Chemical Bonding

Ionic
Covalent
Polar covalent
Coordinate covalent
Resonance
Valence Bond – Chapter 10
Molecular Orbitals - Chapter 10

Molecular Geometry – Chapter 10

Ionic Bonds

Macromolecular compound
Held together by electrostatic forces
Large ΔEN

Ionic Compound Nomenclature Revisite

Cations – M^{+n} (-n e^{-})

Less EN – element name first

Na^{+1}, Mg^{+2}, Al^{+3}

Anions – N^{-m} (+me^{-})

More EN – “ide” to end of eleme

F^{-1}, O^{-2}, N^{-3}

Ionic Compounds

NaF 3.1	Na_2O 2.5	Na_3N 2.1
MgF_2 2.9	MgO 2.3	Mg_3N_2 1.9
AlF_3 2.6	Al_2O_3 2.0	(AlN) 1.6

Covalent Compounds

(Number prefix)(element) +

(Number prefix)(element)ide

Which element first?

(Less EN element)(more EN +ide)

Use Greek number prefixes

Covalent Chemical Bonding - Approxir

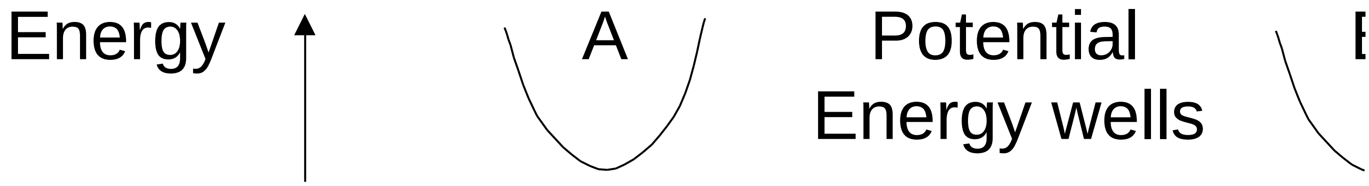
Lewis Dot Structures – shared electrons

Valence Bond Theory (VB) - overlap of atomic orbitals → bonds

Molecular orbital theory (MO) - e^- spread ove of molecule
shared by 2 or more atoms

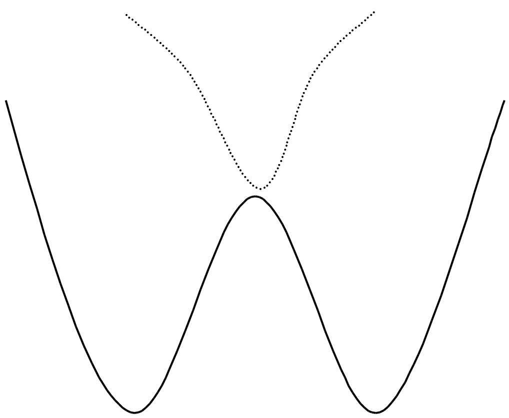
Covalent Bonding

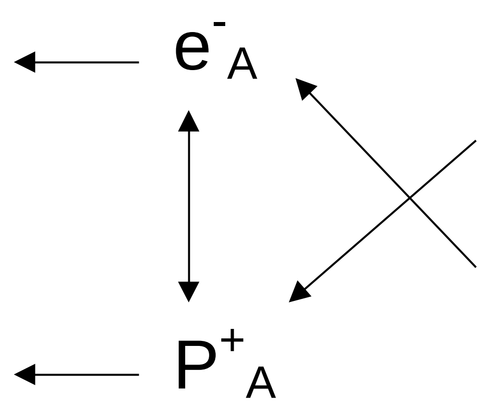
Two hydrogen atoms, A and B, at infinite dis
each other:



Attraction energy $= \dfrac{-\,q_1 q_2}{r} = \dfrac{-\,e^2}{r_{A\ or\ B}}$

Bring together the two atoms A and B. The t
electrons e_A and e_B are each shared by both
The result is a lower energy system.



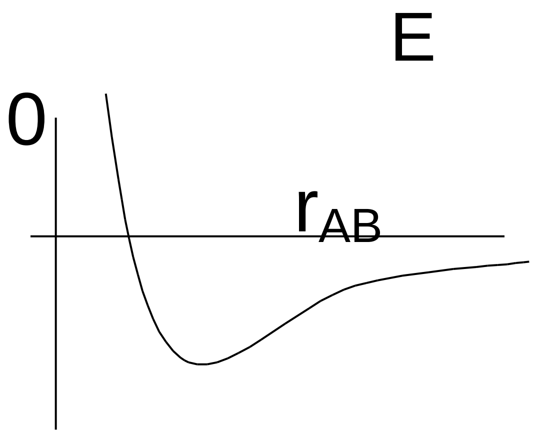


$$\text{Attraction energy} = -\frac{e^2}{r_{P_A e_A}} - \frac{e^2}{r_{P_B e_B}} - \frac{e^2}{r_{P_A e_B}} - \frac{}{r}$$

$$\text{Repulsion energy} = \frac{e^2}{r_{P_A P_B}} + \frac{e^2}{r_{e_A e_B}}$$

Net energy = (E of combined atoms) – (E of separate atoms)

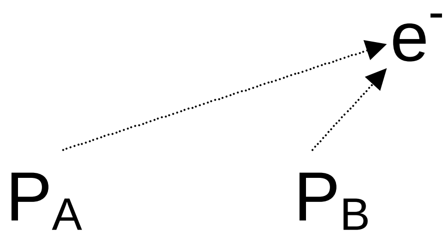
$$= -\frac{e^2}{r_{P_A e_B}} - \frac{e^2}{r_{P_B e_A}} + \frac{e^2}{r_{P_A P_B}} + \frac{e^2}{r_{e_A e_B}}$$



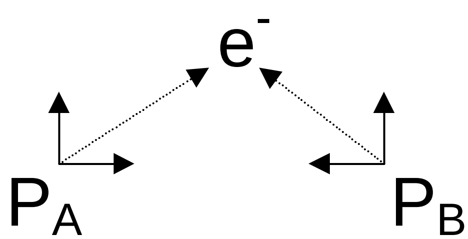
$$r_{P_A P_B} = r_{AB}$$

Bonding → extra electron “density” between molecule, less electron density outside.

In vector terms, for a 2 proton, 1 electron situ
H_2^+):



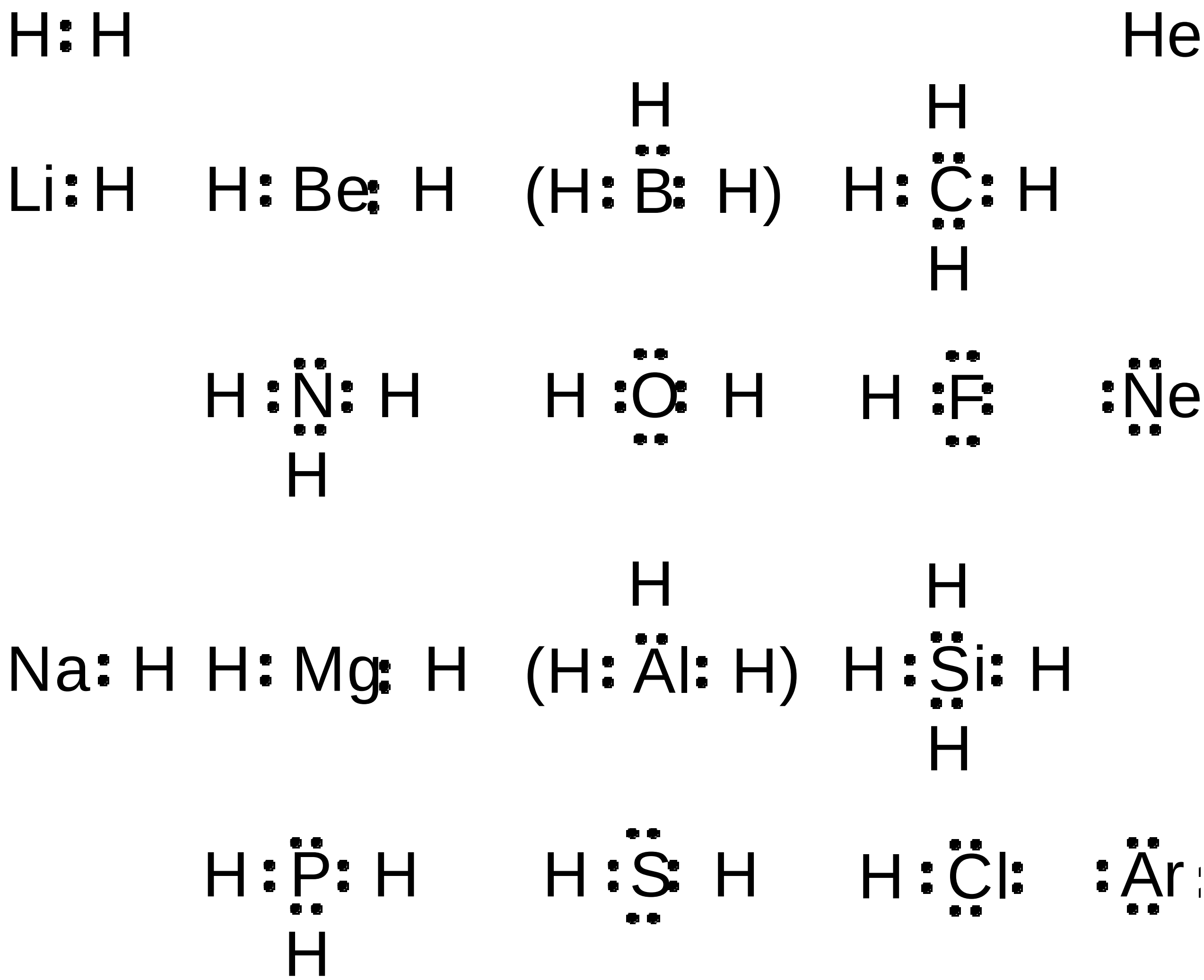
Electron pulls P_A
apart



Electron pulls P_A
together

Electron density between atomic nuclei tends
them together.

covalent – the sharing of electrons between



Period 2 & 3 Hydrides – unstable compound

<u>Covalent Lewis Structures</u>

1. Write skeletal structure (which atoms bon

    H, F, Cl, Br, I usually terminal – only one
    possible
    O, S - single/double
    C, N - single/double/triple
    Central atom <u>usually</u> less EN

2. Count valence electrons

3. Draw single bonds and complete octets

4. If octets not filled, use double or triple bor

Examples:

PF_3

```
    F
F   P   F
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F 3 x 7 = 21e^-

P 1 x 5 = <u>5e^-</u>

26e^-

C_2H_4

```
H   C   C   H
    H   H
```

C 2 x 4 = 8e^-

H 4 x 1 = <u>4e^-</u>

12e^-

CH_4O

C	4 e^-
O	6 e^-
4H	4 e^-
	14 e^-

H

H C H

H

N_2H_4

H N N H

H H

N	2 x 5 =	$10e^-$
H	4 x 1 =	$4e^-$
		$14e^-$

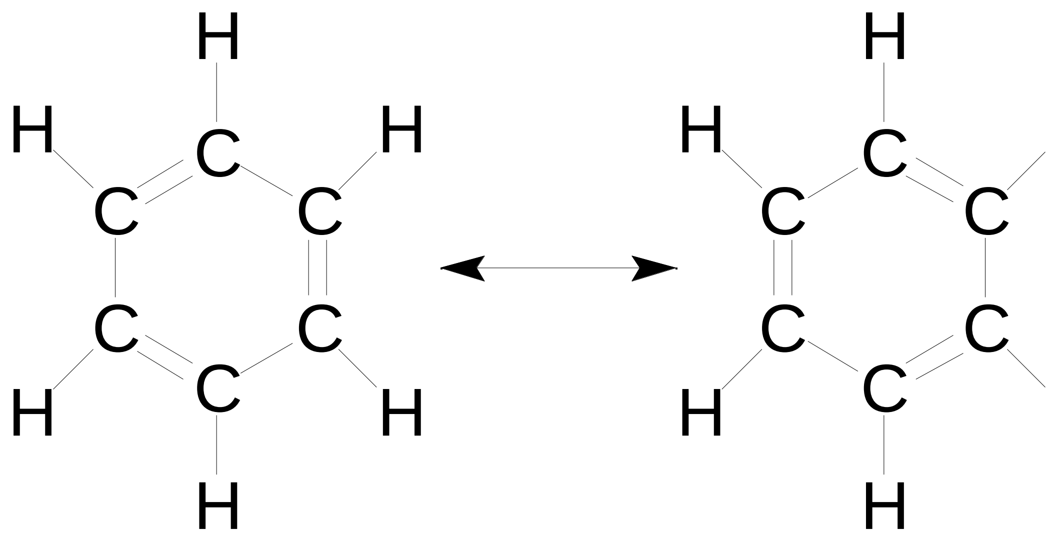
Resonance structures: more than one Lewis possible

a) N_2O_4

$2 \times 5 = 10e^-$
$4 \times 6 = \underline{24e^-}$
$34e^-$

O O
N N
O O

b) C_6H_6 (Benzene)



Bond Lengths

R_{single} (1.54 Å) > R_{double} (1.33 Å) > R_{triple} (1.20

Experimentally, all bonds the same length

Exceptions:

1. Odd electron molecules, e.g., NO, NO_2,
2. S, P compounds SF_6 (12 e^- on S); PF_5 (
3. B, Al previously discussed

Problems:

1. Limited applicability
2. No geometry
3. Resonance structures necessary