

# Chapter 2: Bonding

Dr. Feras Fraige

## Chapter Outline

- **Review of Atomic Structure**  
Electrons, protons, neutrons, quantum mechanics of atoms, electron states, the periodic Table
- **Atomic Bonding in Solids**  
Bonding energies and forces
- **Primary Interatomic Bonding**  
Ionic  
Covalent  
Metallic
- **Secondary Bonding**  
Three types of dipole-dipole bonds
- **Molecules and molecular solids**

Understanding of interatomic bonding is the first step towards understanding/explaining materials properties

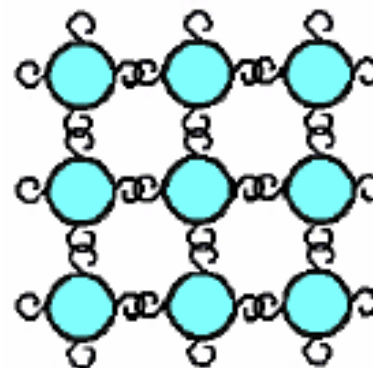
## Nature of Interatomic Bonding

Why the individual atoms coalesce into larger structures and take on the characteristics and properties of many different materials?

People were trying to answer this question for well over two millennia, since the time of the atomic hypothesis of Democritus, 440 B.C.\* Roman poet Lucretius (95-55 B.C.) wrote in *De Rerum Natura* (On the Nature of Things):

“What seems to us the hardened and condensed  
Must be of atoms among themselves more hooked,  
Be held compacted deep within, as it were  
By branch-like atoms- of which sort the chief  
Are diamond stones, despisers of all blows,  
And stalwart flint and strength of solid iron...”

John Dalton (1766-1844) found the evidence of those "hooks" in his quantitative chemical measurements, making the foundation of modern atomic theory of matter.



\* the idea that everything is made of distinct atoms has been a subject of skeptical discussions as recently as the beginning of the twentieth century, before Einstein's observation of Brownian motion in 1905 and Max von Laue's observation of the diffraction of X-rays by crystals in 1912 provided strong support for the atomistic theory.

## Structure of atoms: Brief review

The bonding mechanisms between atoms are closely related to the structure of the atoms themselves.

Atoms = nucleus (protons and neutrons) + electrons

### Charges:

Electrons and protons have negative and positive charges of the same magnitude,  $1.6 \times 10^{-19}$  Coulombs.

Neutrons are electrically neutral.

### Masses:

Protons and Neutrons have the same mass,  $1.67 \times 10^{-27}$  kg.

Mass of an electron is much smaller,  $9.11 \times 10^{-31}$  kg and can be neglected in calculation of atomic mass.

**The atomic mass (A)** = mass of protons + mass of neutrons

# protons gives chemical identification of the element

# **protons = atomic number (Z)**

# neutrons defines isotope number

## Atomic mass units. Atomic weight.

The atomic mass unit (amu) is often used to express atomic weight. 1 amu is defined as 1/12 of the atomic mass of the most common isotope of carbon atom that has 6 protons ( $Z=6$ ) and six neutrons ( $N=6$ ).

$$M_{\text{proton}} \approx M_{\text{neutron}} = 1.66 \times 10^{-24} \text{ g} = 1 \text{ amu.}$$

The atomic mass of the  $^{12}\text{C}$  atom is 12 amu.

The atomic weight of an element = weighted average of the atomic masses of the atoms naturally occurring isotopes. Atomic weight of carbon is 12.011 amu.

The atomic weight is often specified in mass per mole.

A mole is the amount of matter that has a mass in grams equal to the atomic mass in amu of the atoms (A mole of carbon has a mass of 12 grams).

The number of atoms in a mole is called the Avogadro number,  $N_{\text{av}} = 6.023 \times 10^{23}$ .

$$1 \text{ amu/atom} = 1 \text{ gram/mol}$$

Example:

$$\text{Atomic weight of iron} = 55.85 \text{ amu/atom} = 55.85 \text{ g/mol}$$

## Some simple calculations

The number of atoms per  $\text{cm}^3$ ,  $n$ , for material of density  $d$  ( $\text{g}/\text{cm}^3$ ) and atomic mass  $M$  ( $\text{g}/\text{mol}$ ):

$$n = N_{\text{av}} \times d / M$$

Graphite (carbon):  $d = 2.3 \text{ g}/\text{cm}^3$ ,  $M = 12 \text{ g}/\text{mol}$

$$n = 6 \times 10^{23} \text{ atoms}/\text{mol} \times 2.3 \text{ g}/\text{cm}^3 / 12 \text{ g}/\text{mol} = 11.5 \times 10^{22} \text{ atoms}/\text{cm}^3$$

Diamond (carbon):  $d = 3.5 \text{ g}/\text{cm}^3$ ,  $M = 12 \text{ g}/\text{mol}$

$$n = 6 \times 10^{23} \text{ atoms}/\text{mol} \times 3.5 \text{ g}/\text{cm}^3 / 12 \text{ g}/\text{mol} = 17.5 \times 10^{22} \text{ atoms}/\text{cm}^3$$

Water ( $\text{H}_2\text{O}$ )  $d = 1 \text{ g}/\text{cm}^3$ ,  $M = 18 \text{ g}/\text{mol}$

$$n = 6 \times 10^{23} \text{ molecules}/\text{mol} \times 1 \text{ g}/\text{cm}^3 / 18 \text{ g}/\text{mol} = 3.3 \times 10^{22} \text{ molecules}/\text{cm}^3$$

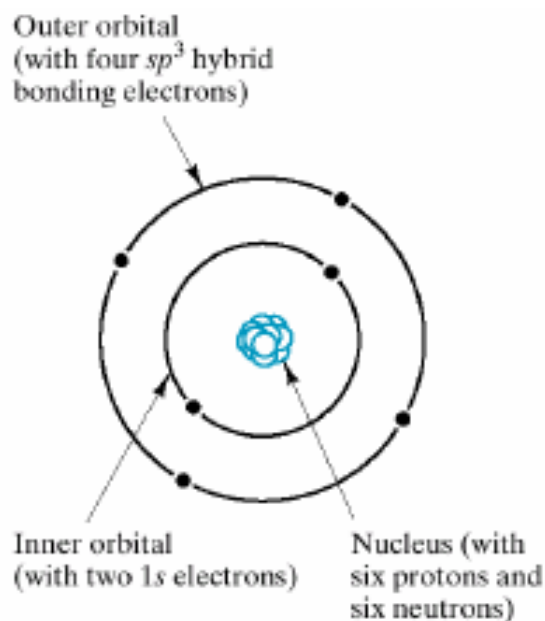
For material with  $n = 6 \times 10^{22} \text{ atoms}/\text{cm}^3$  we can calculate mean distance between atoms  $L = (1/n)^{1/3} = 0.25 \text{ nm}$ .

□ the scale of atomic structures in solids – a fraction of 1 nm or a few Å.

## Electrons in Atoms (I)

The electrons form a cloud around the nucleus, of radius of 0.05 – 2 nm.

This picture looks like a mini planetary system. But quantum mechanics tells us that this analogy is not correct:



Electrons move not in circular orbits, but in 'fuzzy' orbits. Actually, we cannot tell how it moves, but only can say what is the probability of finding it at some distance from the nucleus.

Only certain "orbits" or shells of electron probability densities are allowed. The shells are identified by a **principal quantum number  $n$** , which can be related to the size of the shell,  $n = 1$  is the smallest;  $n = 2, 3 \dots$  are larger.

**The second quantum number  $l$** , defines subshells within each shell. Two more quantum numbers characterize states within the subshells.

## Electrons in Atoms (II)

- The quantum numbers arise from solution of Schrodinger's equation
- Pauli Exclusion Principle: only one electron can have a given set of the four quantum numbers.

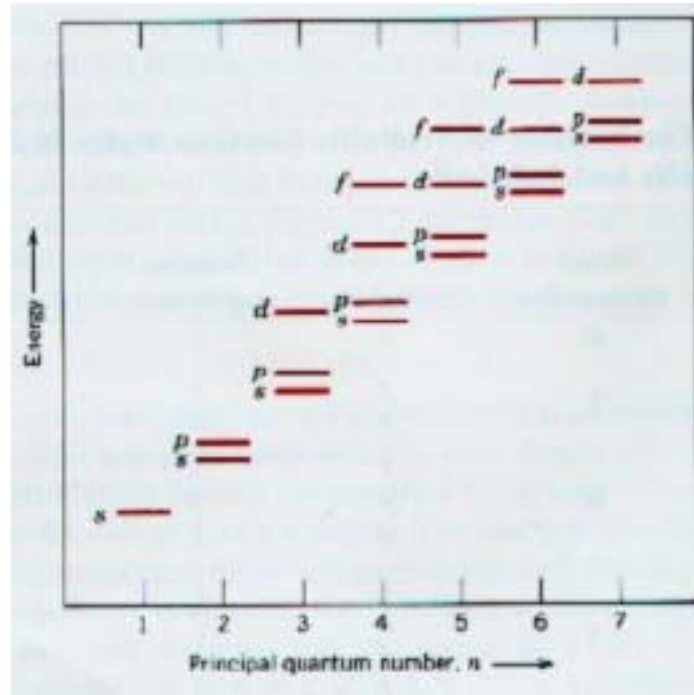
### The number of available states in electron shells & subshells

Principal Q. N., $n$	Subshells	Number of States	Number of Electrons	
			Per Subshell	Per Shell
1 ( $l=0$ )	<i>K-shell</i> s	1	2	2
2 ( $l=0$ )	s	1	2	8
2 ( $l=1$ )	<i>L-shell</i> p	3	6	
3 ( $l=0$ )	s	1	2	18
3 ( $l=1$ )	<i>M-shell</i> p	3	6	
3 ( $l=2$ )	d	5	10	
4 ( $l=0$ )	s	1	2	32
4 ( $l=1$ )	<i>N-shell</i>	p	3	6
4 ( $l=2$ )		d	5	10
4 ( $l=3$ )		f	7	14

Each "orbit" or shell can accommodate only a maximum number of electrons, which is determined by quantum mechanics. In brief, the most inner *K-shell* can accommodate only two electrons, called *s-electrons*; the next *L-shell* two *s-electrons* and six *p-electrons*; the *M-shell* can host two *s-electrons*, six *p-electrons*, and ten *d-electrons*; and so on.



## Electrons in Atoms (III)



Subshells by energy: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, ...

- Electrons that occupy the outermost filled shell – **the valence electrons** – they are responsible for bonding.
- Electrons fill quantum levels in order of increasing energy (only n,  $\ell$  make a significant difference).

Examples: Argon,  $Z = 18$ :  $1s^2 2s^2 2p^6 3s^2 3p^6$

Iron,  $Z = 26$ :  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

## Electrons in Atoms (IV)

<u>Element</u>	<u>Atomic #</u>	<u>Electron configuration</u>
Hydrogen	1	$1s^1$
Helium	2	$1s^2$ (stable)
Lithium	3	$1s^2 2s^1$
Beryllium	4	$1s^2 2s^2$
Boron	5	$1s^2 2s^2 2p^1$
Carbon	6	$1s^2 2s^2 2p^2$
...	...	...
Neon	10	$1s^2 2s^2 2p^6$ (stable)
Sodium	11	$1s^2 2s^2 2p^6 3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
...	...	...
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$ (stable)
...	...	...
Krypton	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ (stable)

Electron configurations where all states within valence electron shell are filled are stable → unreactive inert or noble gas.

## Periodic Table

The first accepted periodic table of elements was published in 1869 by Mendeleev. In the same year, a German chemist Lothar Meyer independently published a very similar table, but his contribution is generally ignored.

All elements in the periodic table have been classified **according to the electron configuration.**

Periodic Table

Li	Na	K	Ca	Mg	Zn	Fe	Ni	Co	Cu	Ag	Au	Pt	Pd	Rh	Ru	Rd	Rf	Rg	Ca	Sc	Ti	V	Cr	Mn	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	I	Xe	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
----	----	---	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	---	----	----	----	----	----	----	----	----	----	----	----	---	----	----	----	----	----	---	----	----	----	----	----	----	----	----	----	----	----

Draft of the periodic table,  
Mendeleev, 1869

# Periodic Table

The periodic table is color-coded into three regions: Metal (light blue), Nonmetal (medium blue), and Intermediate (dark blue). Annotations indicate electron transfer tendencies: red text for elements that 'give up' electrons (IA, IIA, and the first two columns of the transition metals), blue text for elements that 'accept' electrons (groups VIA, VIIA, and the first two columns of the transition metals), and green text for inert gases (group 0).

1	2											3	4	5	6	7	8	9	10								
IA	IIA											IIIA	IVA	VA	VIA	VIIA	0										
H	He											B	C	N	O	F	Ne										
3	4											13	14	15	16	17	18										
Li	Be											Al	Si	P	S	Cl	Ar										
11	12	13	14	15	16	17	18											31	32	33	34	35	36				
Na	Mg	Al	Si	P	S	Cl	Ar											Ga	Ge	As	Se	Br	Kr				
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36										
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr										
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54										
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe										
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86										
Cs	Ba	Rare earth series	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn										
87	88	Actinide series	104	105	106	107	108	109	110																		
Fr	Ra	Actinide series	Rf	Db	Sg	Bh	Hs	Mt	Ds																		

Electropositive elements:  
Readily give up electrons  
to become + ions.

Electronegative elements:  
Readily acquire electrons  
to become - ions.

Elements in the same column (Elemental Group) share similar properties. Group number indicates the number of electrons available for bonding.

0: Inert gases (He, Ne, Ar...) have filled subshells: chem. inactive

IA: Alkali metals (Li, Na, K...) have one electron in outermost occupied s subshell - eager to give up electron - chem. active

VIIA: Halogens (F, Br, Cl...) missing one electron in outermost occupied p shell - want to gain electron - chem. active

# Periodic Table - Electronegativity

IA																0	
1	IIA											IIIA	IVA	VA	VIA	VIIA	8
H	He											B	C	N	O	F	Ne
2.1	1.5											2.0	2.5	3.0	3.5	4.0	-
3	4																10
Li	Be											Al	Si	P	S	Cl	Ar
1.0	1.5											1.5	1.8	2.1	2.5	3.0	-
11	12																18
Na	Mg	IIIB	IVB	VB	VIB	VII B	VIII			IB	IIB						18
0.9	1.2																-
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.8	1.6	1.8	2.0	2.4	2.8	-
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
0.8	1.0	1.2	1.4	1.5	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	-
55	56	57-71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
0.7	0.9	1.1-1.2	1.5	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2	-
87	88	89-103															
Fr	Ra	Ac-No															
0.7	0.9	1.1-1.3															

Figure 2.7 from the textbook. The electronegativity values.

Electronegativity - a measure of how willing atoms are to accept electrons

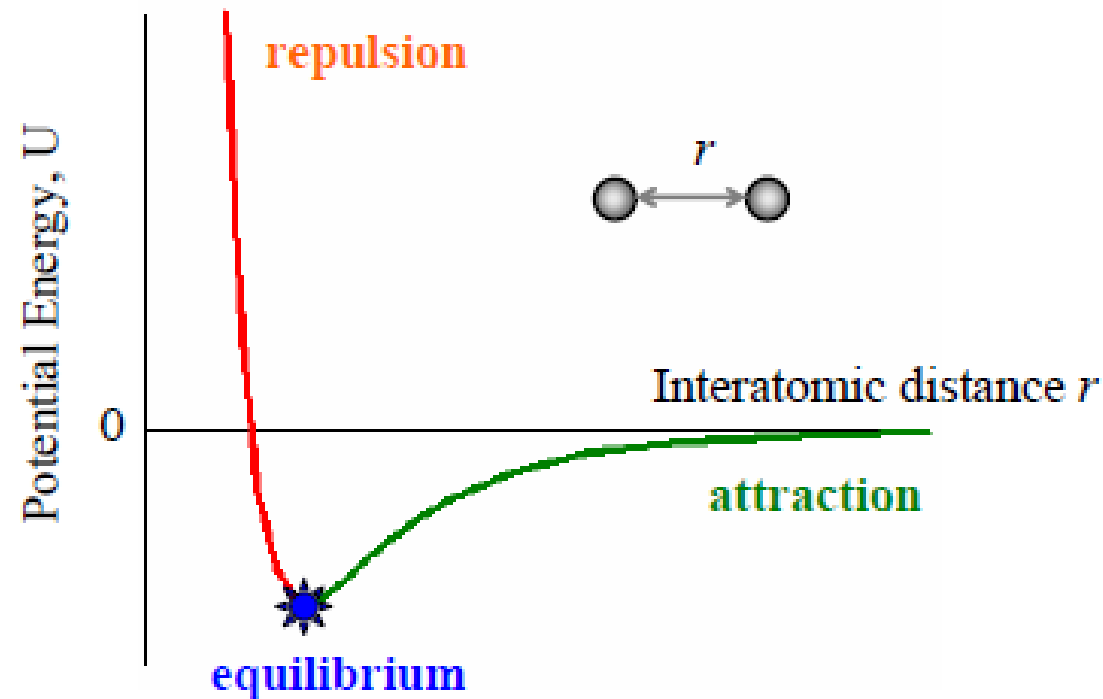
Subshells with one electron → low electronegativity

Subshells with one missing electron → high electronegativity

Electronegativity increases from left to right

Metals are electropositive – they can give up their few valence electrons to become positively charged ions

## Bonding Energies and Forces



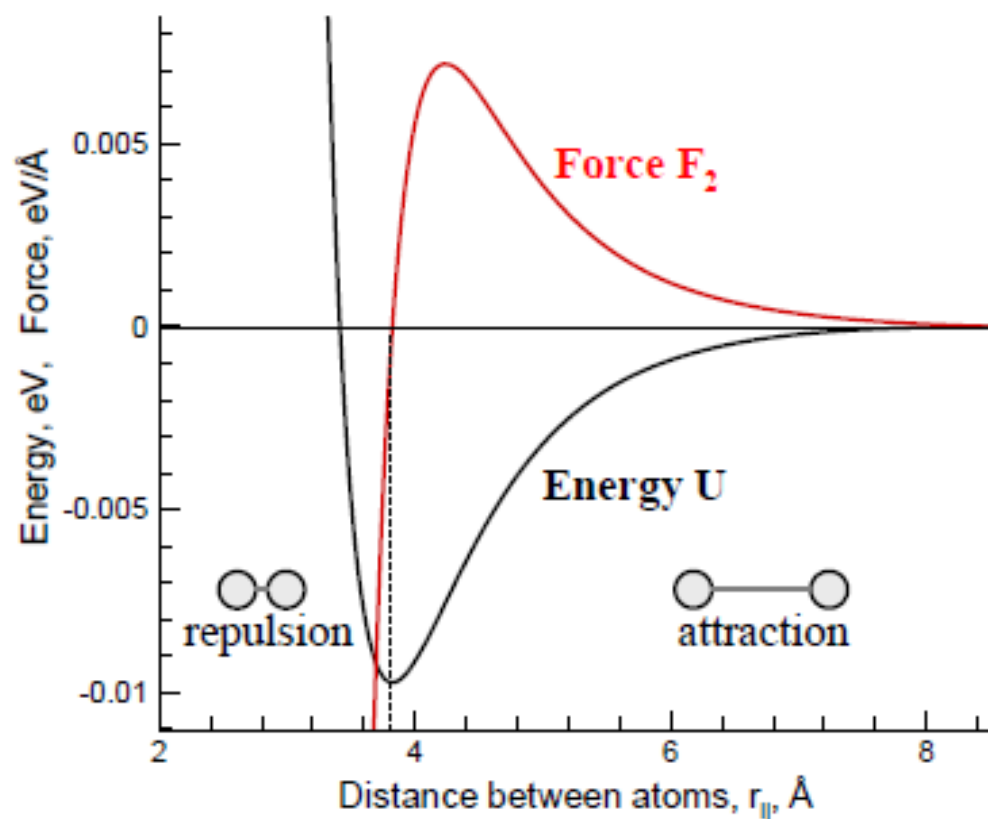
This is typical potential well for two interacting atoms

The repulsion between atoms, when they are brought close to each other, is related to the Pauli principle: when the electronic clouds surrounding the atoms starts to overlap, the energy of the system increases abruptly.

The origin of the attractive part, dominating at large distances, depends on the particular type of bonding.

## Bonding Energies and Forces

Forces can be calculated from the potential energy of interatomic interaction. For example, for a system of two atoms (e.g. a diatomic molecule), the potential depends only on the distance between the two atoms  $U(r_{12})$



$$\vec{F}_1 = -\vec{F}_2 = -\frac{dU(r_{12})}{dr_{12}}$$

Diagram showing two atoms (labeled 1 and 2) with forces  $\vec{F}_1$  and  $\vec{F}_2$  acting on them. The distance vector is defined as  $r_{12} = |\vec{r}_1 - \vec{r}_2|$ .

## The electron volt (eV) – energy unit convenient for description of atomic bonding

Electron volt - the energy lost / gained by an electron when it is taken through a potential difference of one volt.

$$E = q \times V$$

for  $q = 1.6 \times 10^{-19}$  Coulombs and  $V = 1$  volt

$$1 \text{ eV} = 1.6 \times 10^{-19} \text{ J}$$

## Types of Bonding

The electronic structure of atoms defines the character of their interaction among each other. Filled outer shells result in a stable configuration as in noble inert gases. Atoms with incomplete outer shells strive to reach this noble gas configuration by sharing or transferring electrons among each other for maximal stability. Strong “primary” bonding results from the electron sharing or transfer.



# Types of Bonding

**Primary bonding:**  $e^-$  are transferred or shared

**Strong (100-1000 KJ/mol or 1-10 eV/atom)**

- Ionic: Strong Coulomb interaction among negative atoms (have an extra electron each) and positive atoms (lost an electron). Example -  $\text{Na}^+\text{Cl}^-$
- Covalent: electrons are shared between the molecules, to saturate the valency. Example -  $\text{H}_2$
- Metallic: the atoms are ionized, losing some electrons from the valence band. Those electrons form an electron sea, which binds the charged nuclei in place

**Secondary Bonding:** no  $e^-$  transferred or shared

**Interaction of atomic/molecular dipoles**

**Weak ( $< 100$  KJ/mol or  $< 1$  eV/atom)**

- Fluctuating Induced Dipole (inert gases,  $\text{H}_2$ ,  $\text{Cl}_2$ ...)
- Permanent dipole bonds (polar molecules -  $\text{H}_2\text{O}$ ,  $\text{HCl}$ ...)
- Polar molecule-induced dipole bonds (a polar molecule induces a dipole in a nearby nonpolar atom/molecule)

## Ionic Bonding (I)

**Ionic Bonding** is typical for elements that are situated at the horizontal extremities of the periodic table.

Atoms from the left (metals) are ready to give up their valence electrons to the (non-metallic) atoms from the right that are happy to get one or a few electrons to acquire stable or noble gas electron configuration.

As a result of this transfer mutual ionization occurs: atom that gives up electron(s) becomes positively charged ion (**cation**), atom that accepts electron(s) becomes negatively charged ion (**anion**).

Formation of ionic bond:

1. Mutual ionization occurs by electron transfer (remember electronegativity table)
  - Ion = charged atom
  - Anion = negatively charged atom
  - Cation = positively charged atom
2. Ions are attracted by strong coulombic interaction
  - Oppositely charged atoms attract each other
  - An ionic bond is non-directional (ions may be attracted to one another in any direction)

## Ionic Bonding (II)

*Example: table salt (NaCl)*

Na has 11 electrons, 1 more than needed for a full outer shell (Neon)

11 Protons Na  $1S^2 2S^2 2P^6 3S^1$

11 Protons Na<sup>+</sup>  $1S^2 2S^2 2P^6$



donates  $e^-$   
10  $e^-$  left

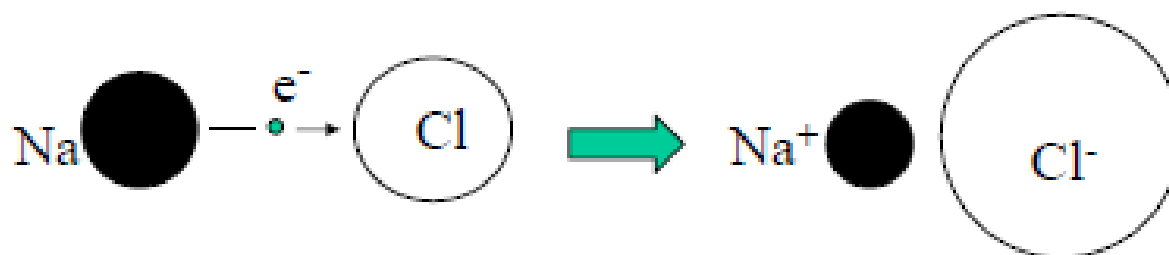
Cl has 17 electron, 1 less than needed for a full outer shell (Argon)

17 Protons Cl  $1S^2 2S^2 2P^6 3S^2 3P^5$

17 Protons Cl<sup>-</sup>  $1S^2 2S^2 2P^6 3S^2 3P^6$



receives  $e^-$   
18  $e^-$

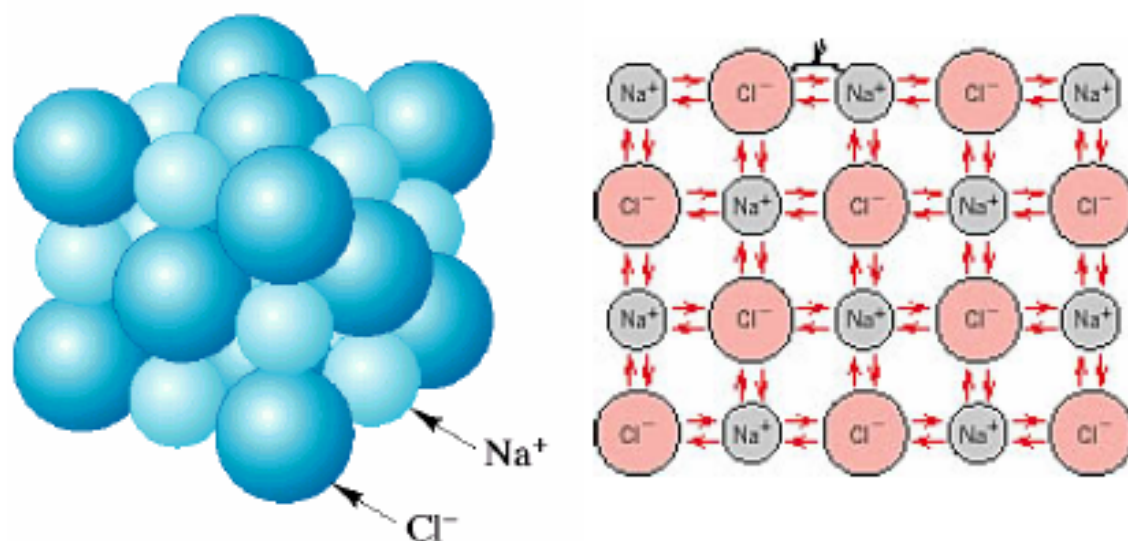


- Electron transfer reduces the energy of the system of atoms, that is, electron transfer is energetically favorable
- Note relative sizes of ions: Na shrinks and Cl expands

## Ionic Bonding (III)

A strong electrostatic attraction between positively charged  $\text{Na}^+$  ions and negatively charged  $\text{Cl}^-$  atoms along with  $\text{Na}^+$  -  $\text{Na}^+$  and  $\text{Cl}^-$  -  $\text{Cl}^-$  repulsion result in the  $\text{NaCl}$  crystal structure which is arranged so that each sodium ion is surrounded by  $\text{Cl}^-$  ions and each  $\text{Na}^+$  ion is surrounded by  $\text{Cl}^-$  ions, see the figure on the left.

Any mechanical force that tries to disturb the electrical balance in an ionic crystal meets strong resistance: ionic materials are strong and brittle. In some special cases, however, significant plastic deformation can be observed, e.g.  $\text{NaCl}$  single crystals can be bent by hand in water.



**Ionic bonds: very strong, nondirectional bonds**

## Ionic Bonding (IV)

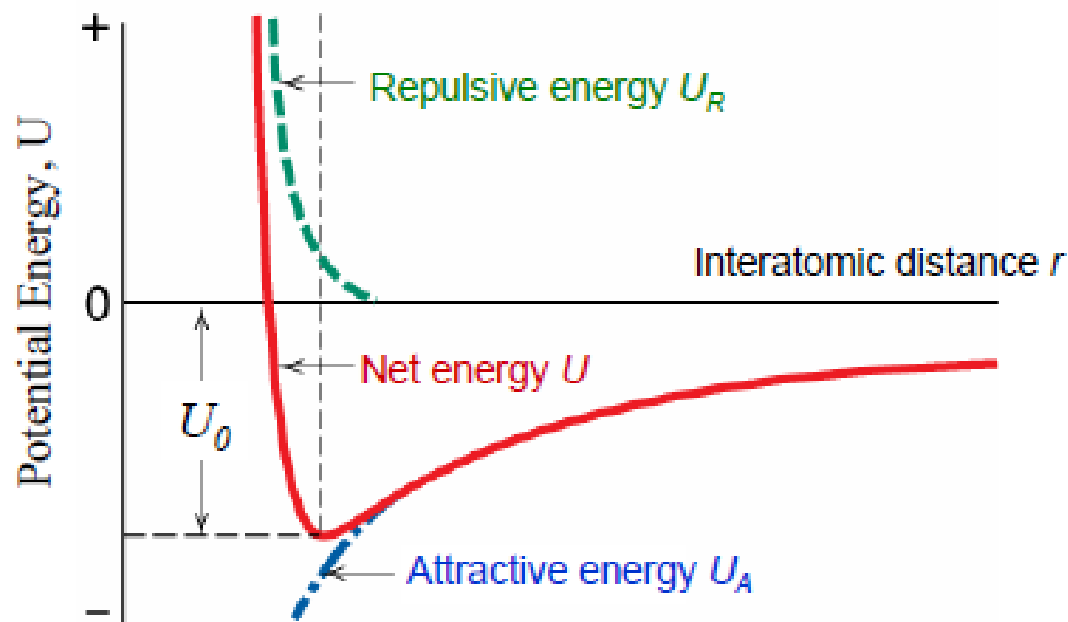
Attractive coulomb interaction between charges of opposite sign:

$$U_A = \frac{1}{4\pi\epsilon_0} \frac{q_1 q_2}{r} = -\frac{A}{r}$$

Repulsion due to the overlap of electron clouds at close distances (Pauli principle of QM):

$$U_R = \frac{B}{r^n}$$

$$U = U_A + U_R = -\frac{A}{r} + \frac{B}{r^n}$$



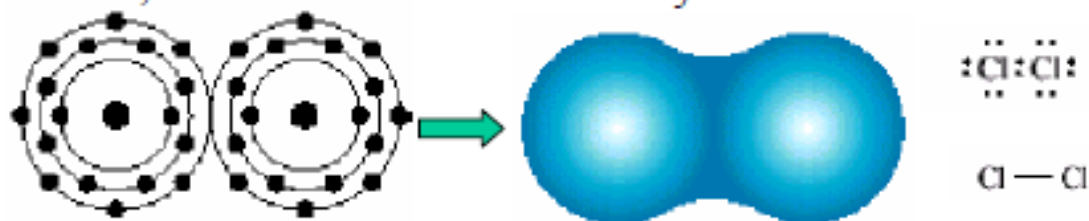
## Covalent Bonding (I)

In covalent bonding, electrons are **shared** between the molecules, to saturate the valency. In this case the electrons are not transferred as in the ionic bonding, but they are localized between the neighboring ions and form **directional** bond between them. The ions repel each other, but are attracted to the electrons that spend most of the time in between the ions.

### Formation of covalent bonds:

- Cooperative sharing of valence electrons
- Can be described by orbital overlap
- Covalent bonds are **HIGHLY directional**
- Bonds - in the direction of the greatest orbital overlap
- Covalent bond model: an atom can covalently bond with at most  $8 - N^v$ ,  $N^v$  = number of valence electrons

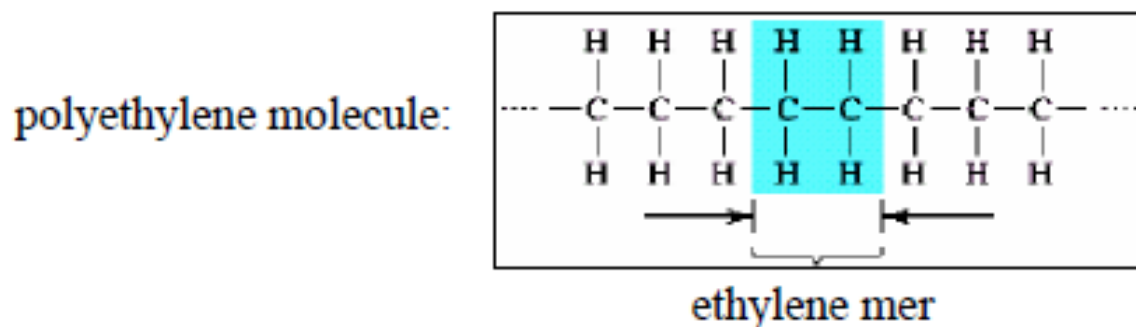
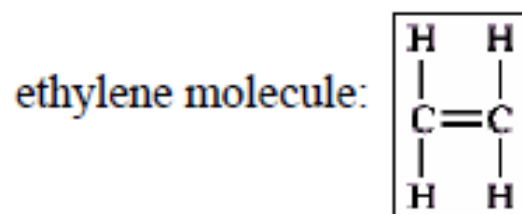
**Example:**  $\text{Cl}_2$  molecule.  $Z_{\text{Cl}} = 17$  ( $1s^2 2s^2 2p^6 3s^2 3p^5$ )  
 $N^v = 7$ ,  $8 - N^v = 1 \rightarrow$  can form only one covalent bond



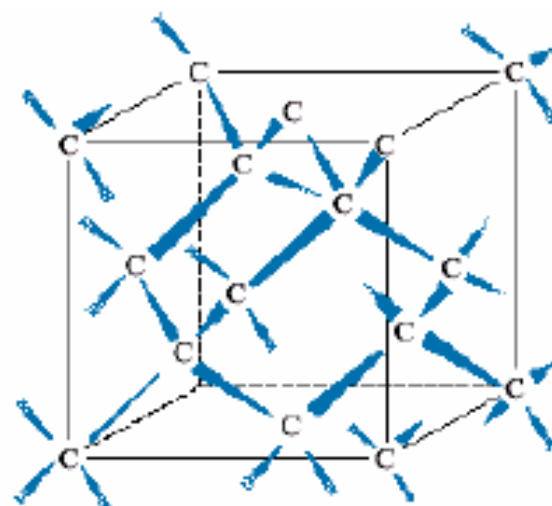
## Covalent Bonding (II)

Example: Carbon materials.  $Z_c = 6$  ( $1S^2 2S^2 2P^2$ )

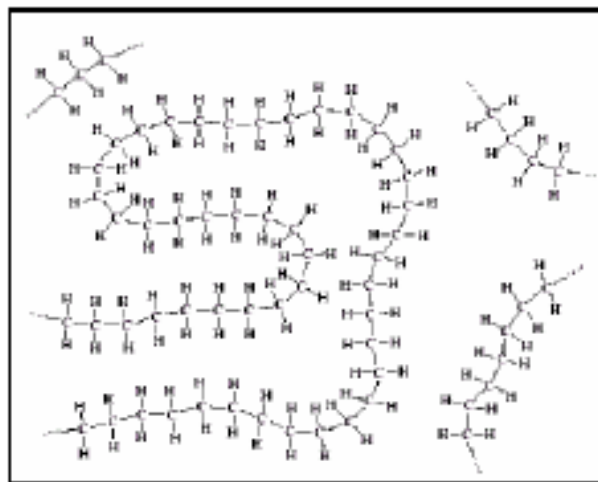
$N^v = 4$ ,  $8 - N^v = 4 \rightarrow$  can form up to four covalent bonds



diamond:  
(each C atom has four  
covalent bonds with four  
other carbon atoms)

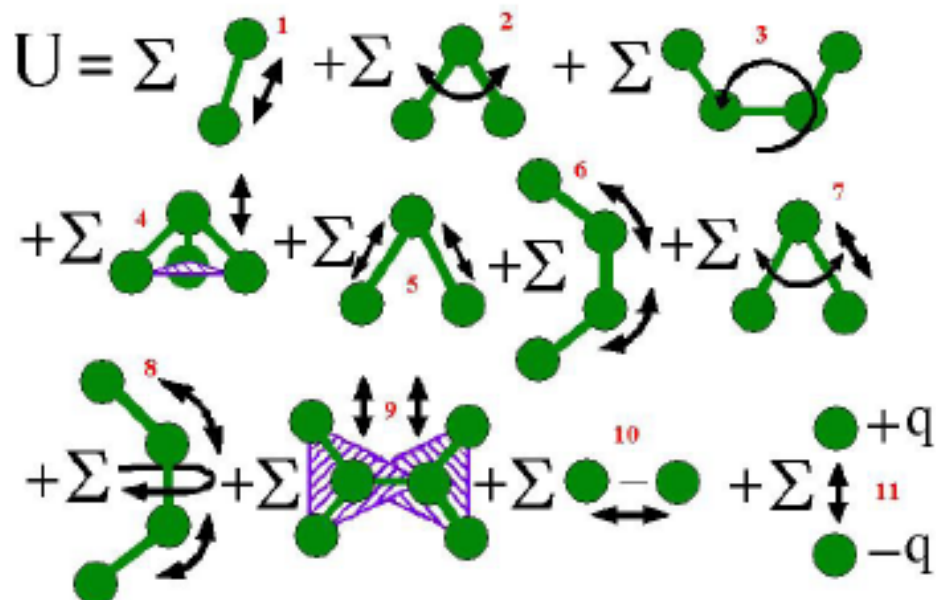


## Covalent Bonding (III)



2-D schematic of the  
“spaghetti-like” structure  
of solid polyethylene

The potential energy of a system of covalently interacting atoms depend not only on the distances between atoms, but also on angles between bonds...

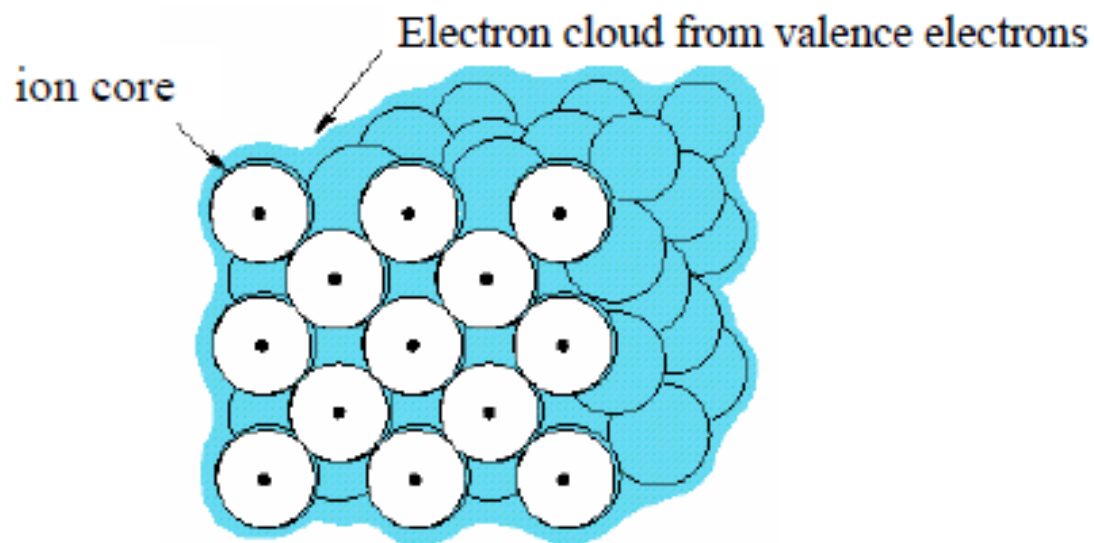




## Metallic Bonding

Valence electrons are detached from atoms, and spread in an 'electron sea' that "glues" the positive ions together.

- A metallic bond is non-directional (bonds form in any direction) → atoms pack closely



The "bonds" do not "break" when atoms are rearranged – metals can experience a significant degree of plastic deformation.

Examples of typical metallic bonding: Cu, Al, Au, Ag, etc. Transition metals (Fe, Ni, etc.) form mixed bonds that are comprising of metallic bonds and covalent bonds involving their 3d-electrons. As a result the transition metals are more brittle (less ductile) than Au or Cu.

## Secondary Bonding (I)

Secondary = van der Waals = physical (as opposite to chemical bonding that involves  $e^-$  transfer) bonding results from interaction of atomic or molecular dipoles and is weak,  $\sim 0.1$  eV/atom or  $\sim 10$  kJ/mol.



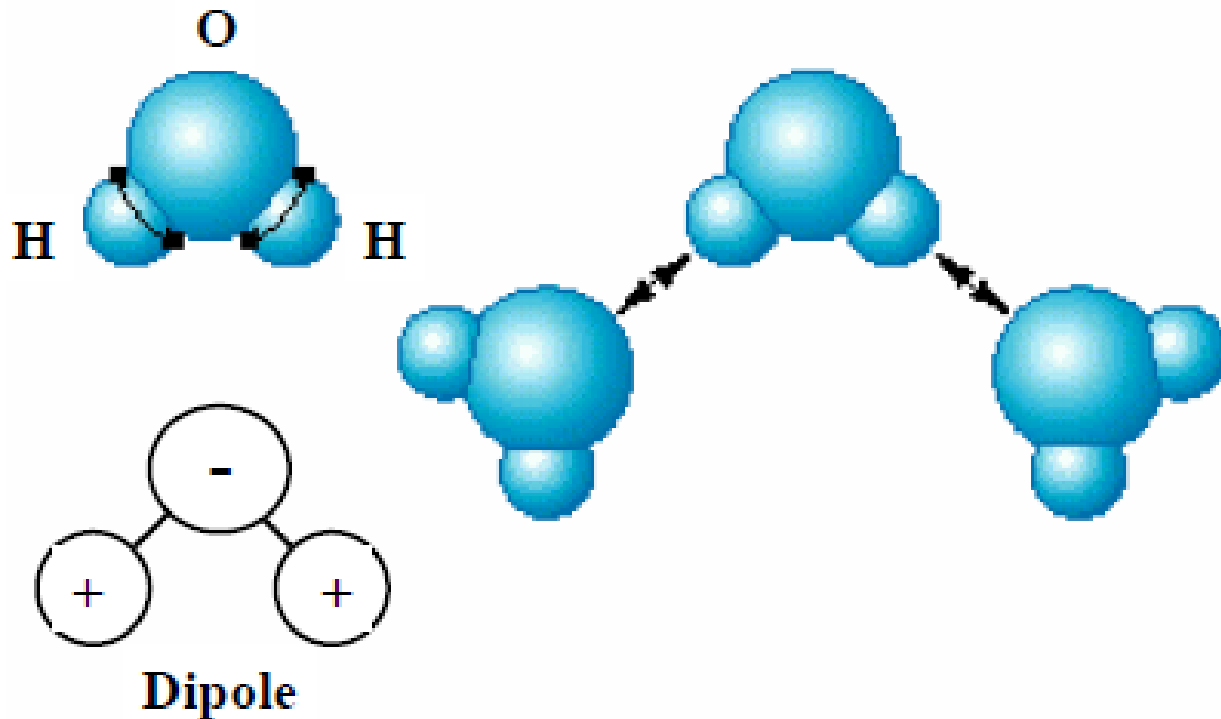
Permanent dipole moments exist in some molecules (called **polar molecules**) due to the asymmetrical arrangement of positively and negatively regions (HCl, H<sub>2</sub>O). Bonds between adjacent polar molecules – **permanent dipole bonds** – are the **strongest among secondary bonds**.

Polar molecules can **induce** dipoles in adjacent non-polar molecules and bond is formed due to the **attraction between the permanent and induced dipoles**.

Even in electrically symmetric molecules/atoms an electric dipole can be created by fluctuations of electron density distribution. Fluctuating electric field in one atom A is felt by the electrons of an adjacent atom, and induce a dipole momentum in this atom. This bond due to fluctuating induced dipoles is the **weakest** (inert gases, H<sub>2</sub>, Cl<sub>2</sub>).

## Secondary Bonding (II)

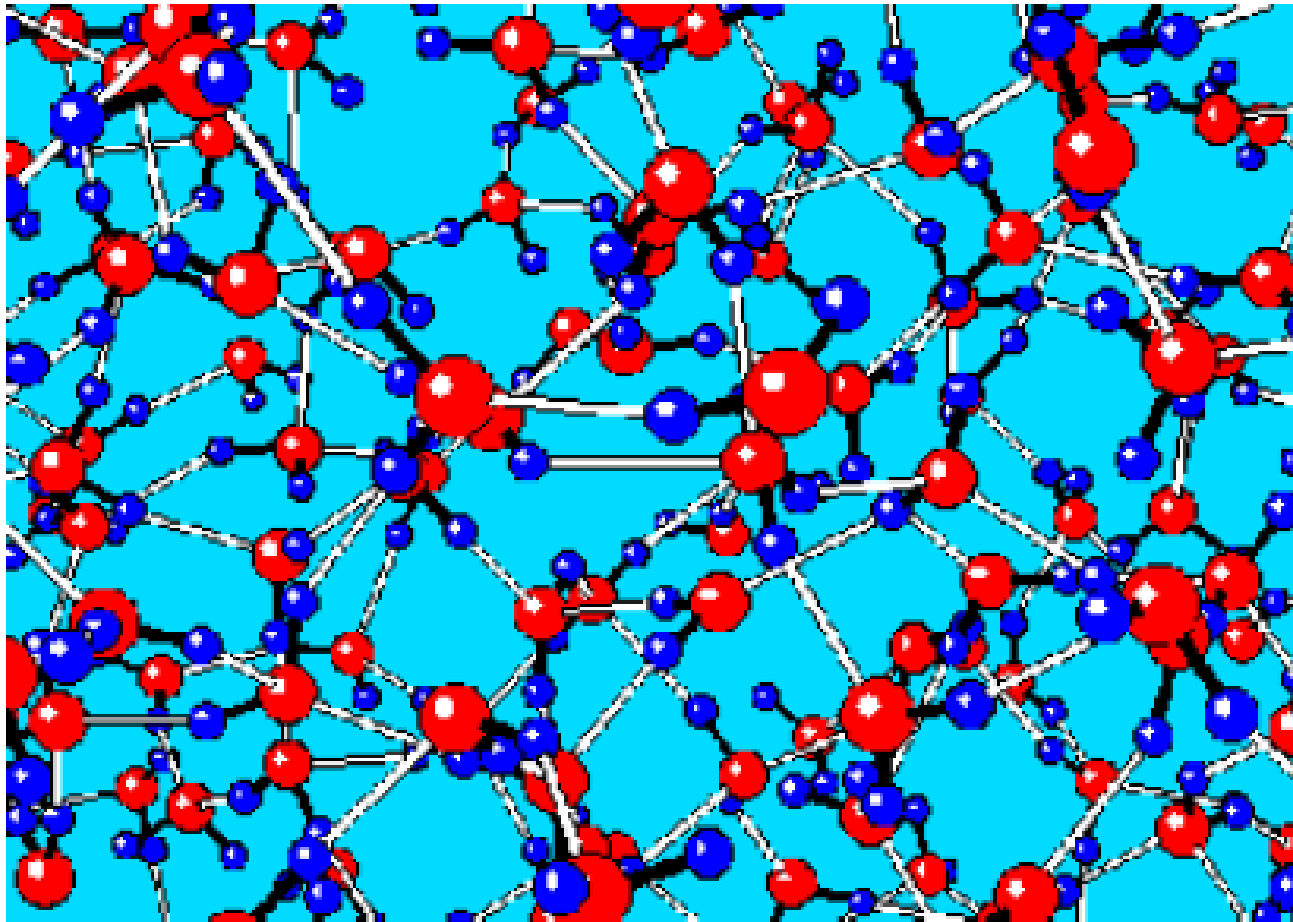
*Example:* hydrogen bond in water. The H end of the molecule is positively charged and can bond to the negative side of another H<sub>2</sub>O molecule (the O side of the H<sub>2</sub>O dipole)



“Hydrogen bond” – secondary bond formed between two permanent dipoles in adjacent water molecules.

## Secondary Bonding (III)

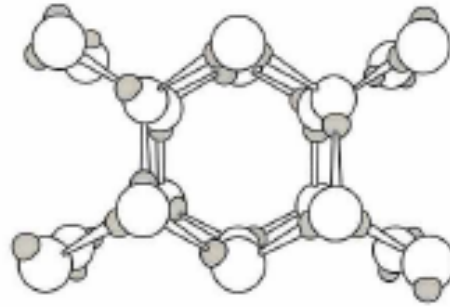
Hydrogen bonding in liquid water  
from a molecular-level simulation



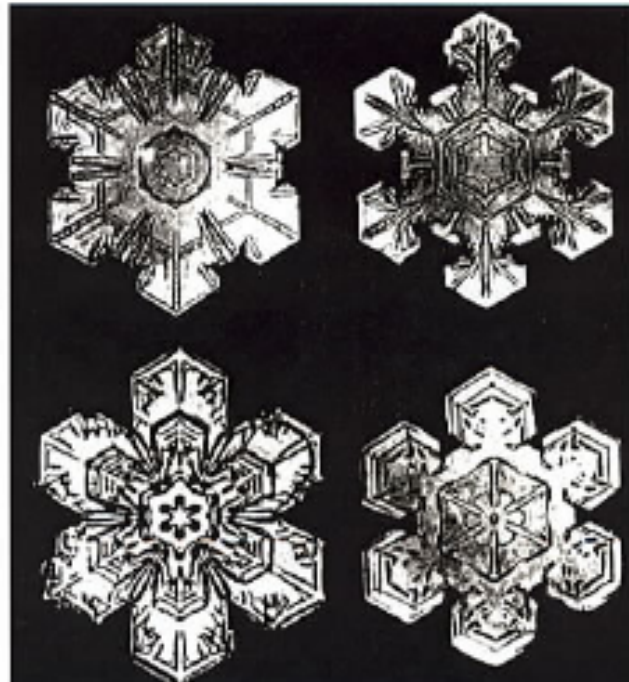
**Molecules:** Primary bonds inside, secondary bonds  
among each other

## Secondary Bonding (IV)

### The Crystal Structures of Ice



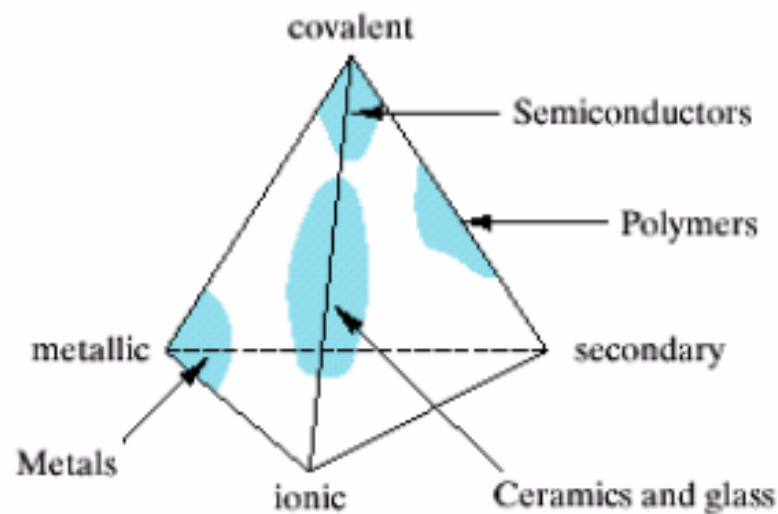
### Hexagonal Symmetry of Ice Snowflakes



Figures by Paul R. Howell

## Bonding in real materials

In many materials more than one type of bonding is involved (ionic and covalent in ceramics, covalent and secondary in polymers, covalent and ionic in semiconductors).



### Examples of bonding in Materials:

Metals: Metallic

Ceramics: Ionic / Covalent

Polymers: Covalent and Secondary

Semiconductors: Covalent or Covalent / Ionic

## Correlation between bonding energy and melting temperature

**Table 2.3** Bonding Energies and Melting Temperatures for Various Substances

<i>Bonding Type</i>	<i>Substance</i>	<i>Bonding Energy</i>		<i>Melting Temperature</i> (°C)
		<i>kJ/mol</i> ( <i>kcal/mol</i> )	<i>eV/Atom,</i> <i>Ion, Molecule</i>	
Ionic	NaCl	640 (153)	3.3	801
	MgO	1000 (239)	5.2	2800
Covalent	Si	450 (108)	4.7	1410
	C (diamond)	713 (170)	7.4	>3550
Metallic	Hg	68 (16)	0.7	-39
	Al	324 (77)	3.4	660
	Fe	406 (97)	4.2	1538
	W	849 (203)	8.8	3410
van der Waals	Ar	7.7 (1.8)	0.08	-189
	Cl <sub>2</sub>	31 (7.4)	0.32	-101
Hydrogen	NH <sub>3</sub>	35 (8.4)	0.36	-78
	H <sub>2</sub> O	51 (12.2)	0.52	0

## Summary

Make sure you understand language and concepts:

- Atomic mass unit (amu)
- Atomic number
- Atomic weight
- Bonding energy
- Coulombic force
- Covalent bond
- Dipole (electric)
- Electron state
- Electronegative
- Electropositive
- Hydrogen bond
- Ionic bond
- Metallic bond
- Mole
- Molecule
- Periodic table
- Polar molecule
- Primary bonding
- Secondary bonding
- Van der Waals bond
- Valence electron



## Reading for next class:

### **Chapter 3: The structure of crystalline solids**

#### **Unit cells**

#### **Crystal structures**

Face-centered cubic

Body-centered cubic

Hexagonal close-packed

#### **Density computations**

#### **Types of solids**

Single crystals

Polycrystalline

Amorphous

*Optional reading (Parts that are not covered / not tested):*

*3.8–3.10 Crystallography*

*3.15 Anisotropy*

*3.16 Diffraction*