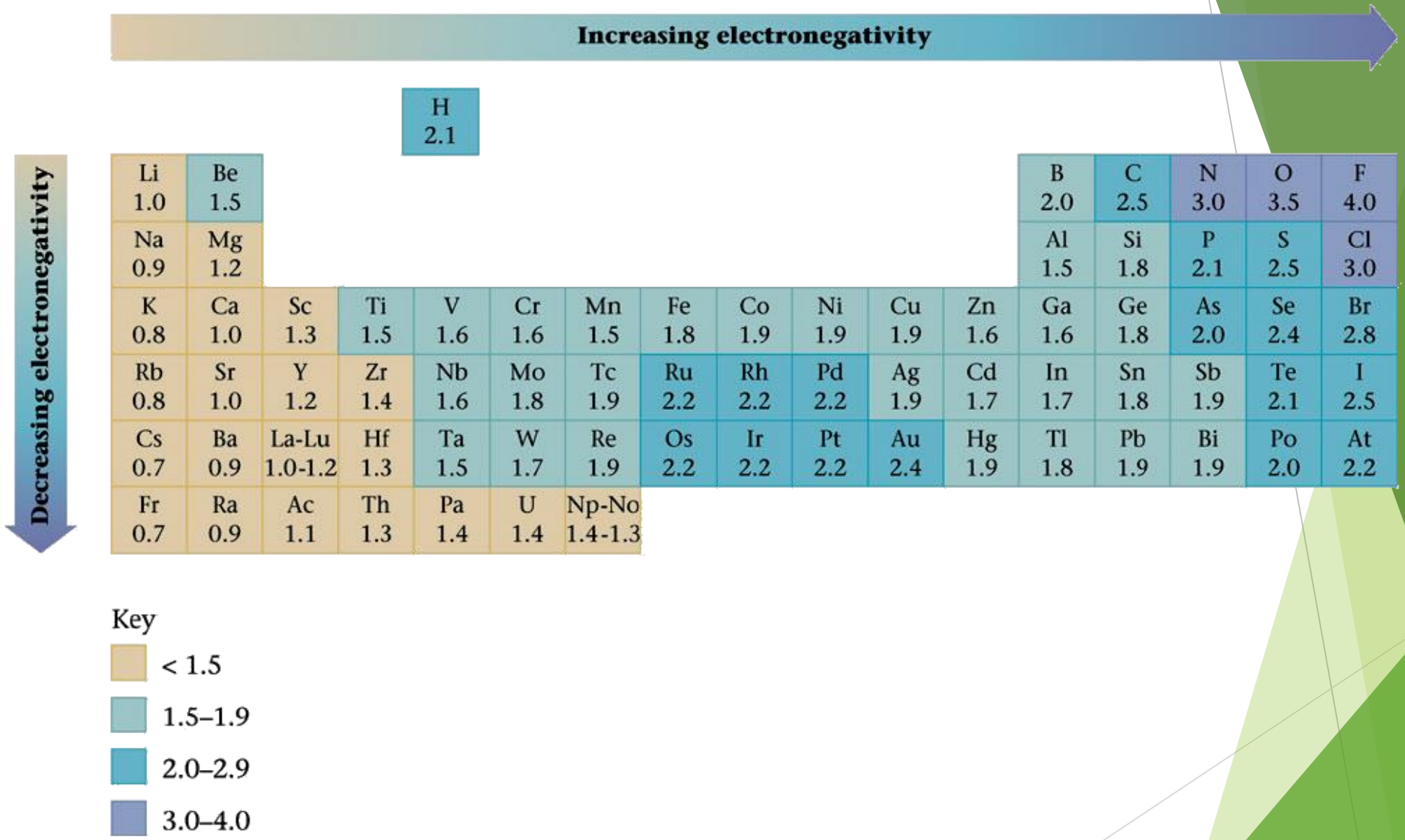

Increasing electronegativity
Decreasing electronegativity
H 2.1
Li 1.0
Be 1.5
B 2.0
C 2.5
N 3.0
O 3.5
F 4.0
Na 0.9
Mg 1.2
Al 1.5
Si 1.8
P 2.1
S 2.5
Cl 3.0
K 0.8
Ca 1.0
Sc 1.3
Ti 1.5
V 1.6
Cr 1.6
Mn 1.5
Fe 1.8
Co 1.9
Ni 1.9
Cu 1.9
Zn 1.6
Ga 1.6
Ge 1.8
As 2.0
Se 2.4
Br 2.8
Rb 0.8
Sr 1.0
Y 1.2
Zr 1.4
Nb 1.6
Mo 1.8
Tc 1.9
Ru 2.2
Rh 2.2
Pd 2.2
Ag 1.9
Cd 1.7
In 1.7
Sn 1.8
Sb 1.9
Te 2.1
I 2.5
Cs 0.7
Ba 0.9
La-Lu 1.0-1.2
Hf 1.3
Ta 1.5
W 1.7
Re 1.9
Os 2.2
Ir 2.2
Pt 2.2
Au 2.4
Hg 1.9
Tl 1.8
Pb 1.9
Bi 1.9
Po 2.0
At 2.2
Fr 0.7
Ra 0.9
Ac 1.1
Th 1.3
Pa 1.4
U 1.4
Np-No 1.4-1.3
Key
< 1.5
1.5–1.9
2.0–2.9
3.0–4.0

Table 11.1 The Relationship Between Electronegativity and Bond Type

Electronegativity Difference Between the Bonding Atoms	Bond Type	Covalent Character	Ionic Character
Zero ↓ Intermediate ↓ Large	Covalent ↓ Polar covalent ↓ Ionic	Decreases ↓	Increases ↓

The three possible types of bonds:

(a) a covalent bond formed between identical atoms; e.g.O2,F2…

(b) a polar covalent bond, with both ionic and covalent components; e.g. HCl,HI…

(c) ionic bond, with no electron sharing..eg. NaCl..

Electron Arrangements And Ion Charge

- We know
  - Group 1A metals form ions with +1 charge
  - Group 2A metals form ions with +2 charge
  - Group 7A nonmetals form ions with -1 charge
  - Group 6A nonmetals form ions with -2 charge
  - Group 8A nonmetals do not form ions, in fact they are extremely unreactive

Electron Arrangements and Ion Charge

- Representative Metals form cations by losing enough electrons to get the same electron configuration as the previous noble gas
- Nonmetals form anions by gaining enough electrons to get the same electron configuration as the next noble gas

Atom	Atoms Electron Config	Ion	Ions Electron Config
Na	$[Ne]3s^1$	Na^{+1}	[Ne]
Mg	$[Ne]3s^2$	Mg^{+2}	[Ne]
Al	$[Ne]3s^23p^1$	Al^{+3}	[Ne]
O	$[He]2s^22p^4$	O^{-2}	[Ne]
F	$[He]2s^22p^5$	F^{-1}	[Ne]

Electron Arrangements and Ionic Bonding

- Representative metals lose their valence electrons to form cations
- Nonmetals gain electrons so their valence shell has the same electron arrangement as the next noble gas
- *There have to be enough electrons from the metals atoms to supply the needed electrons for the nonmetal atoms*
  - Allows us to predict the formulas of ionic compounds
- In Polyatomic ions, the atoms in the ion are connected with covalent bonds. The ions are attracted to oppositely charged ions to form an ionic compound. e.g. $CaCO_3$

Properties of Ionic Compounds

1. All solids at room temperature

   Melting points greater than 300°C

2. Liquid state conducts electricity, solid state does not

   Liquid = molten

3. Brittle and Hard

4. Often soluble in water, and when dissolved the solution becomes an electrical conductor

   1. When ionic compounds containing polyatomic ions dissolve, the covalent bonds holding the polyatomic ion do not break, the ion stays together even though it separates from the other ion.

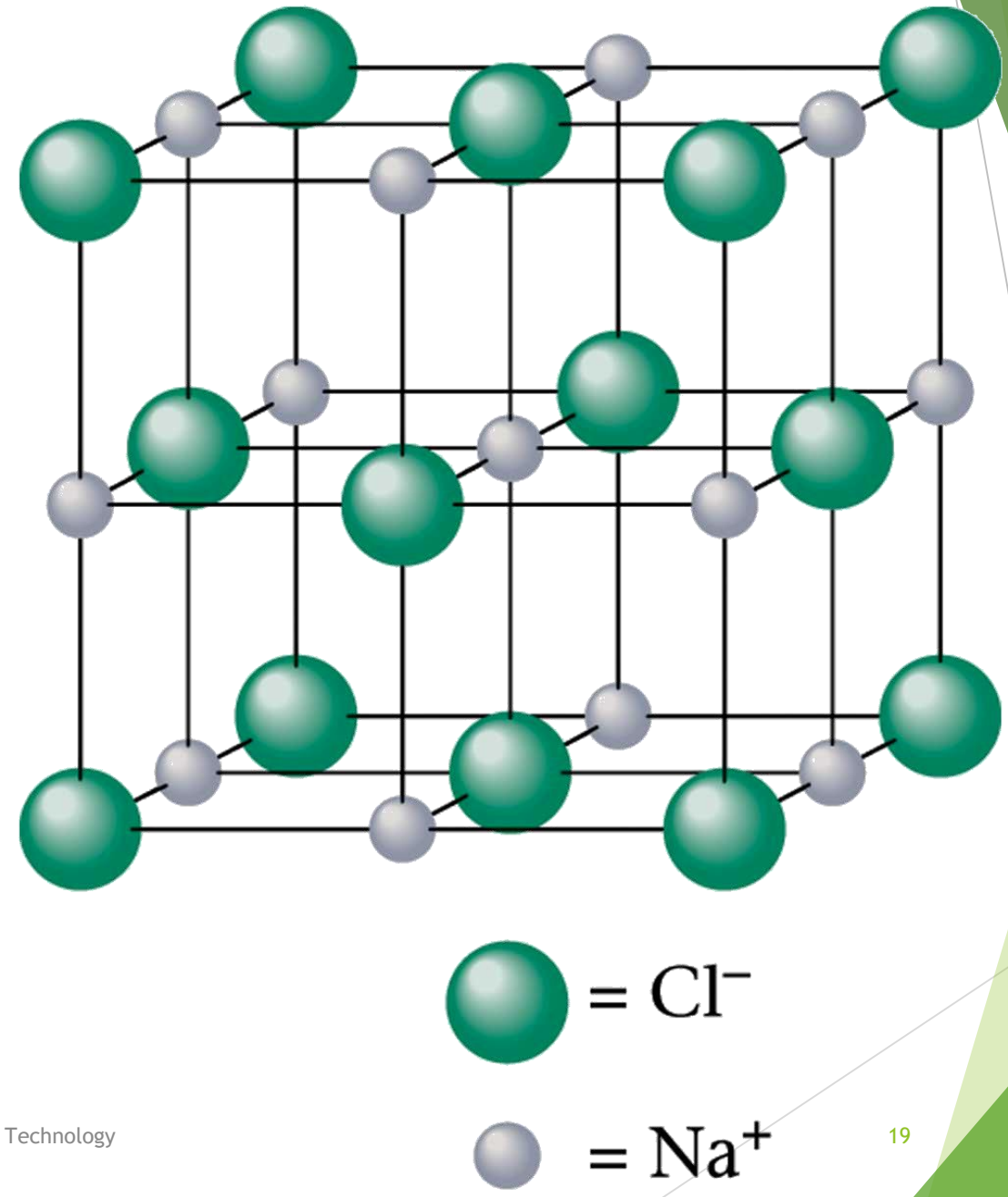
      e.g Na_2SO_4 == $2Na^{+1} + SO_4^{-2}$

   1. All strong electrolytes

Bonding and Structure of Ionic Compounds

- **Crystal Lattice** :- Geometric pattern determined by the size and charge of the ions
  - Anions larger than cation
  - Anions > parent atom ….. Cations < parent atom
- Anions generally considered “hard” spheres packed as efficiently as possible, with the cations occupying the “holes” in the packing
- Arrangement results in each cation being surrounded by as many anions as will fit
  - And visa versa
  - **The arrangement Maximizes attractions between ions**

The regular arrangement of sodium and chloride ions in sodium chloride, a crystalline solid.



End