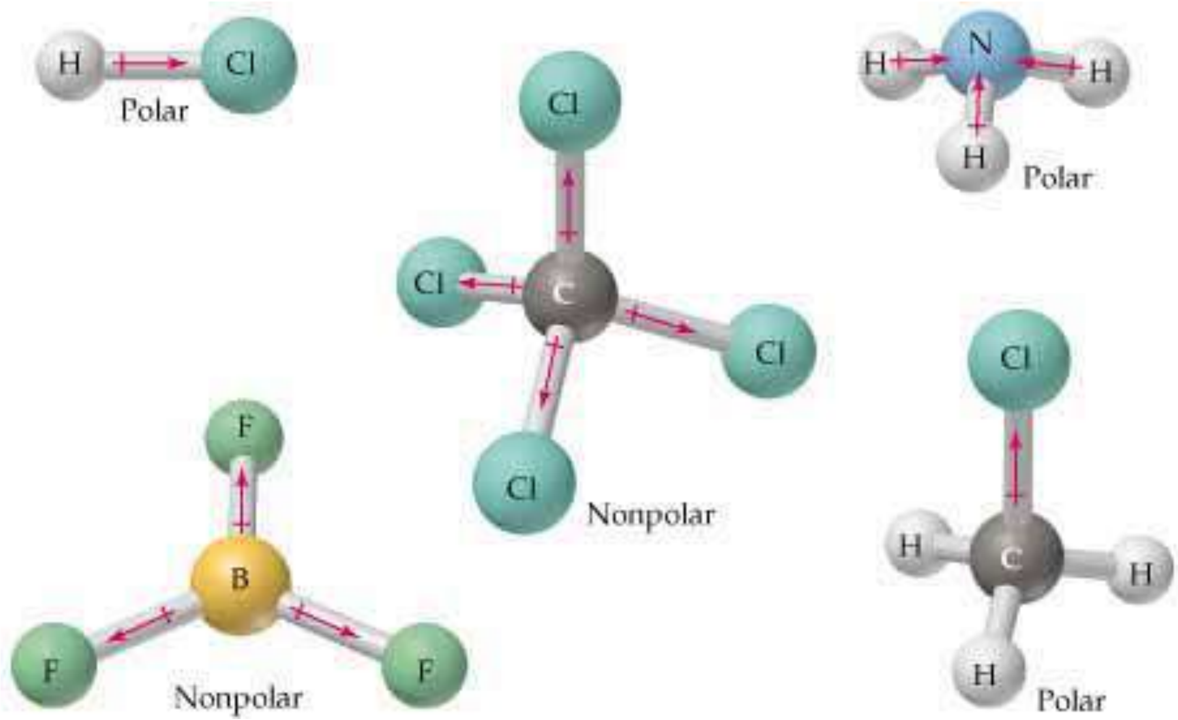
NOTES: 8.4 – Polar Bonds and Molecules



ELECTRONEGATIVITY:

- We’ve learned how valence electrons are shared to form covalent bonds between elements.
- So far, we have considered the electrons to be shared equally.
- However, in most cases, **electrons are NOT shared equally** because of a property called **electronegativity**.

<u>ELECTRONEGATIVITY:</u>

- **<u>ELECTRONEGATIVITY</u>** = the tendency for an atom to attract electrons to itself when it is chemically combined with another element.
- ***The result: <u>a "tug-of-war" between the nuclei of the atoms.</u>***

<u>ELECTRONEGATIVITY:</u>

- Electronegativities are given numerical values (the most electronegative element has the highest value; the least electronegative element has the lowest value)

****See table 6.2 p. 181**

- Most electronegative element: <u>Fluorine (4.0)</u>
- Least electronegative elements: <u>Fr (0.7), Cs (0.7)</u>

Electronegativity

- 0.5 - 0.9
- 1.0 - 1.4
- 1.5 - 1.9
- 2.0 - 2.4
- 2.5 - 2.9
- 3.0 - 3.5
- 3.6 - 3.9
- 4.0 +

1	2	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)
H 2.1																	He --
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.3	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.6	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr --
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9	Ac 1.1	Rf --	Db --	Sg --	Bh --	Hs --	Mt --	Uun --	Uuu --	Uub --		Uuq				

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

<u>ELECTRONEGATIVITY:</u>

- <u>Notice the periodic trend:</u>

  ➔ As we move from **left to right** across a row, electronegativity **<u>increases</u>** (metals have low values nonmetals have high values – excluding noble gases)

  ➔ As we move down a column, electronegativity **<u>decreases</u>**.

****The higher the electronegativity value, the greater the ability to attract electrons to itself.***

Electronegativity

- 0.5 - 0.9
- 1.0 - 1.4
- 1.5 - 1.9
- 2.0 - 2.4
- 2.5 - 2.9
- 3.0 - 3.5
- 3.6 - 3.9
- 4.0 +

1	2	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)
H 2.1																	He --
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.3	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.7	Ga 1.6	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr --
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.7	Xe 2.6
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 2.0	Pb 2.3	Bi 2.0	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9	Ac 1.1	Rf --	Db --	Sg --	Bh --	Hs --	Mt --	Uun --	Uuu --	Uub --		Uuq				

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

How are different bond types determined?

- Differences in electronegativity between atoms in a compound are used to determine bond type.
- Three types:

1) Nonpolar covalent bond

2) Polar covalent bond

3) Ionic bond

NONPOLAR BONDS:

- When the atoms in a molecule are the same (or have the same, or very close electronegativities), the bonding electrons are **shared equally**.

- **Result**: **a nonpolar covalent bond**

  ➔ Examples: **O_2, F_2, H_2, N_2, Cl_2**

<u>POLAR BONDS:</u>

- When 2 **different** atoms are joined by a covalent bond, and the **bonding electrons are shared unequally**, the bond is a polar covalent bond, or **<u>POLAR BOND</u>**.

- **The atom with the stronger electron attraction (the more electronegative element) acquires a <u>slightly negative charge</u>.**

- **The less electronegative atom acquires a <u>slightly positive charge</u>.**

<u>POLAR BONDS:</u>

- **Example: <u>HCl</u>**

$**\delta$ = "partial"

- Electronegativities:

  H = <u>2.1</u>

  Cl = <u>3.0</u>

$$\delta^{+}\,H - Cl^{\,\delta^{-}}$$

POLAR BONDS:

- **Example: $\underline{H_2O}$**

- Electronegativities:

  H = 2.1

  O = 3.5

POLAR BONDS:

- **Example: H_2O**

- Electronegativities:

  H = 2.1

  O = 3.5

POLAR BONDS:

- **Example: $\underline{H_2O}$**

- Electronegativities:

  H = 2.1

  O = 3.5

Predicting Bond Types:

- Electronegativities help us predict the type of bond:

Electronegativity Difference	Type of Bond	Example
0.0 – 0.4	**nonpolar covalent**	H-H (0.0)
0.4 – 1.0	**moderately polar covalent**	HCl (0.9)
1.0 – 2.0	**very polar covalent**	HF (1.9)
$\geq$ 2.0	**ionic**	Na^+Cl^- (2.1)

POLAR MOLECULES:

- A polar bond in a molecule can make the entire molecule polar
- A molecule that has 2 poles (**charged regions**), like H-Cl, is called a dipolar molecule, or **dipole**.

DIPOLE:

- A molecule that has **two electrically charged regions**, or poles
- **EX:** any **molecule** with a polar covalent bond
- **Important Note:**
  - ➔ **Symmetry can cancel out dipoles**
  - ➔ EX: **CF_4**

POLAR MOLECULES:

- Example: **CO_2**

δ- δ-

O = C = O   shape: **linear**

*The bond polarities cancel because they are in opposite directions; CO_2 is a **nonpolar molecule**.

POLAR MOLECULES:

- Water, H_2O, also has 2 polar bonds:

➔ But, the molecule is bent, so the bonds do not cancel.

➔ H_2O is a polar molecule.

δ-

δ+

δ+

How are different molecules held together in groups?

- **Intermolecular forces** (forces of attraction between molecules) <u>hold groups of molecules together</u>
- weaker than either an ionic or a covalent bond
- determine, among other things, whether a compound is a <u>gas, liquid, or solid</u> at a particular temp.

Intermolecular Forces:

Include:

1) van der Waals forces:

-Dispersion forces

-Dipole interactions

2) Hydrogen bonds

van der Waals Forces:

- general term used to describe the weakest intermolecular attractions including: **dispersion forces** and **dipole interactions**

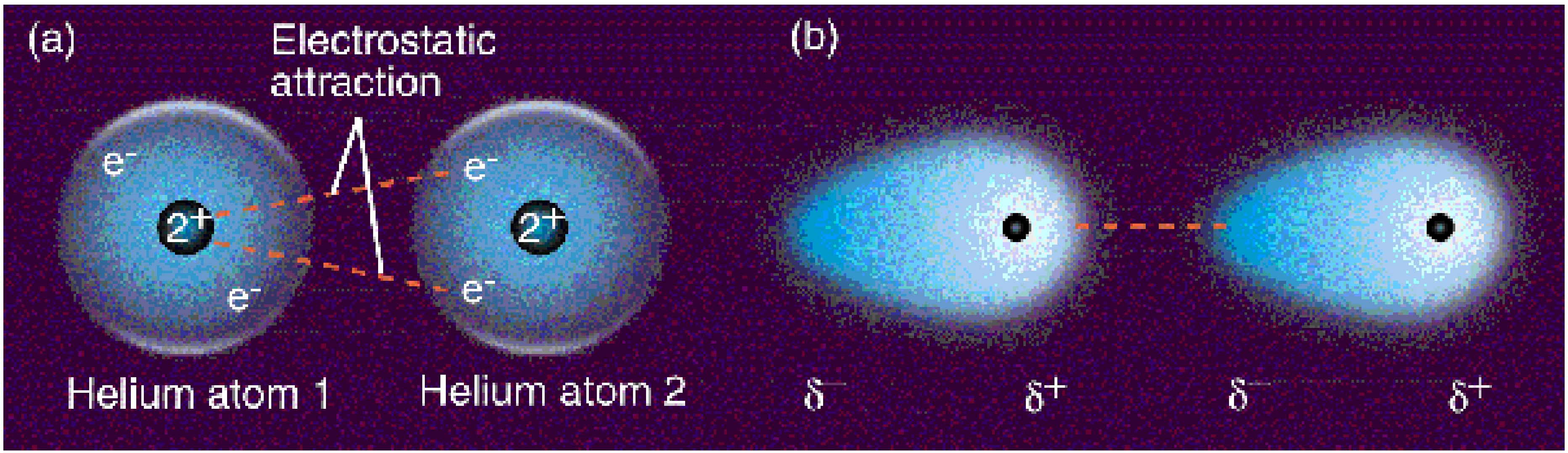
Dispersion Forces:

- **Dispersion Forces** = weakest type of intermolecular attraction, caused by the movement of electrons

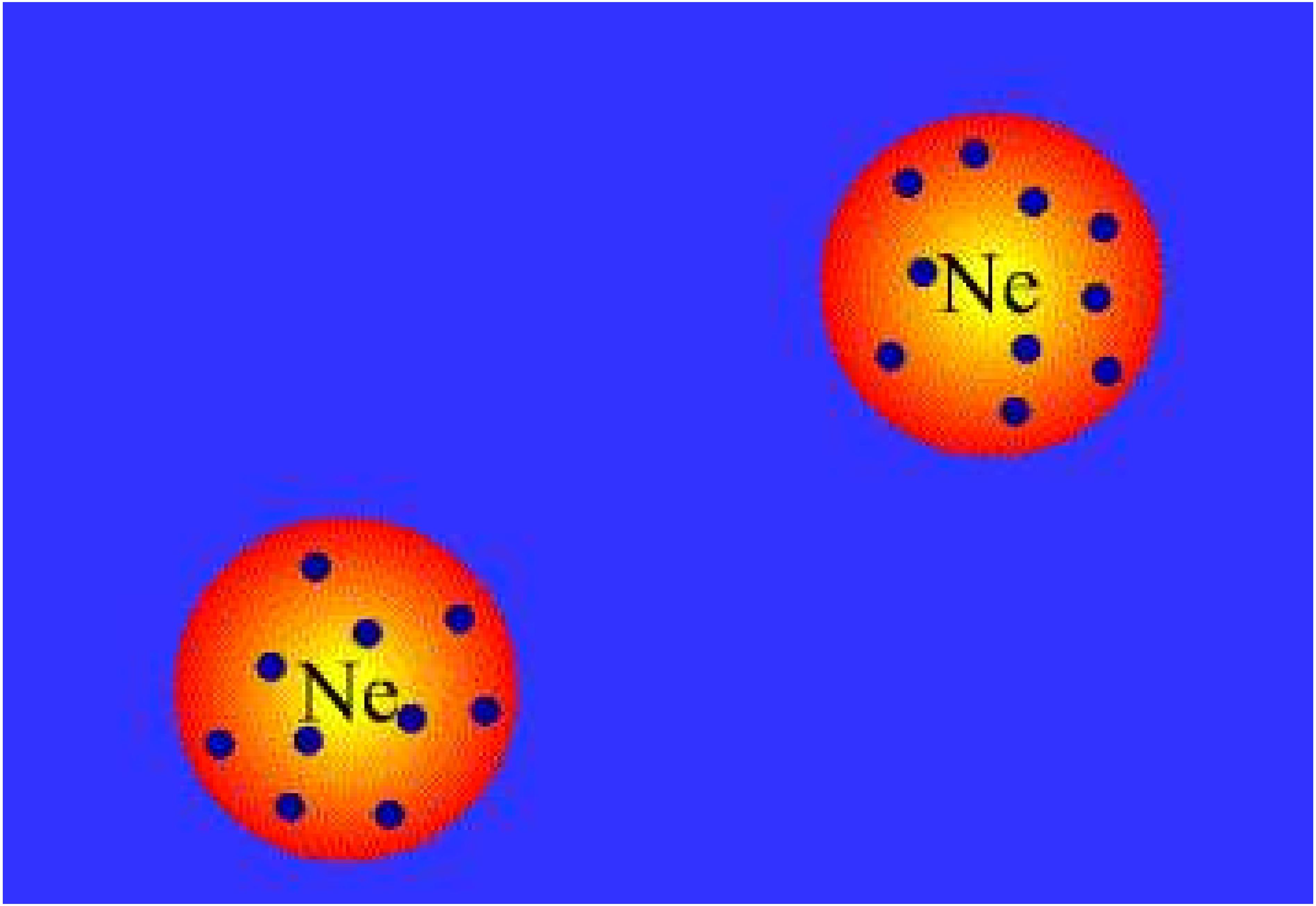
  -more electrons leads to stronger force of attraction between molecules

  EX: $F_2 < Cl_2 < Br_2 < I_2$

  (F_2 and Cl_2 are **gases** at room temp; Br_2 is a **liquid**; I_2 is a **solid**)



(a)
Electrostatic attraction
e⁻
2+
e⁻
e⁻
2+
e⁻
Helium atom 1
Helium atom 2
(b)
δ−
δ+
δ−
δ+

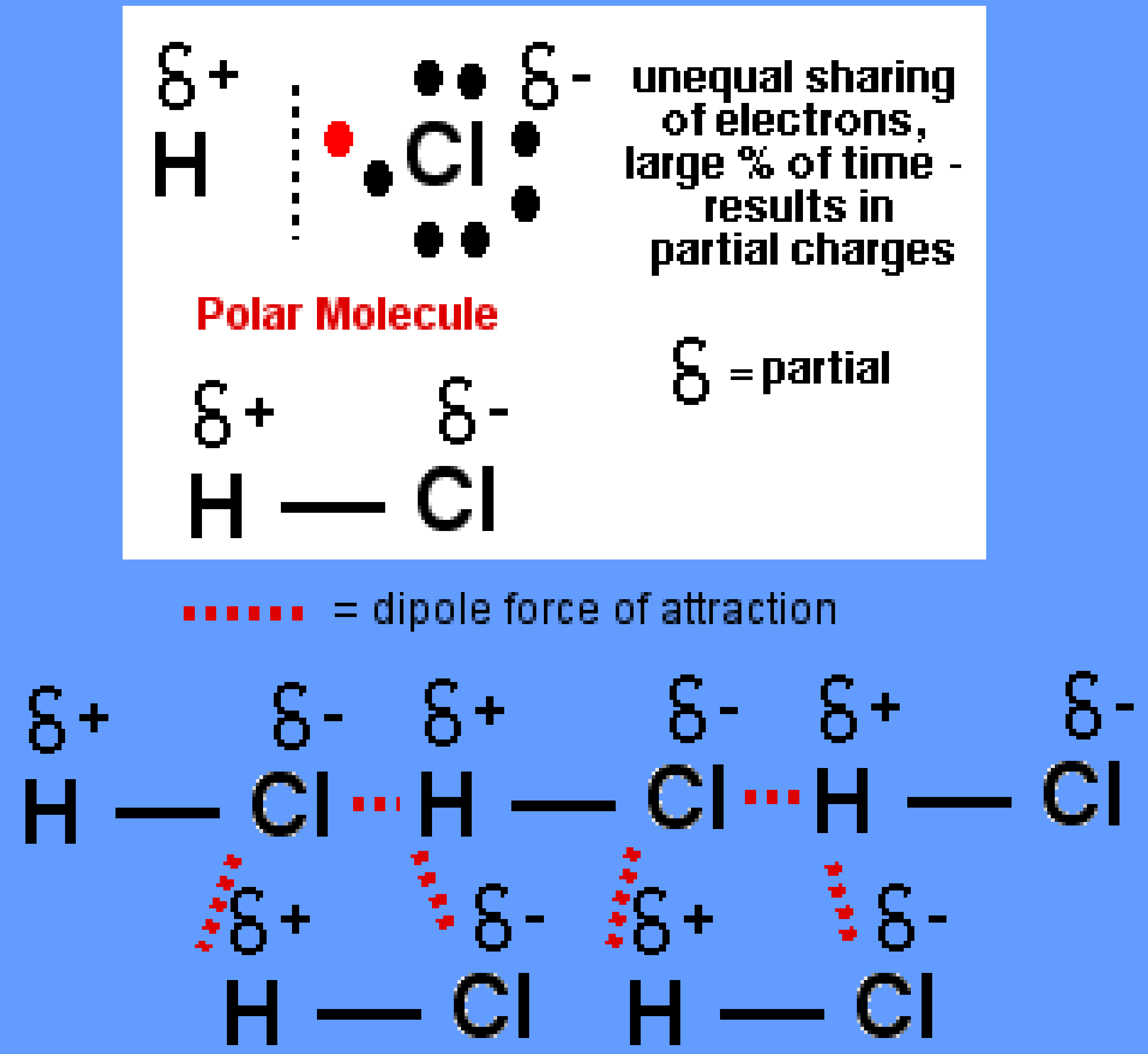

Ne
Ne

Dipole Interactions:

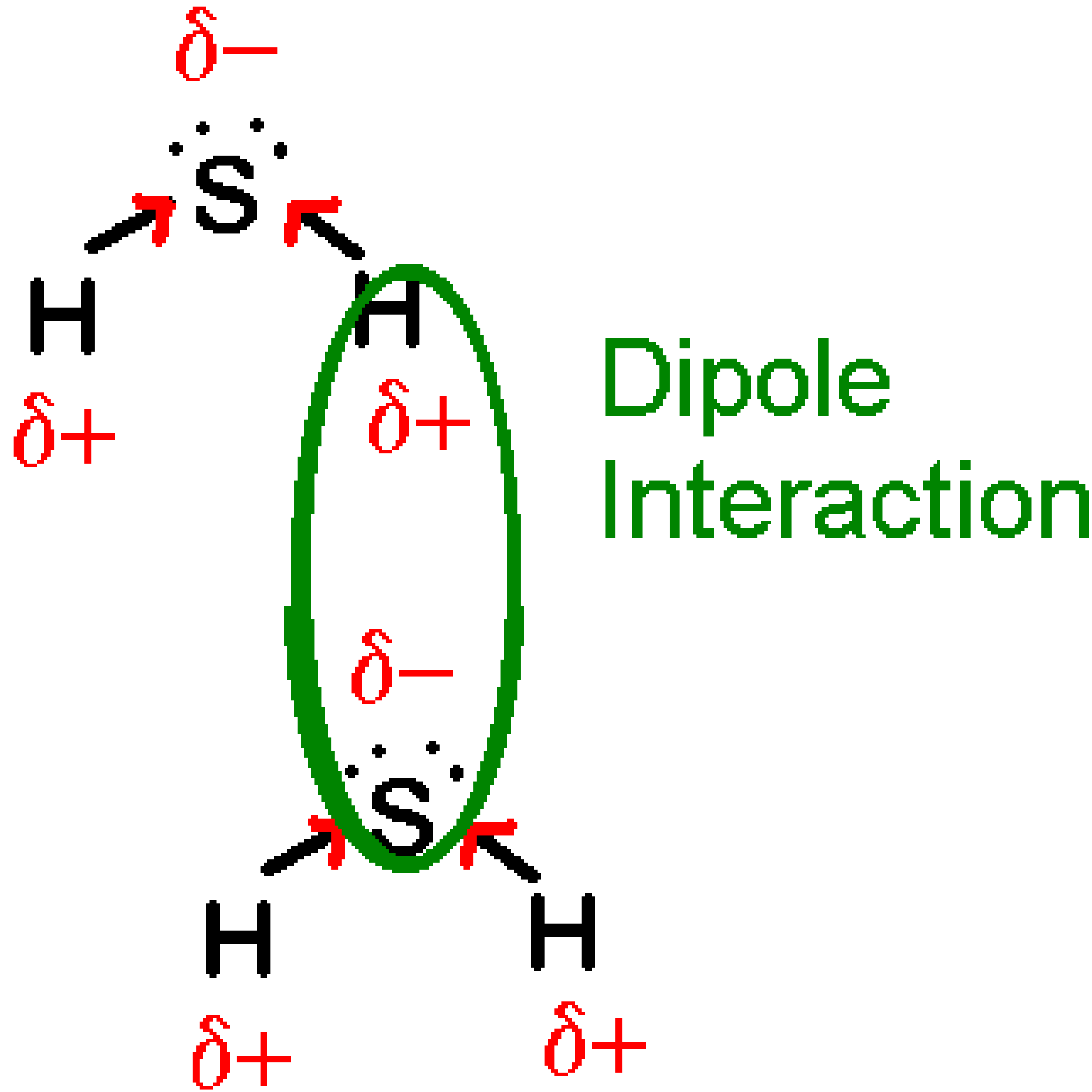
- **Dipole Interactions** = a weak intermolecular force resulting from the attraction of **oppositely charged regions of polar molecules**

**the slightly negative region of a polar molecule is attracted to the slightly positive region of another polar molecule*

Dipole Forces

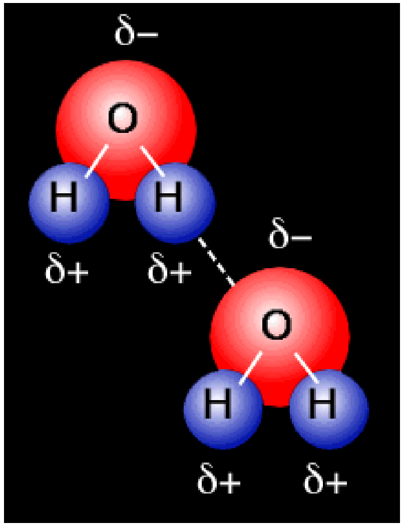


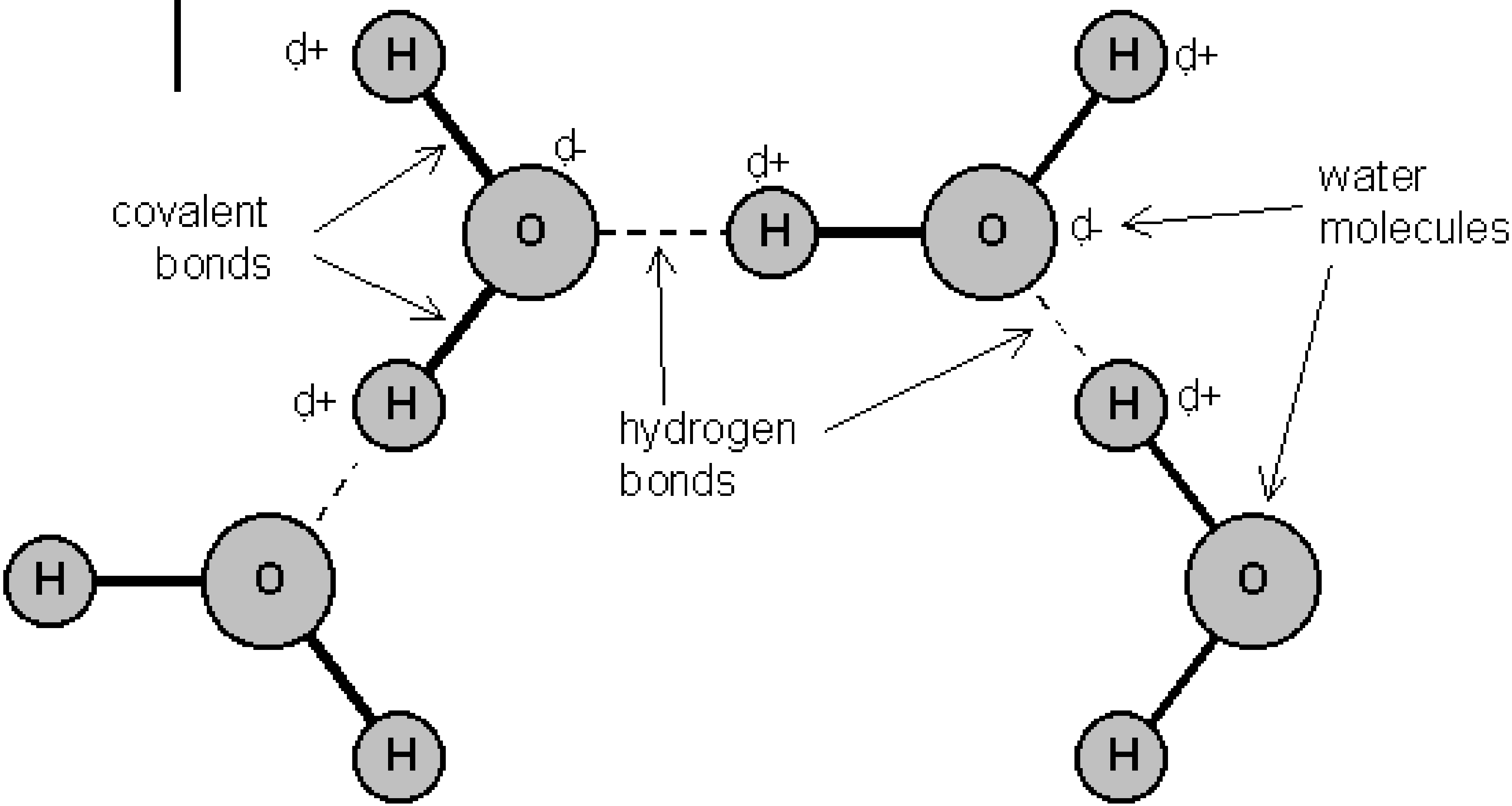
C. Ophardt, c. 2003

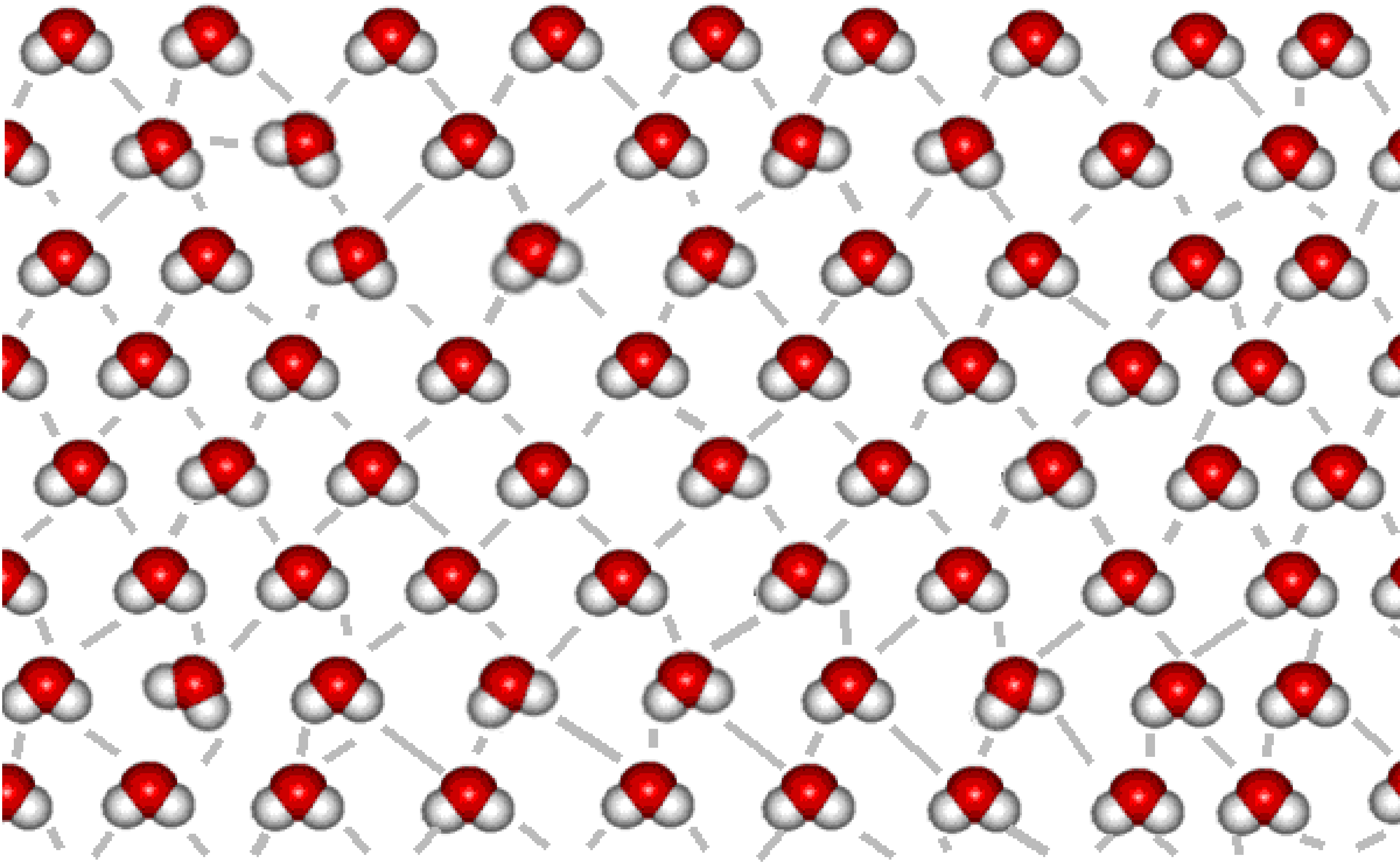

δ−
S
H
H
δ+
δ+
Dipole
Interaction
δ−
S
H
H
δ+
δ+

Hydrogen Bonds:

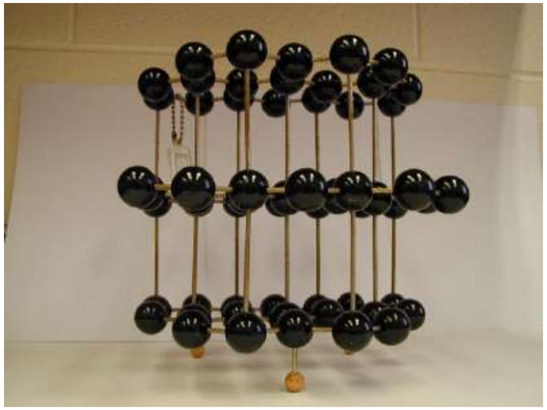
- A relatively strong intermolecular force in which a hydrogen atom that is covalently bonded to a very electronegative atom is weakly bonded to an **unshared electron pair** of another electronegative atom in a nearby molecule
- **EX: H_2O, DNA**




d+
H
d-
covalent
bonds
O
d+
H
O
d-
water
molecules
d+
H
hydrogen
bonds
H
O
H
H
O
H

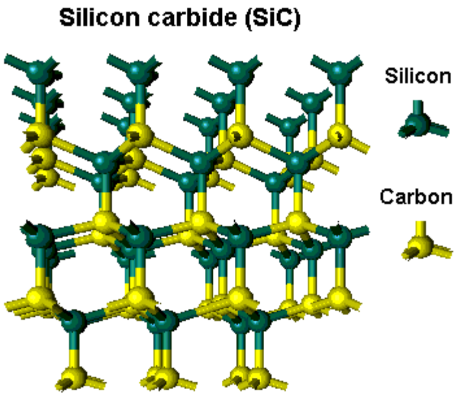


Network Solids:



- **Network** solid = a substance in which all atoms are covalently bonded to each other
- **EX:**

diamond, silicon carbide



and….

- Bond types and intermolecular attraction give rise to the physical and chemical properties of a compound!!
- **EX:**

  Physical state, melting or boiling point, solubility, conductivity, etc.

Comparing Molecular & Ionic Compounds:

	IONIC COMPOUNDS:	MOLECULAR COMPOUNDS:
Representative unit?	FORMULA UNIT	MOLECULE
Melting & boiling pts:	HIGH	LOW
Physical state at room temp:	SOLID	LIQUID or GAS
Formed by:	Transfer of electrons	Sharing of electrons
Formed from:	Metal + nonmetal	Nonmetal + nonmetal
Solubility in water:	Usually high	High to low
Conduct electricity?	Yes – good conductor	Poor to nonconducting