Chemical (Potential) energy is stored in bonds. * It always takes energy to break a bond
* Energy is always released when a new bond is formed

mainly soluble in H_2O
hard substances
very high mp
electrolytes
crystalline solids

Ionic

* Transfer of e^- when one element has a much smaller EN than the other (diff ≥ 1.7)
* Forms Salts

① Transfer of e^-
② Balance charges to make compounds
③ Tend to be very hard solids w/ high melting points

Insoluble
metallic properties
variable mp's

metallic

sea of e^-

Bonding
⇒ What holds atoms together

many are insoluble in H_2O
low mp / decompose
nonelectrolytes
Soft

covalent (intramolecular)

* Sharing electrons b/w atoms with similar electronegativities
* Forms molecules

① Sharing e^-
EVENLY = nonpolar
UNEVENLY = Polar
② All electrons must be paired
③ Tend to have low melting points (many decompose w/ heat)

Review of dot diagrams of Salts

NaCl $Na^+[:\ddot{Cl}:]^{-1}$

CaS $Ca^{+2}[:\ddot{S}:]^{-2}$

CaF_2 Ca^{+2} $[:\ddot{F}:]^{-1}$ $[:\ddot{F}:]^{-1}$

Na_3P Na^+ Na^+ Na^+ $[:\ddot{P}:]^{-3}$

Now we need to include polyatomic ions

$NaNO_3$ $Na^+[\ddot{O}-N(=O)-\ddot{O}]^{-1}$

Molecules have Lewis Dot Diagrams w/ lines in them
① Determine total # of electrons
② Draw the involved atoms
③ Connect atoms using single bonds ($2e^-$)
④ Place remaining electrons around noncentral atoms to complete their octets THEN place any extra electrons on central atom.
⑤ Use multiple bonds (double or triple) to complete octets as necessary.
⑥ Use Resonance to minimize Formal Charges and show possible structures.

Examples

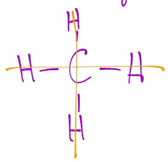
CO_2 $:\ddot{O}=C=\ddot{O}:$
NH_3 $H-\ddot{N}-H$ (H)
H_2 $H-H$
Br_2 $:\ddot{Br}-\ddot{Br}:$
H_2O $H-\ddot{O}:$ (H)
O_2 $:\ddot{O}=\ddot{O}:$
N_2 $:N\equiv N:$
C_2H_6 H_3C-CH_3
C_2H_4 $H_2C=CH_2$
C_2H_2 $H-C\equiv C-H$
H_2CO $H_2C=\ddot{O}:$
ClO^{-1} $[:\ddot{Cl}-\ddot{O}:]^{-1}$
CO_3^{-2} $[\ddot{O}-C(=O)-\ddot{O}]^{-2}$

Just interesting
I_3^- $:\ddot{I}-\ddot{I}-\ddot{I}:$
XeF_4 $F_2>\ddot{X}e<F_2$
PCl_5 $Cl_2>P(Cl)_3$

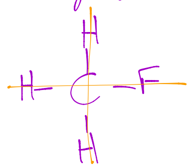
Molecules can be polar or nonpolar as well -
It depends on their overall symmetry
We need to understand the possible molecular shapes

① tetrahedron: 4 atoms ⊕ 1 central CH_4 CH_3F may be polar or nonpolar molecules
② trigonal pyramid: 3 atoms / 1 lone pair ⊕ 1 central NH_3 must be polar molecules
③ bent: 2 atoms / 2 lone pairs ⊕ 1 central H_2O must be polar molecules
④ trigonal planar: 3 atoms / ∅ lone pairs ⊕ 1 central H_2CO CO_3^{-2} BF_3 may be polar or nonpolar molecules
⑤ linear: 2 atoms / ∅ lone pairs ⊕ 1 central CO_2 HCN may be polar or nonpolar molecules

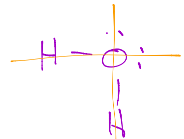
For these shapes we can use a simple 2D test to determine the polarity of 3D molecules
Using the stick model diagrams above for each shape, identify symmetry about the central atom



tetrahedron
symmetrical about both axes
∴ NONPOLAR molecule

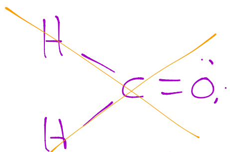


tetrahedron
asymmetrical about the vertical axis
∴ POLAR molecule

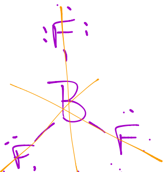


bent
asymmetrical about both axes
∴ POLAR molecule

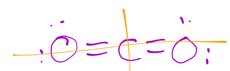
trigonal pyramid
asymmetrical about horizontal axis
∴ POLAR molecule



trigonal planar
asymmetrical about both axes
∴ POLAR molecule



trigonal planar
symmetrical about all axes
∴ NONPOLAR molecule



linear
symmetrical on both axes
∴ NONPOLAR molecule

How does that happen? From where do the shapes come? HYBRIDIZATION

2 or more atomic orbitals combine to form an equal number of EQUIVALENT hybrid molecular orbitals

1 "s" + 3 "p" = 4 "sp^3" orbitals all tetrahedron, trigonal pyramid, bent

1 "s" + 2 "p" = 3 "sp^2" orbitals trigonal planar

1 "s" + 1 "p" = 2 "sp" orbitals linear

The number of directions electrons come off central atom = the # of orbitals needed

tetrahedron, trigonal pyramid, bent: sp^3 — 4 directions, 4 orbitals

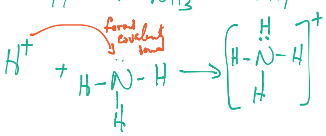
trigonal planar: sp^2 — 3 directions, 3 orbitals

linear: sp — 2 directions, 2 orbitals

Special Types of Covalent Bonding

① Coordinate Covalent Bonding

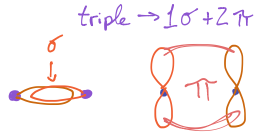
$H^+ + NH_3 \rightarrow NH_4^+$



Single, double, triple bonds are made from σ + π bonds:

Single → 1 σ bond
double → 1 σ + 1 π bond
triple → 1 σ + 2 π bonds

σ = occur on the axis b/w atoms
π = go around the axis b/w atoms



② Network Covalent Bonds: like in diamond. A VERY STRONG network of interconnected covalent bonds forms macromolecules - diamond, sand.

Bond Energy   Bond   Types of bonding

What does it say about bond length and the strength of bonds?

Shorter bonds tend to be stronger: It'll take more energy to break them

The more bonds there are b/w two atoms, the shorter ⊕ stronger it will be.

Strength Single < Double < triple

346 kJ/mol	C—C	1σ
610 kJ/mol	C=C	1σ ⊕ 1π
835 kJ/mol	C≡C	1σ ⊕ 2π

Intermolecular Forces: What causes a molecule to be attracted to another particle?

ⓐ It's always about CHARGE

ⓑ opposites attract

ⓒ Some things:

- always have a charge = ions
- always have a partial charge = polar
- Can possibly have a charge sometimes: nonpolar

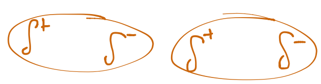
BEWARE

Sometimes, vdW is mistakenly used to describe nonpolar interaction

Intermolecular forces

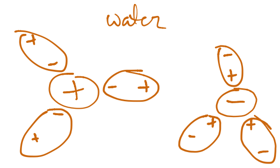
van de Waals forces (not covalent or ionic)

London Dispersion forces
- forces b/w nonpolar molecules
- induced dipole
- momentary dipole



Dipole - Dipole
- attraction b/w polar molecules
- a special case is hydrogen bonding
  - HNOF
  - accounts for special properties of H_2O

Ion - Dipole
- attraction b/w an ion and a polar molecule
- dissolving a salt in water



Effects of intermolecular forces:

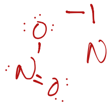
MP Stronger → higher

BP Stronger → higher

P_{vap} Stronger → lower

Summary of Bonding Theories: Do Not Try To memorize this Chart. You need to LEARN how it WORKS. These are not all of the possibilities, just most of them. If you try to memorize this without understanding how these different theories work together, you will get tripped up on the test.

Example	Lewis Dot	VSEPR (Shapes)	σ π l.p.	hybridization	Bond ∠
CH_4, CH_3Cl	−X− (4 bonds)	tetrahedron	4 σ	sp^3	109.5°
NH_3, PH_3	−Ẍ−	trigonal pyramid	3 σ ⊕ 1 l.p.	sp^3	< 109.5
H_2O, H_2S	−Ẍ:	Bent	2 σ ⊕ 2 l.p.	sp^3	<< 109.5°
H_2CO, BF_3	X (3 bonds)	trigonal planar	3 σ ⊕ Ø π ⊕ Ø l.p. 3σ ⊕ 1π ⊕ Ø l.p.	sp^2	120°
O_2	:Ẋ=	linear	1σ ⊕ 1π ⊕ 2 l.p.	sp^2	N/A

 NO_2^{-1}	$:X$	bent	$2\sigma \oplus 1\pi \oplus 1$ l.p.	sp^2	$\angle 120°$
CO_2, HCN	$-X-$	Linear	$2\sigma \oplus 2\pi$ $\varnothing$ l.p.	sp	$180°$
N_2	$:X\equiv$	linear	$1\sigma \oplus 2\pi \oplus 1$ l.p.	sp	N/A