9.) - hydrogen bonding, dipole-dipole, and London force.
- London holds non-polar together weakly.
- hydrogen bonding is strongest.

10.) metals are malleable/ductile because metallic bonds allow slipping or movement between atoms and ionic bonds do not.

Chemical Bonding Review

This review is only intended to remind you of the type of questions which could be asked. It is your responsibility to study ALL notes and assignments in order to fully understand this unit.

Topic	Learning Outcomes
Bonding and Molecular Shape	It is expected that students will: • define covalent and ionic bonding • define valence electrons • demonstrate a knowledge that bonding involves valence electrons • draw an electron dot diagram for an atom • identify from a chemical formula the probable type of bond (ionic or covalent) • draw electron dot diagrams and structural formulae for simple molecules and ions and deduce molecular formulae

1a.) Ionic bond is attractive force from opposite electrostatic attraction. Covalent bond is attractive force due to sharing of e⁻(s).

1. What is the difference between an ionic and covalent bond and an ionic or covalent compound?
2. Compare CH_3F and $MgCl_2$ in terms of:
   a. Type of bonding
   b. Structure/breakability
   c. Melting point/boiling point
   d. Electrical conductivity
3. How many valence e- does Cl have? How about Na^+?
4. Draw a Lewis dot diagram for Ca, Ca^{2+}, F, and S.
5. What is the difference between polar and non-polar covalent bonds? Is CO a polar or non-polar BOND?
6. What is a polyatomic ion? Give an example.
7. Draw CH_2O. Is it a polar MOLECULE? ~~Name the shape of the molecule.~~
8. Compare/Draw/name the shape of H_2S (BP= -85.6°C) and H_2O ((BP= 100°C)to show their polarity. Explain why they boil at such different temperatures.
9. What are the three types of Intermolecular forces? Which is the only one which holds non-polar molecules together weakly? Which is the strongest?
10. Describe why metals are ductile/malleable but ionic solids shatter. ~~Explain why solid metals conduct electricity but ionic solids don~~'t.

1b.) Ionic = crystalline structure / covalent not. (brittle)

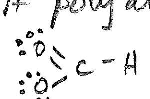
2.) CH_3F = Polar covalent, non crystalline and brittle, high melting/boiling point, non-conductive.
$MgCl_2$ = Ionic bond, crystalline structure, relatively low melting/boiling point, conductive.

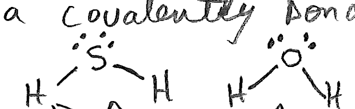
3.) Cl = 7 valence e^- / Na^+ = 0 valence e^-

4.) Ca = $\ddot{Ca}$ / Ca^{+2} = $[Ca]^{+2}$ / F = :F: / S = ·S:

5.) Polar = Uneven sharing and non-polar covalent = even sharing. CO = Polar covalent

6.) A polyatomic ion is a covalently bonded compound with a charge. PO_4^{-3}

7.)  - Polar!

8.) 

non-polar covalent / Polar covalent

- both bent or angular shape
← polar / - water has hydrogen bonding due to polar nature.
- dihydrogen sulphide in not polar.

SOLUBILITY (POLAR VS. NONPOLAR)

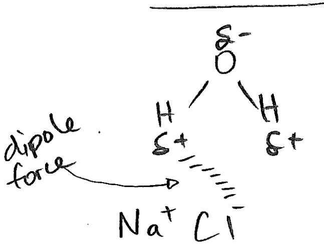
Generally, "like dissolves like." Polar molecules dissolve other polar molecules and ionic compounds. Nonpolar molecules dissolve other nonpolar molecules. Alcohols, which have characteristics of both, tend to dissolve in both types of solvents, but will not dissolve ionic solids.

Check the appropriate columns as to whether the solute is soluble in a polar or nonpolar solvent.

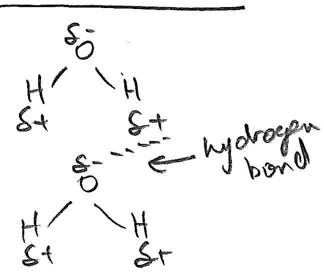
SOLUTES	SOLVENTS Water	 CCl_4	 Alcohol
1. NaCl	✓		
2. I_2		✓	✓
(alcohol) 3. ethanol (CH_3CH_2OH)	✓	✓	✓
4. benzene all C-H bonds		✓	✓
5. Br_2		✓	✓
6. KNO_3	✓		
7. toluene All C-H bonds		✓	✓
8. $Ca(OH)_2$	✓		

Why? (Explain using the concepts of electrostatic attraction and polarity)

Ionic to polar

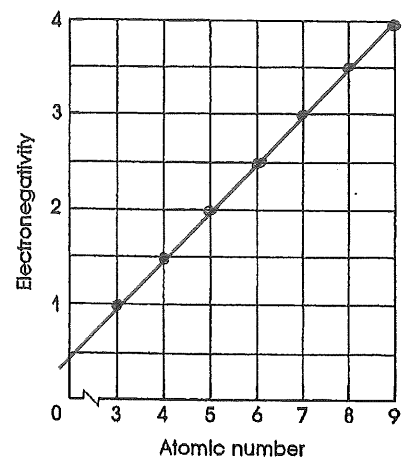


Polar to Polar



2. Using the grid below, graph electronegativity versus atomic number for the elements in Period 2. How does electronegativity vary with the atomic number across a period?

electronegativity increases across a period



Account for this fact.

more protons in nucleus to pull electrons in closer as one goes across the period.

Why is no electronegativity shown for element 10?

noble gases have full p-orbitals, resulting in stability. As such no e⁻ are needed and none are allowed to be taken.

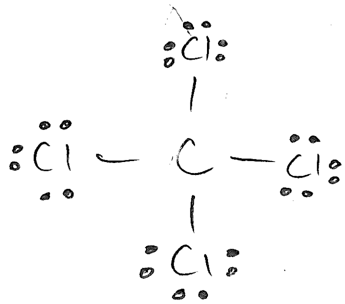
B. Polarity of Bonds

Differences in electronegativity can be used to determine how polar a bond is between two atoms. If the difference in the electronegativities of the atoms is 0.2 or less, the bond is considered to be nonpolar covalent, and the electron sharing is more or less equal. If the difference is more than 0.2 but less than 1.7, the bond is polar covalent, which means that the sharing is unequal. If the difference is greater than 1.7, the bond is considered to be ionic, and the bonding electrons are essentially transferred to one of the atoms.

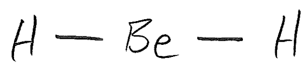
Using the table of electronegativities from Part A, calculate the electronegativity difference for the atoms that are bonded in the following diatomic molecules. Then tell whether the bond is nonpolar covalent, polar covalent, or ionic. Also, tell which atom has the greater share of the bonding electrons.

FORMULA	ELECTRONEGATIVITY DIFFERENCE	TYPE OF BOND	ATOM WITH GREATER ELECTRON SHARE
NO	0.5	PC	Nitrogen
MgO	2.0	I	Oxygen.
Br_2	0	C	N/A
LiH	1.1	PC	Hydrogen.
LiBr	1.8	I	Bromine.
CuF	2.1	I	Fluorine
CO	1.0	PC	Oxygen
HAt	0.1	C	Astatine.

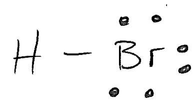
CCl_4



BeH_2



HBr



E. Intermolecular Forces

The forces between molecules include dipole-dipole forces (for polar molecules), London forces (for nonpolar molecules), and hydrogen bonding (for molecules in which hydrogen is attached to fluorine, oxygen, or nitrogen). Indicate, for each of the following substances, the intermolecular forces involved.

1. H_2 London
2. Ne London.
3. CH_2Cl_2 - Polar dipde-dipole
4. NH_3 hydrogen bonding.
5. HCl - Polar dipole-dipole
6. H_2O - Polar hydrogen bonding
7. NF_3 London.