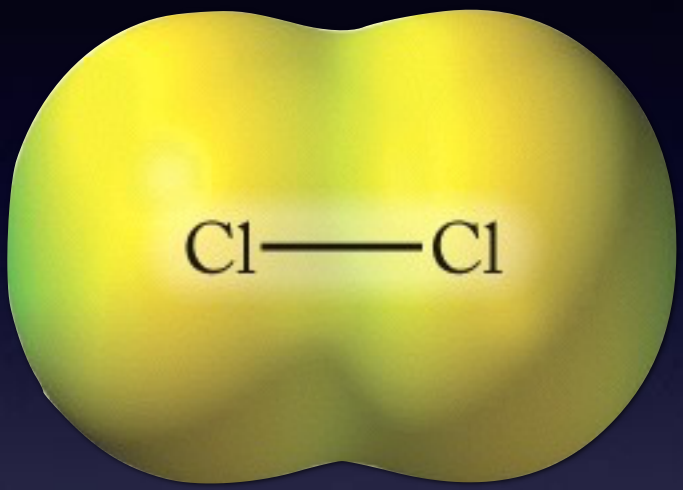
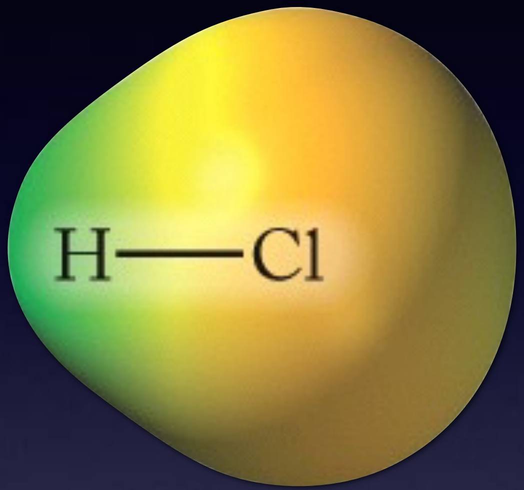
Previously in Molecularity…

The nature of the chemical bond

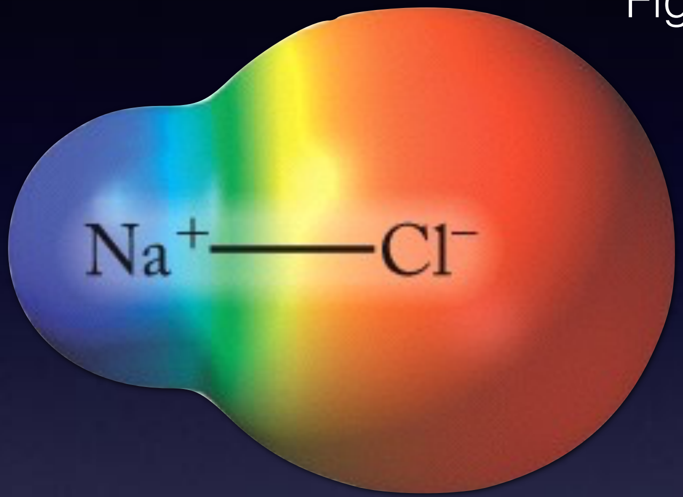
Fig. 4.7
GKF



Cl_2
Nonpolar covalent
$\Delta EN < 0.5$



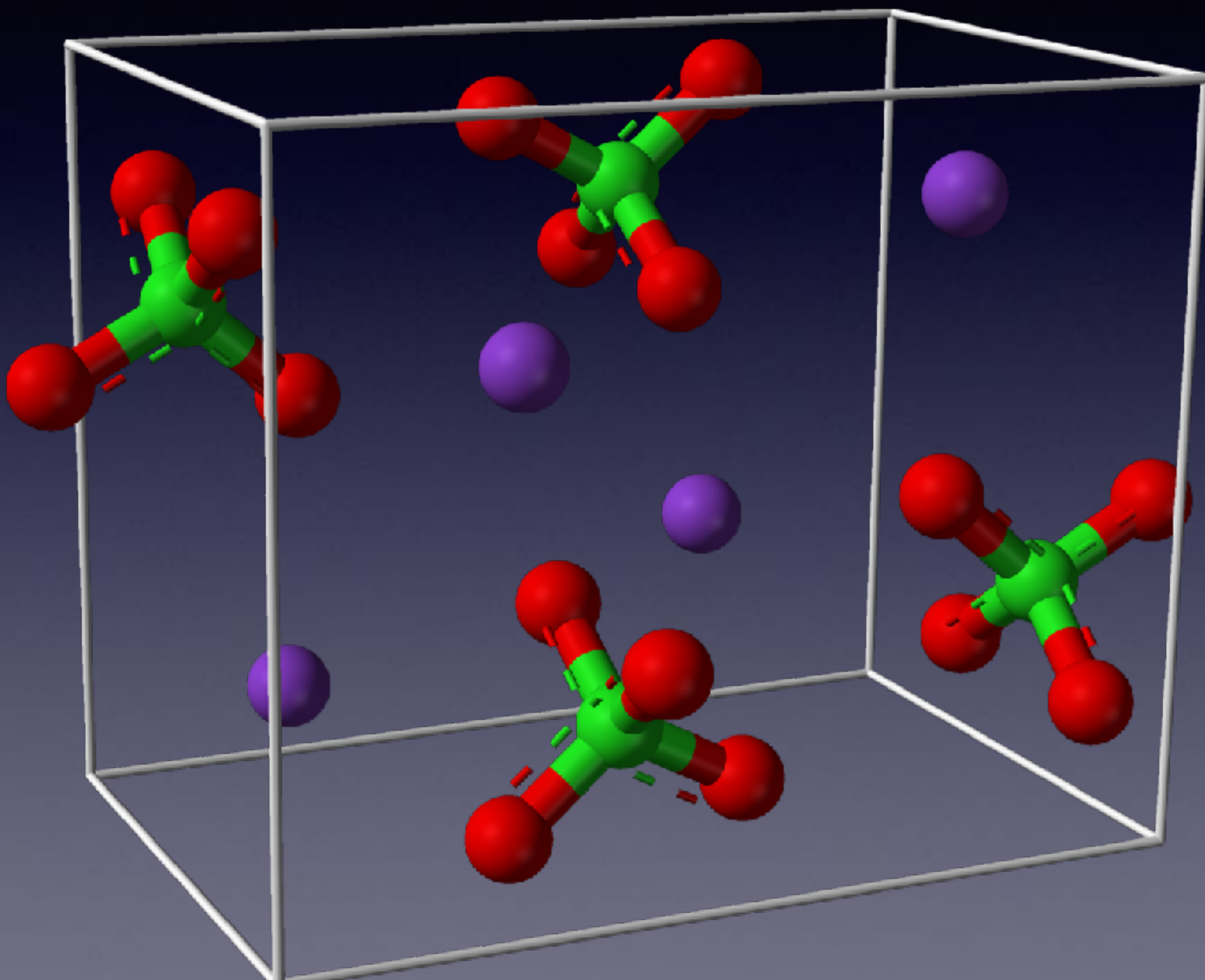
HCl
Polar covalent
$0.5 < \Delta EN < 1.6$



NaCl
Ionic bonds
$\Delta EN \geq 2.0$

Also metallic bonds.

Both ionic and covalent bonds?



- Potassium permanganate, $KMnO_4$
- Permanganate anions, MnO_4^- are covalently bound
- Anions and potassium cations bond ionically

Now you try…

- Classify compounds as forming ionic/covalent/metallic bonds. If covalent, specify polar or nonpolar. If multiple, explain.

Nitric acid (consider bonding between H and nitrate)

Sodium nitrate

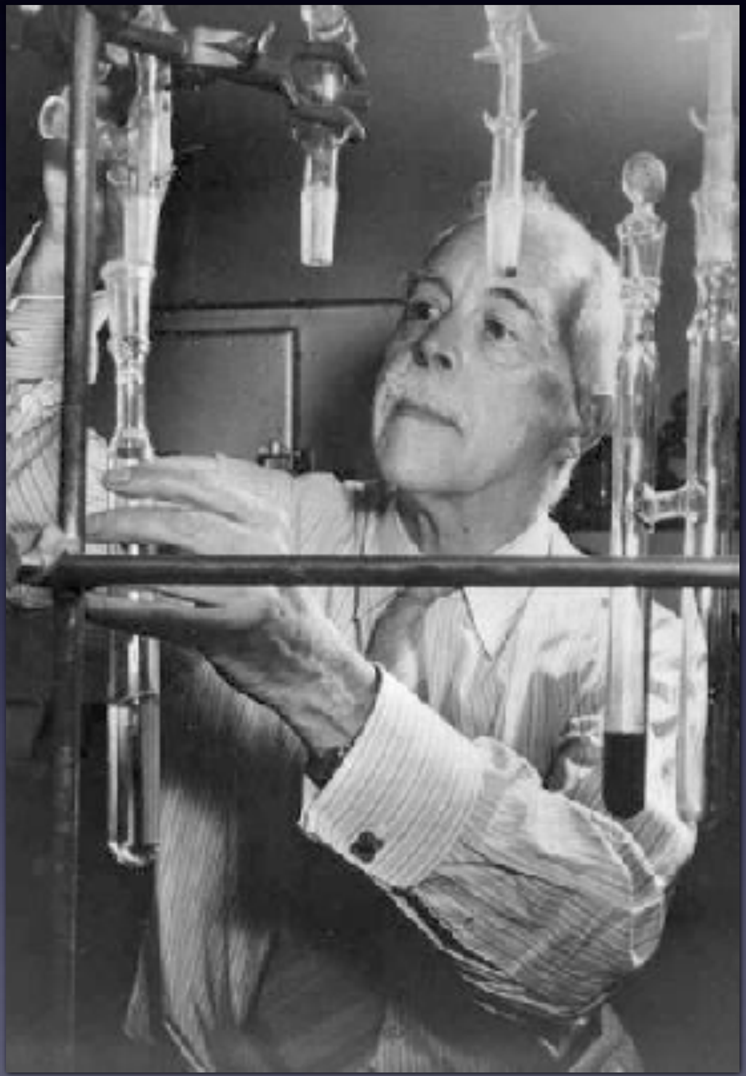
Cesium bromide

Hydrobromic acid

Ammonia

Understanding / predicting covalent bonding

- Gilbert Newton Lewis developed a dot notation to understand and predict covalent bonding
- We will focus on **molecular** compounds in class, but …
- …you are also responsible for understanding Lewis dot structures of **ionic** compounds (GKF §4.3)



G. N. Lewis

www.atomicheritage.org
Gilbert%20Lewis.jpg

Where are we going today?

Ch1010-A17-A03 Lecture 12

- §4.4–5, 4.7, 5.3 Gilbert gets the measles

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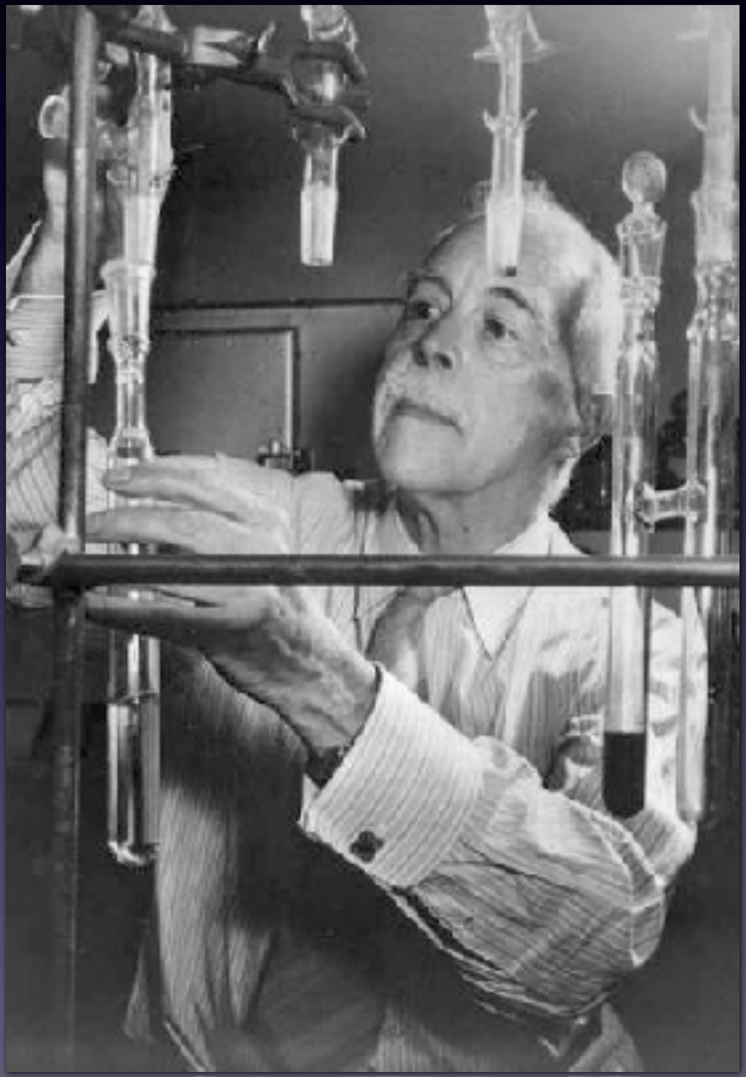
The Octet "Rule"

- Atoms form chemical bonds by sharing electrons
- Atoms tend to gain, loose, or share electrons to have 8 valence electrons
- Theory predates quantum mechanics
- <u>Generally</u> accurate for s- and p-block.
- Totally fails for d-block (see 18 e^- rule) and totally fails for f-block (crazytown)



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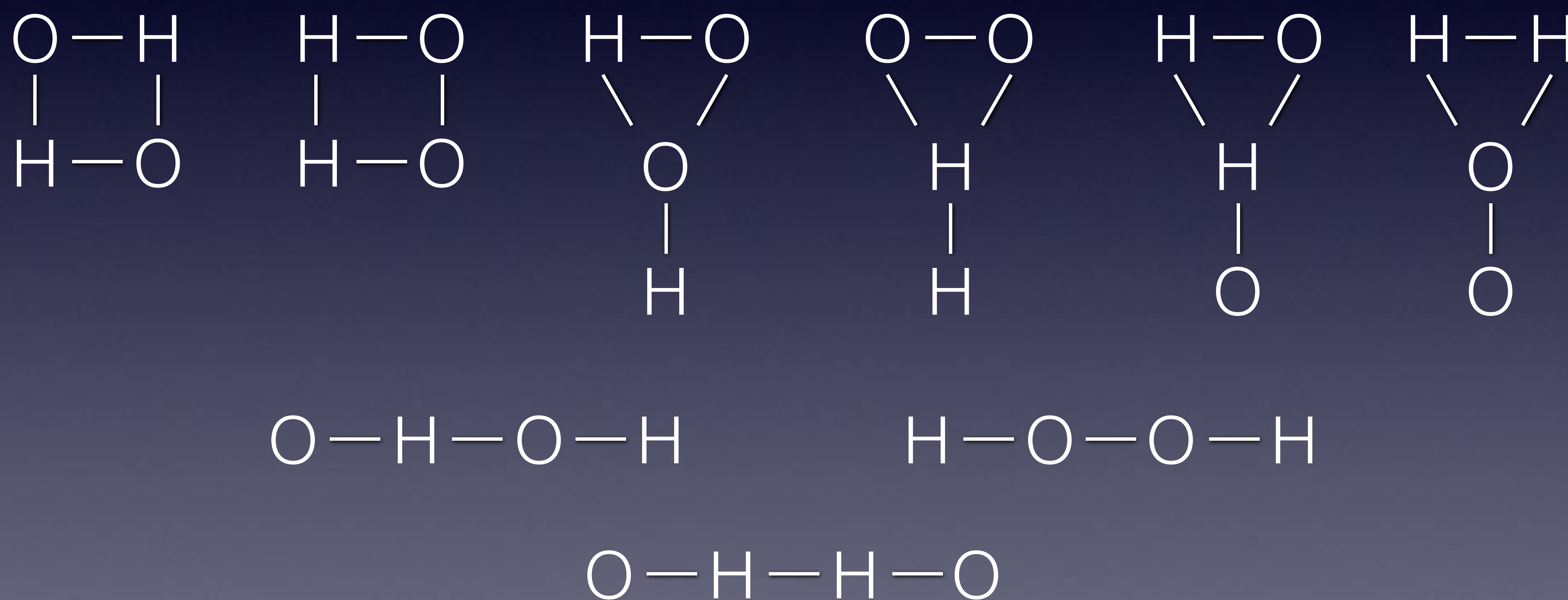
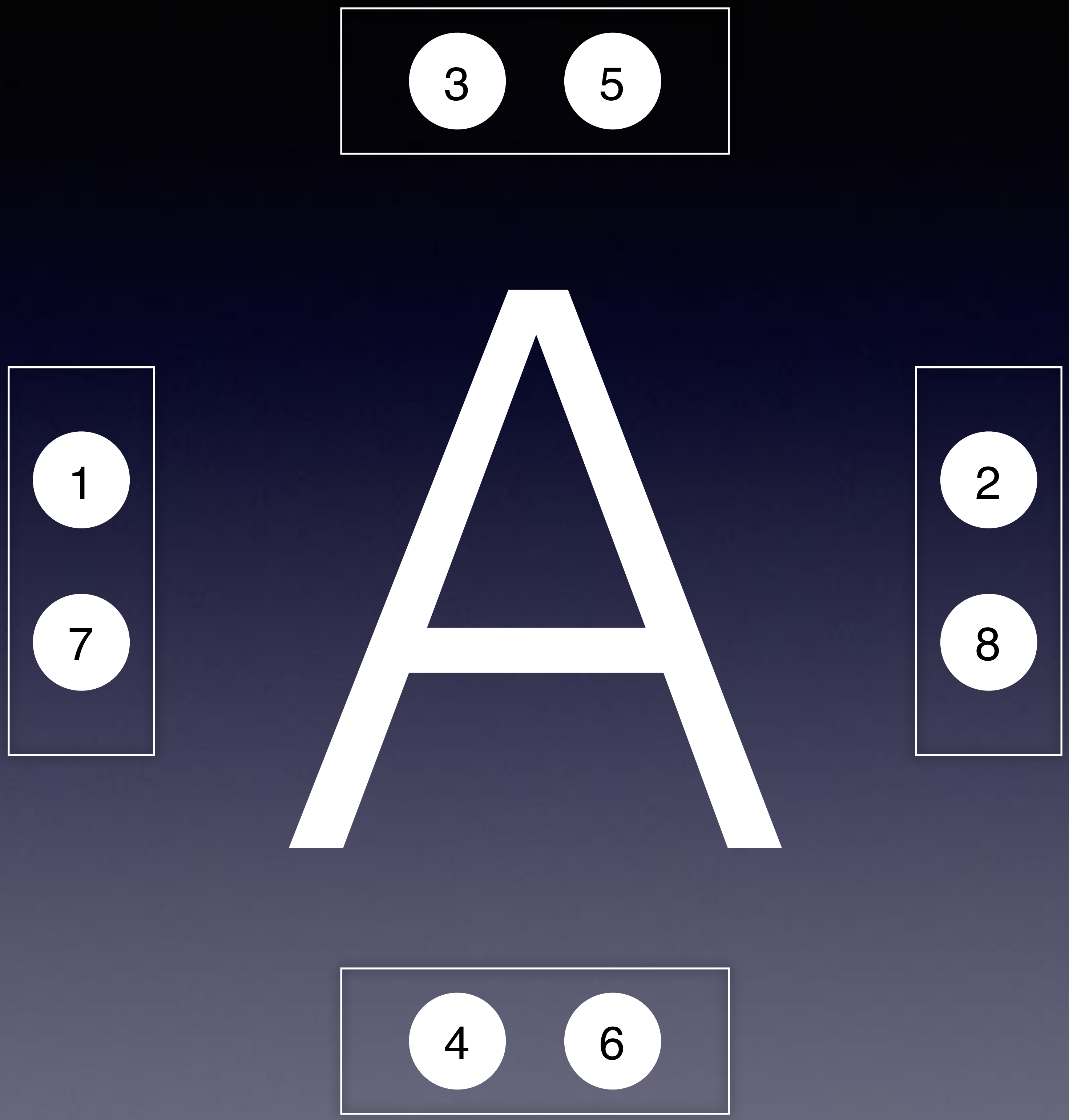
Lewis structures are predictive

- Consider hydrogen peroxide (H_2O_2):

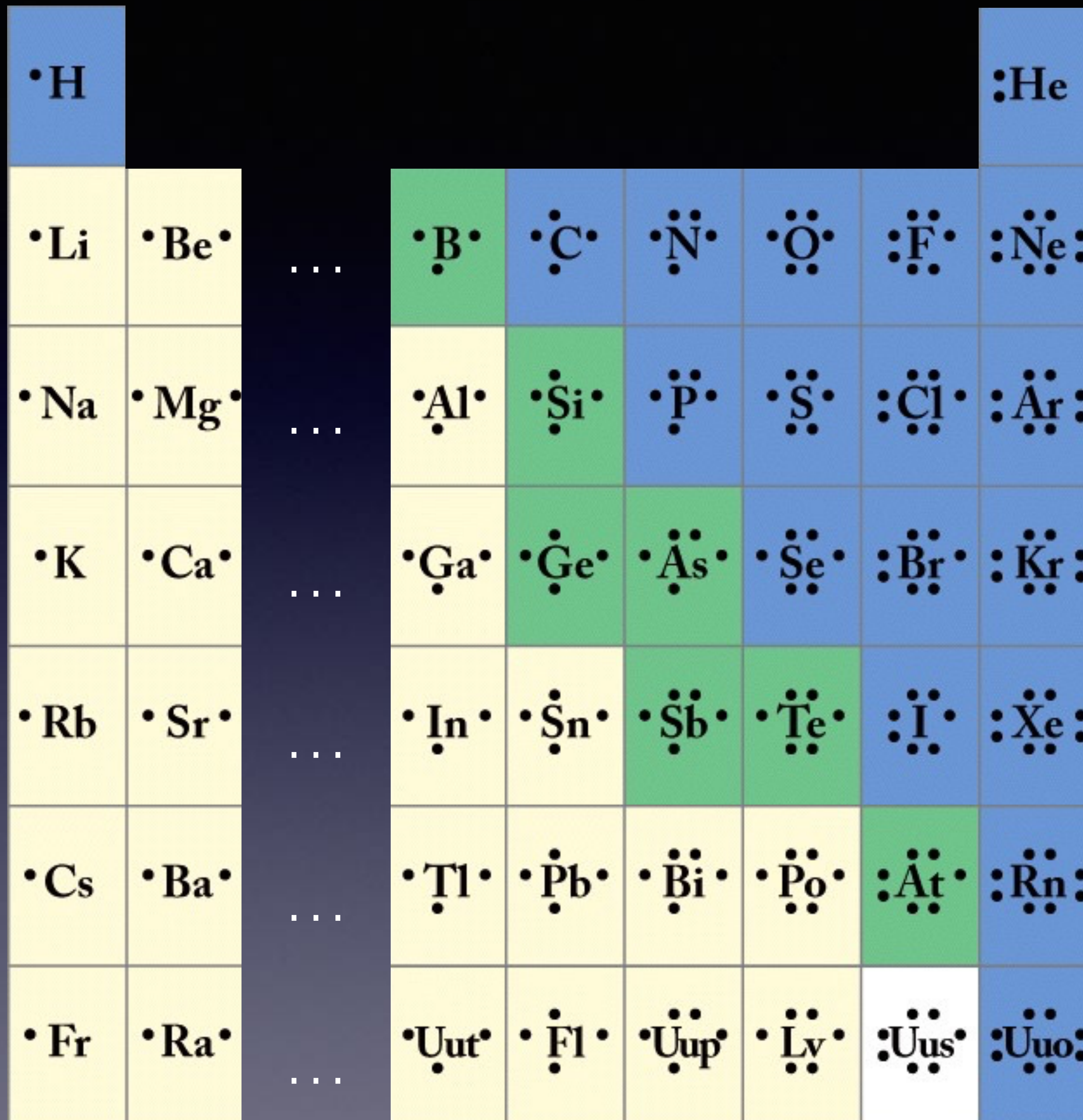


- Useful for resonance and formal charge (down the road).

Octet Rules



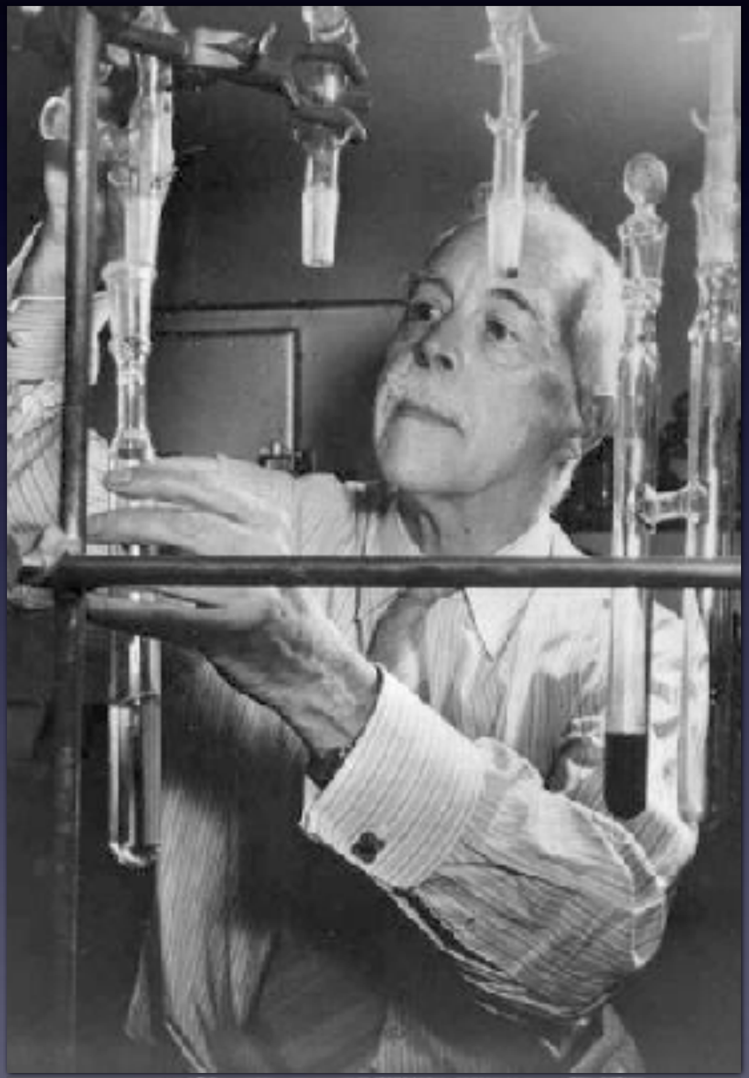
- Each of 4 sides gets up to 2 electrons
- Fill across for 1 to 2; Don't pair until #5
- Pairs are called "lone pairs"
- Shared electrons are "bonds"
- Pretend H and He follow a *duet* rule


H
He
Li Be … B C N O F Ne
Na Mg … Al Si P S Cl Ar
K Ca … Ga Ge As Se Br Kr
Rb Sr … In Sn Sb Te I Xe
Cs Ba … Tl Pb Bi Po At Rn
Fr Ra … Uut Fl Uup Lv Uus Uuo

GKF Fig 4.5

How to win at measles

- Sum **valence** e^- including overall charge.
  (This determines total # of bonds and lone pairs)
- Arrange around a central atom…
  1. Greatest bonding capacity
  2. Lowest electronegativity
- Draw single bonds to central atom
- Complete octets around periphery
- Fix central atom octet by converting peripheral lone pairs to bonds as needed
- Determine resonance / formal charge



Gilbert N. Lewis

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Gilbert%20Lewis.jpg

Example: phosphene, PH_3

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Examples in class

need more examples

- Some easy ones:
  - Difluorodichloromethane (which is a CFC by the way) CCl_2F_2
  - Phosphine (PH_3)
- Organic molecules including ethene, C_2H_4, formaldehyde CH_2O, formic acid HCOOH (both oxygens connect to carbon), and ethanol C_2H_5OH ($H_3C–CH_2OH$)
- More challenging ones involving ions… we'll only get partway…
  - Perchlorate anion $[ClO_4]^–$
  - Chlorate anion $[ClO_3]^–$
  - Ammonium cation $[NH_4]^+$

Guiding principle…
Least electronegative atom at center!

Notable s- and p-block exceptions

- Hydrogen $1s^1$ and helium $1s^2$ follow a *duet* rule
  (n = 1 can only hold two electrons)

- Alkali earths and the boron family sometimes do their own thing
  Example: magnesium hydride (MgH_2)
  Example: boron trifluoride (BF_3)

- Molecules with an odd total of valence electrons
  (one atom will end up with an odd count, duh!)
  Example: nitrogen monoxide

- *Hypervalent* molecules or *expanded octets*:
  large atoms in the 3p, 4p, … block

Group 2 and Group 3 exceptions

- Alkali earths and the boron family sometimes do their own thing
  Example: magnesium hydride (MgH_2)
  Example: boron trifluoride (BF_3)

- We'll understand more exceptions as we learn about resonance and formal charge.

Where did we go today?

Ch1010-A17-A03 Lecture 12

- §4.4–5, 4.7, 5.3 Gilbert gets the measles

Next time…

- §5.4 Resonance and formal charge

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