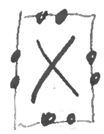
Ch. 8 Bonding

(Starting in the middle of the chapter)

Lewis Structures

1. Neutral atoms or ions showing valence e^-s 
☆ 2. Molecules with shared e^-s

Key Things to Know

See also, the Learning Targets

- Drawing Lewis Structures
  - Formal charge
  - Resonance
- Lattice Energy
- Bond Polarity
- Bond Enthalpy & ΔH_{rxn}

We base drawings on the number of valence electrons. The no. of val. e^-s depends on the group number.

Group	1	2	13	14	15	16	17	18
No. of Val. e^-s	1	2	3	4	5	6	7	8

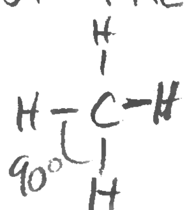
Electrons in drawings are nearly always in pairs:

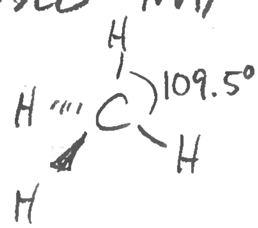
Non-bonding pair/lone pair: X:

Bonding pair: X—Y (X•—•Y) (X:Y)

Multiple bonds: X=Y double ($4e^-$s); X≡Y triple ($6e^-$s)

As we draw molecules, we'll use right angles but the real angles may vary!

2D 
(not real)

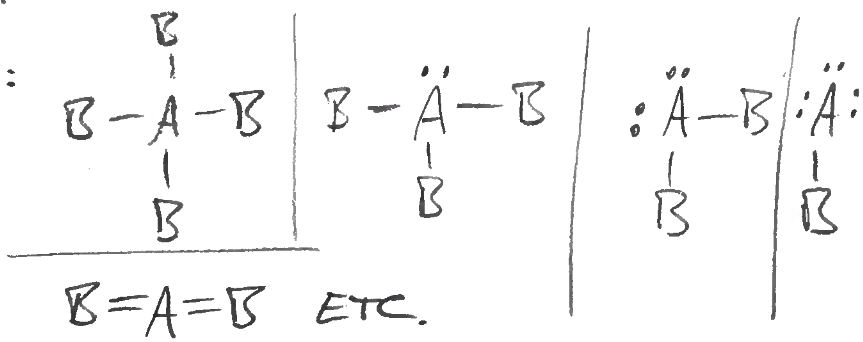
3D 

Rules for Lewis Structures

I. All molecules' drawings must include ALL of the valence electrons incl. any added due to a negative charge or minus any for a positive charge.

II. The Octet Rule: Atoms usually form bonds and/or have lone pairs so that, counting all bonding and non-bonding electrons gives a total of eight.

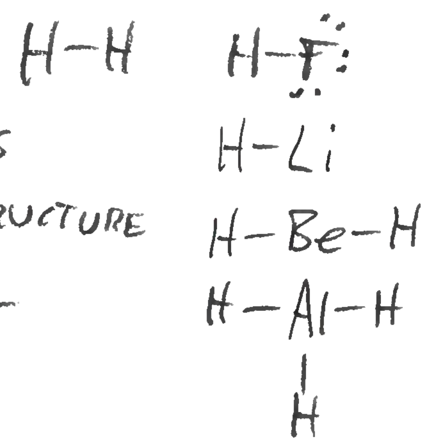
Octets look like this:
(for the "A")



Initial exceptions to rule II.

H only ever makes one bond:

Groups 1, 2, 3 almost never have octets
if we draw a molecular structure



Atoms which can exceed an octet
(if abs. necessary) incl.

P S Cl
As Se Br Kr
Sb Te I Xe

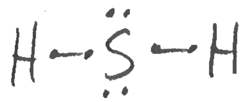
Note: Never B, C, N, O, or F

Always rule-followers

SOME EXAMPLE MOLECULAR STRUCTURES

TO WALK YOU THROUGH THE PROCEDURE ON YOUR HANDOUT

H_2S



$2(1) + 6 = 8e^-$ ✓ S = OCTET ✓

HS^-

$1(1) + 1(6) + 1 = 8e^-$

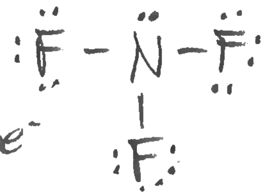
VAL e^- − INSIDE CIRCLE

$6 - 7 = -1$

FORMAL CHARGE

NF_3

$1(5) + 3(7) = 26e^-$



CN^-

$1(4) + 1(5) + 1$

$10e^-$

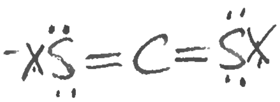
1ST DRAFT

FC FC

$4 - 5 = -1$ $5 - 5 = \varnothing$

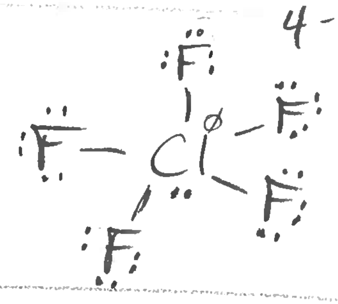
CS_2

$1(4) + 2(6) = 16e^-$



ClF_5

$6(7) = 42e^-$



Cl

FC

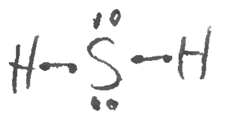
$7 - 7 = \varnothing$

STOPPED HERE GROUP Y Tu 2025-12-16

FOR GROUP X:

H_2S

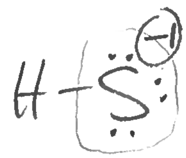
$2(1) + 1(6) = 8e^-$



HS^-

$1(1) + 1(6) + 1$

$8e^-$

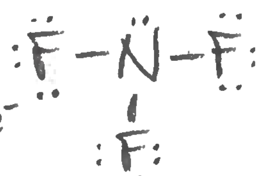


S

$6 - 7 = -1$

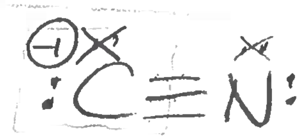
NF_3

$1(5) + 3(7) = 26e^-$



CN^-

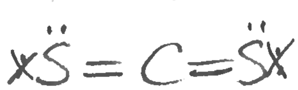
$1(4) + 1(5) + 1 = 10e^-$



C N

$4 - 5 = -1$ $5 - 5 = \varnothing$

CS_2

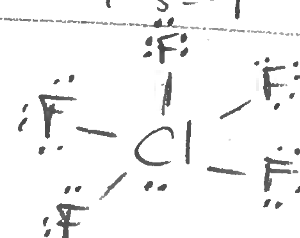


$1(4) + 2(6) = 16e^-$

ClF_5

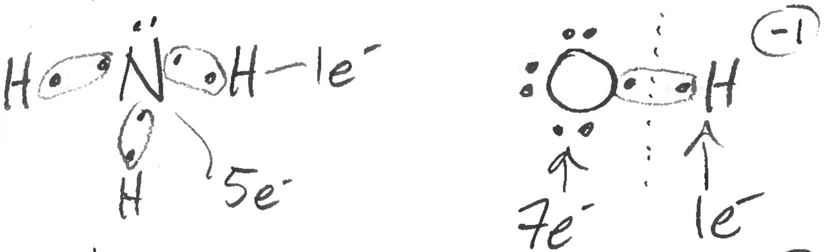
$1(7) + 5(7)$

$42e^-$



FORMAL CHARGE

F.C. KEEPS TRACK OF HOW ELECTRONS ARE SHARED IN STRUCTURAL DRAWINGS — EQUALLY OR UNEQUALLY.

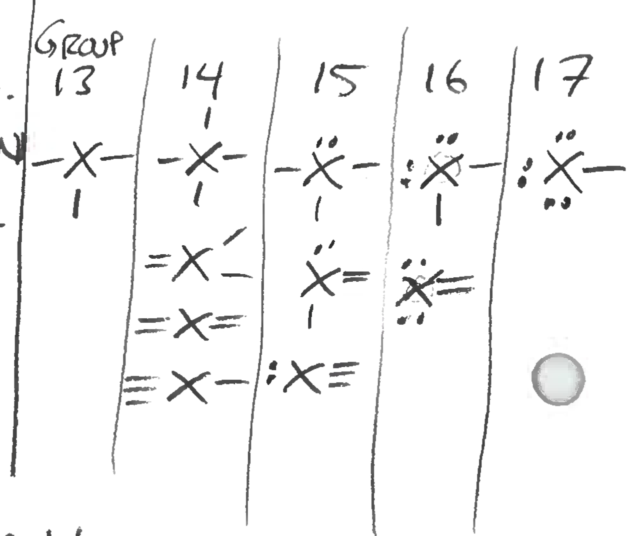


OH^- ION

WHEN SHOULD YOU CALC. F.C.?

1. CALC. F.C. FOR ANY ATOMS DO NOT FOLLOW THEIR USUAL BONDING PATTERN.
2. CALC. F.C. IF IT'S A POLYATOMIC ION TO ACCOUNT FOR THE ION'S CHARGE: ALL F.C. VALUES MUST ADD UP TO THE OVERALL CHARGE (INCL. FOR NEUTRAL ONES).
3. CALC. F.C. IF AN ATOM IN A DRAWING HAS LESS THAN OR MORE THAN AN OCTET. A RULE OF THUMB IS THAT FOR SUCH ATOMS F.C. = ∅ OR +1

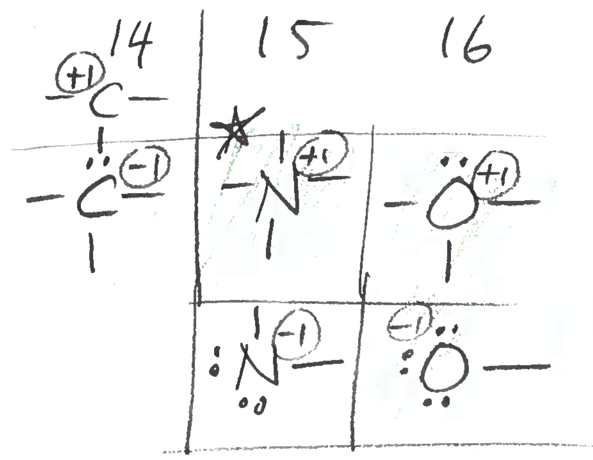
USUAL BONDING



FOR STRUCTURE DRAWING EVALUATION, THE MOST ACCEPTABLE STRUCTURE (VALID) WILL ...

1. HAVE AS FEW ATOMS AS POSSIBLE WITH ANY F.C.
2. HAVE NO F.C. > +1
3. NEG. F.C. VALUES ARE ON MORE ELECTRONEGATIVE ATOMS

COMMON F.C. VALUES



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M 2026-01-05

"Resonance" or Contributing Structures

Section 8.6

Lewis structure drawings have limitations and isolated drawings of molecules may leave out important aspects of the molecule's structure.

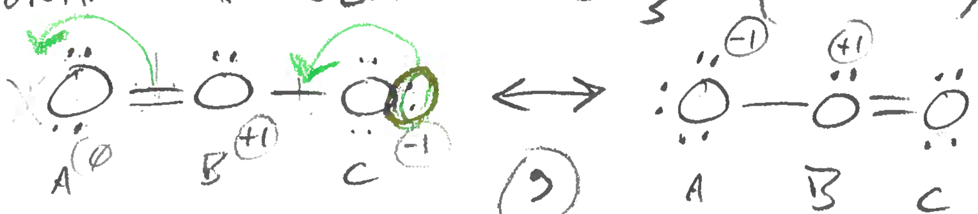
Drawings $\neq$ molecular structures

Contributing structures are multiple drawings which we combine mentally into a single model for understanding bonding in a molecule.

Rules for res.

1. Different contr. structures differ in the placement of lone pairs, double bonds, and/or triple bonds.
2. Atoms remain connected up in the same order in contr. structures for the same molecule. C–O–C $\neq$ C–C–O
3. Although many molecules have contr. structures which appear to be reflections or rotations, they are really rearrangements of electron pairs.

Resonance in ozone: O_3 (18e⁻s)



I. II.

F.C.

I. A 6-6=∅
B 6-5=+1
C 6-7=-1

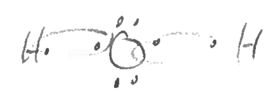
II A -1
B +1
C ∅

THE TWO CONTR. STRUCTURES OF O_3 SHOW US THAT FOR OUR MENTALLY COMBINED MODEL, BONDS ARE EQUIVALENT. NOT SINGLE NOT DOUBLE BUT 1.5 BONDS (3 BONDS, 2 LOCATIONS SHARING e^-s)

IN FACT, ACCORDING TO EXP. MEAS. BOTH BONDS ARE 1.278 Å LONG.

TYPICALLY, O—O 1.48 Å ← 1.278 Å IS RIGHT IN BTWN!
O=O 1.21 Å

THE SHIFTING OF ELECTRON PAIRS IN RESONANCE DRAWINGS SHOULD BE UNDERSTOOD AS A <u>DELOCALIZATION</u> OF ELECTRONS IN THE MOLECULE. FOR OZONE, IT MEANS SOME ELECTRONS ARE SHARED BTWN. ALL THREE ATOMS, NOT JUST TWO NEIGHBORING ATOMS.

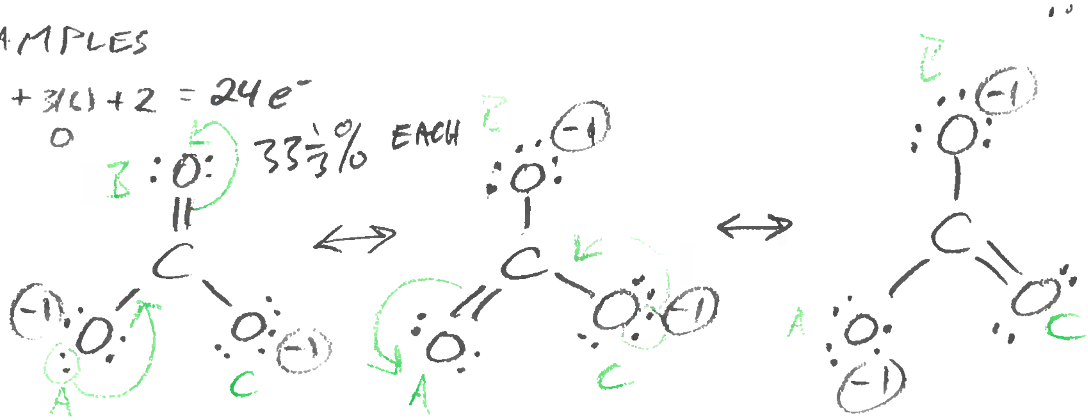


OTHER EXAMPLES

CO_3^{2-} $\underset{C}{4} + \underset{O}{3(6)} + 2 = 24e^-$

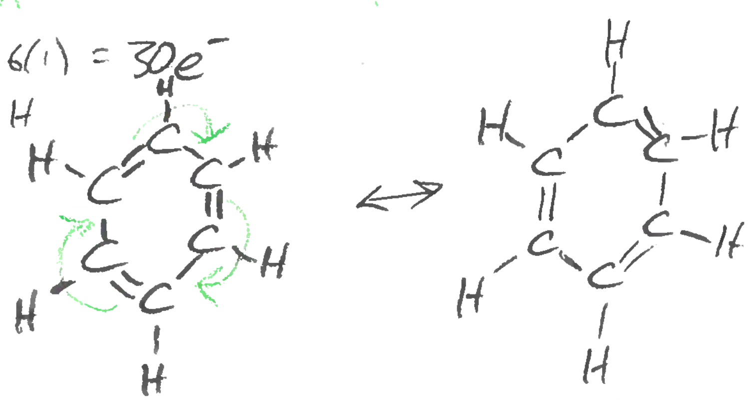
CARBONATE

33⅓% EACH



C_6H_6 $\underset{C}{6(4)} + 6(1) = 30e^-$

BENZENE



50% 50%

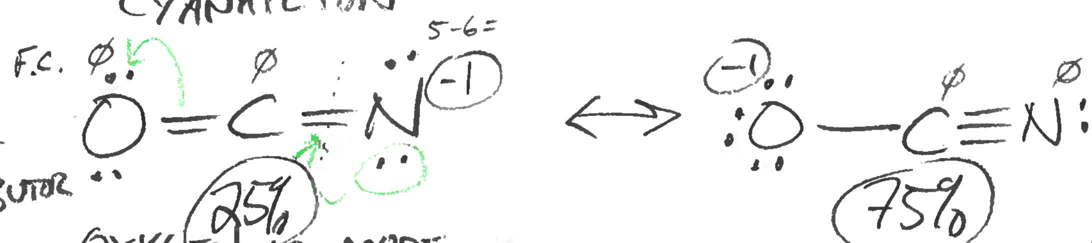
$$F.C. = \left(\begin{matrix}\text{No. of}\\ \text{Valence}\\ e^-\end{matrix}\right) - \left(\begin{matrix}\text{Lone pair } e^-\text{s}\\ + \\ \tfrac{1}{2}\text{ Bonding } e^-\text{s}\end{matrix}\right)$$

WE OFTEN TEACH RESONANCE USING ONLY MOLECULES WITH EQUIVALENT RES. FORMS, THAT LOOK LIKE REFLECTIONS OR ROTATIONS. THINGS GET INTERESTING FOR MOLECULES WITH NON-EQUIVALENT RES FORMS.

FOR EX. OCN^- $\quad \underset{O}{6} + \underset{C}{4} + \underset{N}{5} + 1 = 16e^-$

CYANATE ION

MINOR CONTRIBUTOR



MAJOR CONTRIBUTOR

- OXYGEN IS MORE ELECTRONEGATIVE THAN NITROGEN SO THE STRUCTURE ON THE RIGHT IS PREFERRED.

[TRY THIS ON YOUR OWN: $CH_3OCH_2^{+1}$]

EXCEPTIONS TO THE OCTET RULE

① AN ODD NO. OF e^-s (NO $11e^-$; NO_2 $17e^-$)

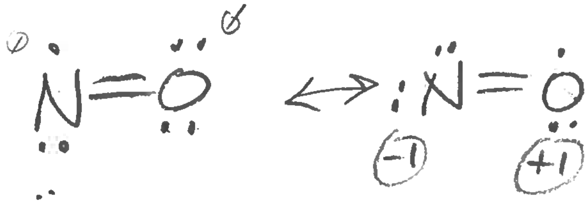
② FEWER THAN AN OCTET. FOR EX., METALS IN LEWIS STRUCTURE DRAWINGS (FROM GROUPS 1, 2, & 13) WILL ALWAYS HAVE LESS THAN $8e^-$s.

③ MORE THAN AN OCTET. SOME ATOMS EXCEED $8e^-$s IN SOME STRUCTURAL DRAWINGS: PERIOD 3 AND HIGHER NON-METALS (P, S, Cl, As, Se, Br, Te, I).
[NEVER C N O OR F]

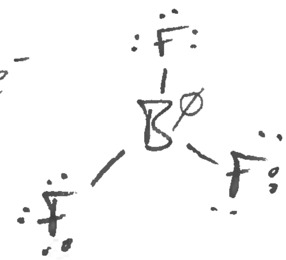
(MY PREF. IS NOT TO DRAW A STRUCTURE WHICH EXCEEDS THE OCTET UNLESS IT CAN'T BE AVOIDED).

FOR EX.

(1) NO $11e^-$

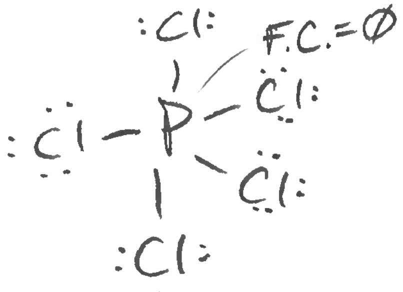


(2) BF_3 $24e^-$

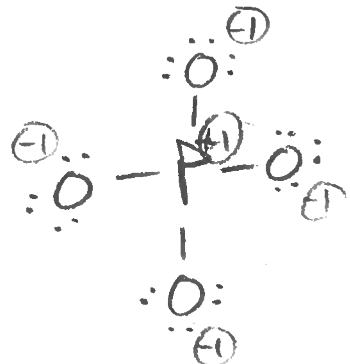


(3) $PCl_5 = 40e^-$ VS PO_4^{3-} $5+4(6)+3=32e^-$

EXCEEDS OCTET



DOES NOT HAVE TO



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8.2 Ionic Bonding

Types of Bonds

(WE ARE INTERESTED IN BONDING B/C THE NATURE OF BONDS IN A SUBSTANCE DETERMINES MELTING POINT, BOILING POINT, CHEMICAL REACTIVITY, AND MANY OTHER THINGS!).

1. COVALENT
2. METALLIC
3. IONIC

Covalent Bonds (BTWN. NON-METALS)

MODELLED AS A SHARING OF AN e^- PAIR BTWN. TWO ATOMS DUE TO OVERLAPPING ATOMIC ORBITALS. THE e^-'S IN BTWN. ATTRACT BOTH NUCLEI. COVALENT BONDS ARE DIRECTIONAL (REQUIRING A SPECIFIC ORIENTATION IN ORDER NOT TO BREAK).

METALLIC BONDS

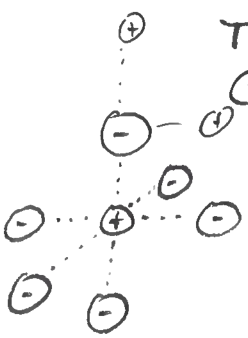
- METAL ATOMS IN A PURE ELEMENT HAVE CORE AND BONDING ELECTRONS. THE SHARED BONDING ELECTRONS ARE IN A CONDUCTING BAND WHICH ALLOWS THEM TO MOVE RANDOMLY BTWN. METAL ATOMS.
- ELECTRONS ARE DELOCALIZED THROUGHOUT THE ENTIRE PIECE OF METAL, WHICH IS WHY METALS ARE GOOD CONDUCTORS OF HEAT AND ELECTRICITY.
- BONDS ARE NOT DIRECTIONAL AND CAN CHANGE ORIENTATION WITHOUT BREAKING. THIS IS WHY METALS ARE MALLEABLE (HAMMER) AND DUCTILE (WIRE).

IONIC BONDS

IONS ATTRACT DUE TO OPPOSITE CHARGE — NO ELECTRONS ARE SHARED.

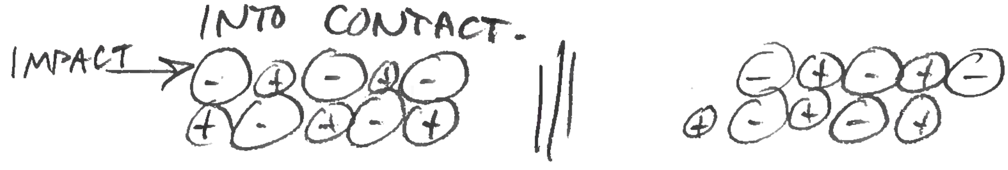
THERE ARE NO DISCRETE (OR SEPARATE) MOLECULES.

BONDING IS UNDERSTOOD TO BE DISTRIBUTED THROUGHOUT THE SOLID. BONDING IS ORGANIZED IN A 3D CRYSTAL LATTICE.



IONIC COMPOUNDS HAVE HIGH MELTING AND BOILING POINTS. THIS IS B/C THE VERY STRONG IONIC BONDS HAVE TO BE BROKEN. BY CONTRAST MOLECULAR COMPOUNDS MAY HAVE LOW MELT/BOIL. POINTS B/C THEIR COVALENT BONDS DO NOT BREAK — ONLY WEAK INTER MOLECULAR BONDS BREAK.

IONIC COMPOUNDS ARE BRITTLE B/C STRONG IMPACTS CAUSE LIKE-CHARGED IONS TO COME INTO CONTACT.



STRONG REPULSIONS SPRINGS THIS APART

IONIC COMPOUNDS ARE NON-CONDUCTORS AS SOLIDS B/C CHARGED PARTICLES ARE FIXED IN PLACE. BUT AS LIQUIDS THEY DO CONDUCT ELECTRICITY B/C THEIR IONS CAN MOVE AROUND.

LATTICE ENERGY

BOND STRENGTH IN IONIC COMPOUNDS CAN'T BE MEASURED PER MOLECULE AS FOR MOLECULAR COMPOUNDS. THIS IS B/C BONDING IS DISTRIBUTED THROUGHOUT THE ENTIRE COMPOUND, NOT JUST WITHIN INDIVIDUAL MOLECULES.

SO WE USE LATTICE ENERGY TO GIVE A MEASURE OF RELATIVE BONDING STRENGTH IN DIFF. IONIC COMPOUNDS.

LATTICE ENERGY IS THE ENERGY RELEASED (EXOTHERMIC) WHEN A SOLID 3D LATTICE FORMS BY A COMBINATION OF GASEOUS IONS. ALTERNATELY, IT'S THE ENERGY ABSORBED (ENDOTHERMIC) WHEN AN IONIC COMPOUND IS VAPORIZED.

WHERE THIS COMES FROM IS A HESS'S LAW CALC. OF THE FORMATION OF AN IONIC COMPOUND FROM ITS ELEMENTS. LATTICE ENERGY IS THE FINAL STEP.

$$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \rightarrow NaCl_{(s)} \quad \Delta H_f^\circ = -410.9\text{ kJ}$$

TO FIND LATTICE ENERGY, WE BREAK THIS DOWN:

SECTION 8.8

1. $Na_{(s)} \rightarrow Na_{(g)}$ $\Delta H = +108\text{ kJ}$ (VAPORIZATION)
2. $\frac{1}{2}Cl_{2(g)} \rightarrow Cl_{(g)}$ $\Delta H = +122\text{ kJ}$ (Cl–Cl BOND STRENGTH)
3. $Na_{(g)} \rightarrow Na^{+1}_{(g)} + 1e^-$ $\Delta H = +496\text{ kJ}$ (1ST IONIZATION ENERGY)
4. $Cl_{(g)} + 1e^- \rightarrow Cl^{-1}_{(g)}$ $\Delta H = -394\text{ kJ}$ (e^- AFFINITY)
5. $Na^{+1}_{(g)} + Cl^{-1}_{(g)} \rightarrow NaCl_{(s)}$ $\boxed{\Delta H = -788\text{ kJ}}$ (LATTICE ENERGY)

$-411\text{ kJ} = (+108) + (+122) + (+496) + (-394) + (-788)$ ✓

WE SOLVE FOR X = LATTICE ENERGY TO CALC. IT.

WHAT DO YOU NEED TO BE ABLE TO DO?

YOU NEED TO QUALITATIVELY COMPARE THE LATTICE ENERGIES OF COMPOUNDS BASED ON THEIR CHEMICAL FORMULAS. WE DO THIS BASED ON COULOMB'S LAW.

$$\boxed{E = k\frac{q_1 q_2}{d}}$$

q = CHARGE ±1, ±2, ±3, +4

d = DISTANCE (SUM OF ~~ATOMIC~~ IONIC RADIUSES)

CHARGE IN COMPARISON BTWN IONS CAN BE 2x, 3x OR EVEN 4x AS BIG. IONIC RADIUS CAN BE A FEW PERCENT BIGGER OR SMALLER. FOR THIS REASON, CHARGE MATTERS MORE.

TRENDS IN LATTICE ENERGY (EXAMPLE COMPARISONS)

(1) THE LARGER THE CHARGES, THE LARGER $E_{LATTICE}$.

$$E_{LATTICE} \quad \underset{+1/-1}{NaCl} < \underset{+2/-1}{CaCl_2} < \underset{+2/-2}{MgO} < \underset{+3/-2}{Al_2O_3}$$

COMPARE ION CHARGES FIRST B/C THEY HAVE A BIGGER EFFECT THAN DIFFERENCES IN ION SIZE.

(2) THE LARGER THE COMBINED RADIUSES ARE, THE SMALLER $E_{LATTICE}$.

ALL COMPOUNDS +1/-1

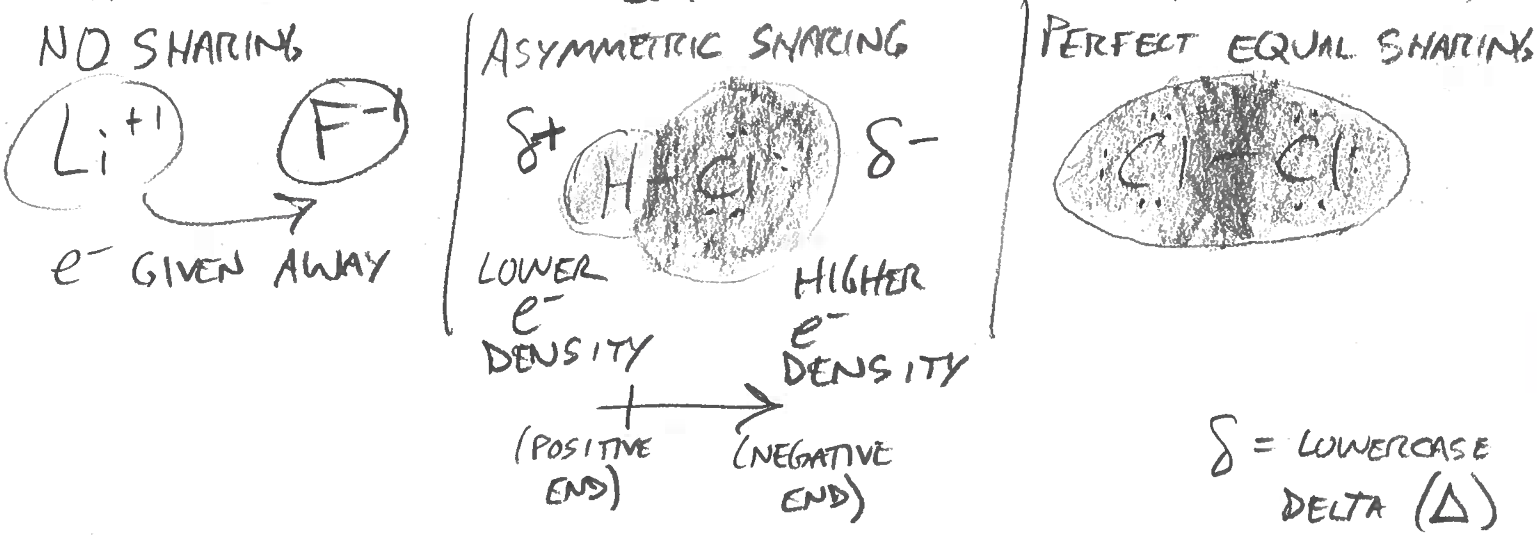
$$KBr < NaCl < LiF$$

THE RELATIVE SIZE OF IONS ONLY MATTERS FOR COMPOUNDS WITH THE SAME CHARGES.

ON A QUIZ OR TEST I WILL ONLY USE UNAMBIGUOUS EXAMPLES.

8.4 Bond Polarity and Electronegativity (EN)

Ionic bonds are pure covalent bonds (like in Cl_2) are at opposite ends of a bond-type spectrum.
A pure ionic bond has <u>no</u> shared electrons Li^{+1} F^{-1}
A pure covalent bond shares electrons equally :Cl–Cl:
Bonds in between these extremes share electrons but unequally, with one atom having a greater share than the other. This comes about b/c diff. atoms have a higher or lower Z_{eff} and whichever atom has a higher Z_{eff} attracts e^-'s more strongly.



δ = lowercase delta (Δ)

Dipoles

δ+ δ− indicate partial charges on specific atoms

$+\!\!\!\longrightarrow$ Arrow starts at + and points toward higher e^- density.

(Partial charges are possible in dipoles, but not in isolation).

(Stopped here M 2026-01-05 Group Y)

What do you need to be able to do?

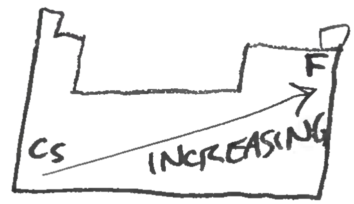
WITH RESPECT TO BOND POLARITY THERE ARE TWO THINGS YOU NEED TO BE ABLE TO DO.

1. DECIDE WHETHER A BOND IS IONIC, POLAR COVALENT, OR NON-POLAR.
2. DETERMINE THE DIRECTION OF BOND POLARITY USING THE PERIODIC TABLE TREND OF ELECTRONEGATIVITY VALUES.

WHEN EN VALUES ARE KNOWN, FIND THE ABSOLUTE VALUE OF THE DIFFERENCE:

ΔEN	
0 – 0.4	NON-POLAR COVALENT
0.4 – 1.9	POLAR COVALENT
>2.0	IONIC

EN TREND



Cs 0.7 $\longrightarrow$ F 4.0

H	2.1
C	2.5
N	3.0
O	3.5

USE THE TREND TO DECIDE WHICH DIRECTION A POLARITY ARROW POINTS:

$\overset{+\!\!\longrightarrow}{H-O}$ OR $\overset{\longleftarrow\!\!+}{O-H}$

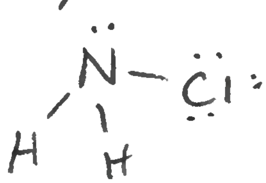
$\overset{+\!\!\longrightarrow}{C-Br}$

$\overset{2.1}{P}-H$

$\Delta EN = \emptyset$!

FOR MOLECULES, CONSIDER EACH BOND SEPARATELY:

NH_2Cl



$\underset{\longleftarrow\!\!+}{N-H}$ $\quad$ $N-Cl$ NON-POLAR

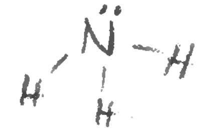
IN CH. 9 WE WILL COMBINE OUR DETERMINATION OF BOND POLARITIES WITH THE 3D SHAPE OF A MOLECULE TO FIND A VECTOR SUM AND THE OVERALL MOLECULAR POLARITY. MOLECULAR POLARITY IS RELEVANT TO UNDERSTANDING INTERMOLECULAR FORCES (CH. 11) AND THE FORMATION OF SOLUTIONS. (CH. 13)

8.8 STRENGTHS OF COVALENT BONDS

A BOND'S STRENGTH IS THE ENERGY REQUIRED TO BREAK IT, MEAS. IN kJ/mol.

RECALL: BREAKING BONDS IS ENDOTHERMIC ($+\Delta H$)
FORMING BONDS IS EXOTHERMIC ($-\Delta H$)

NOTE:
WEAKER BONDS MEANS A MORE REACTIVE MOLECULE B/C BONDS MUST BREAK FOR A REACTION TO OCCUR.

N_2 VS. NH_3, FOR EX. (:N≡N: VS )

GENERAL RULES

① SHORTER BONDS ARE STRONGER BONDS, LONGER ARE WEAKER.

SO B/C BONDING RADIUS OF THE HALOGENS INCREASE AS $F \rightarrow Cl \rightarrow Br \rightarrow I$
BOND STRENGTH FOR $H-X$ DECREASES AS $H-F > H-Cl > H-Br > H-I$

② THE MORE BONDS, THE STRONGER THE BOND.

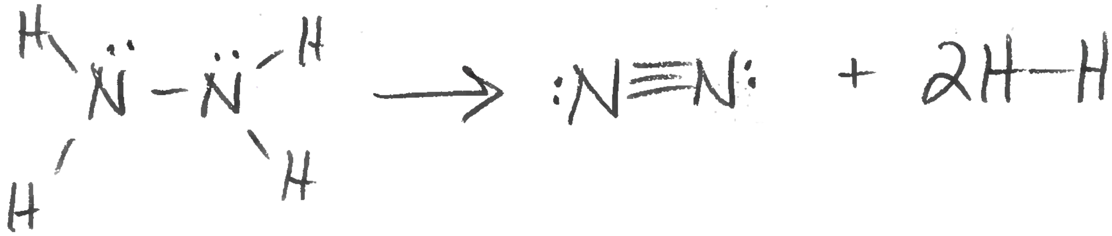
$$\underbrace{(X-X)}_{\text{SINGLE}} < \underbrace{(X=X)}_{\text{DOUBLE}} < \underbrace{(X \equiv X)}_{\text{TRIPLE}}$$

GIVEN A TABLE OF BOND STRENGTHS YOU CAN ESTIMATE THE ΔH_{rxn} BY:

1. FINDING THE SUM OF BONDS BROKEN AND TAKING IT AS $+\Delta H$.
2. FINDING THE SUM OF BONDS FORMED AND TAKING IT AS $-\Delta H$
3. ADDING THESE TOGETHER

EX. $N_2H_4 \rightarrow N_2 + 2H_2$ $\quad \Delta H_{RXN} = ?$

(MUST DRAW LEWIS STRUCTURES, FIRST) [SEE THE MOVIE/BOOK THE MARTIAN]



BONDS BROKEN		BONDS FORMED	
1 N–N	1(163 kJ/mol)	1 N≡N	1(941 kJ/mol)
4 N–H	4(391 kJ/mol)	2 H–H	2(436 kJ/mol)
	+1727 kJ/mol		−1813 kJ/mol

↑ USE B/C EXOTHERMIC

$$\boxed{\Delta H_{RXN} = \Delta H_{BROKEN} + \Delta H_{FORMED}}$$

$$+1727 + (-1813) = \boxed{-86\ kJ/mol}$$

NOTE: IF A BOND IS "BROKEN" AND "FORMED" YOU CAN CANCEL IT OUT DURING CALCULATIONS.

EX. $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

BROKEN		FORMED	
4 C–H	~~4~~ 1(413)	3 C–H	~~3(413)~~
1 Cl–Cl	1(242)	1 C–Cl	1(328)
		1 H–Cl	1(431)
(1894)	+655 kJ/mol		−759 kJ/mol (−1998)

INCL. CROSSED-OUT VALUES

$+655 + (-759) = -104\ kJ/mol$

$(+1894 + (-1998) = -104\ kJ/mol)$