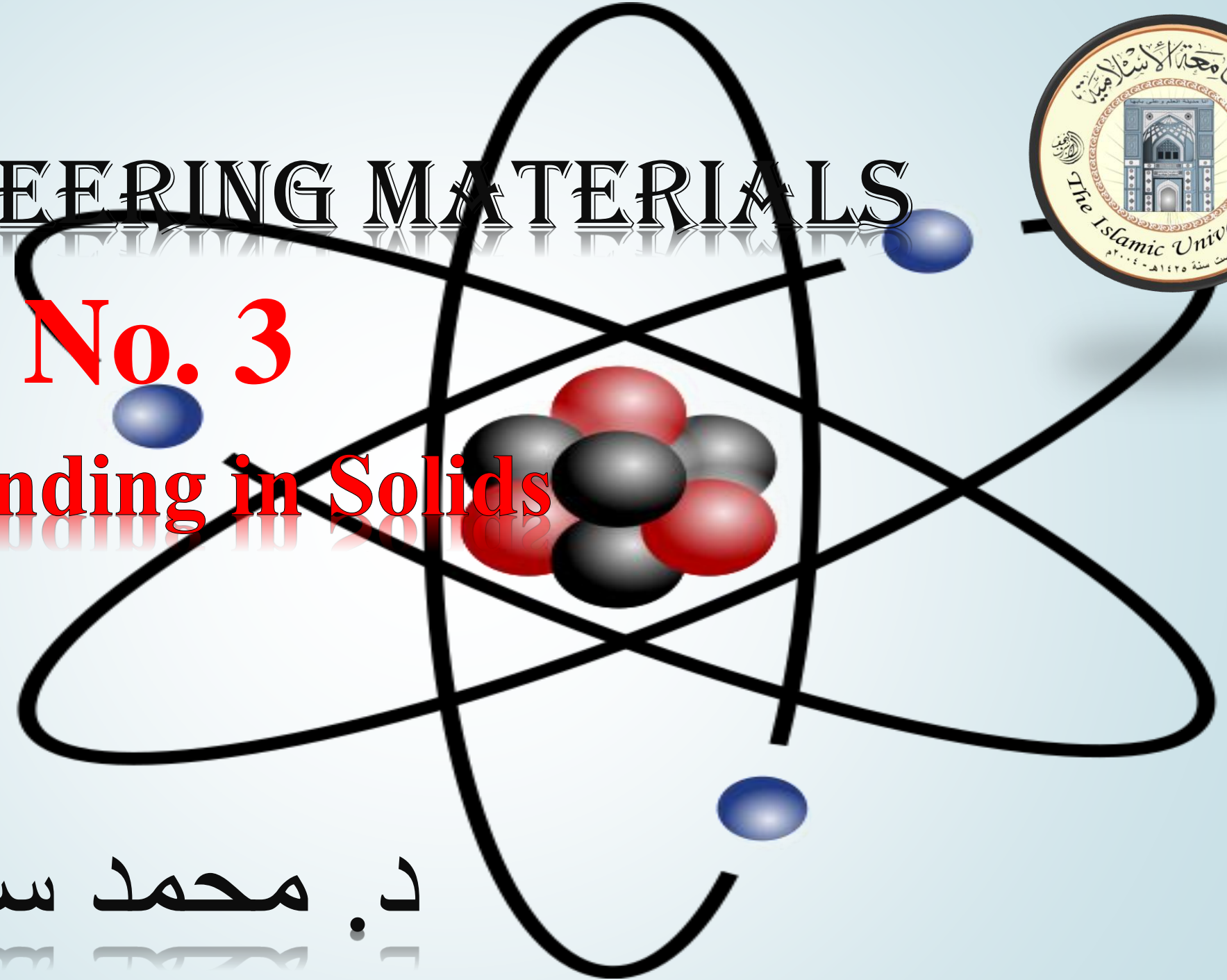




ENGINEERING MATERIALS

Lecture No. 3

Atomic Bonding in Solids



د. محمد سعد

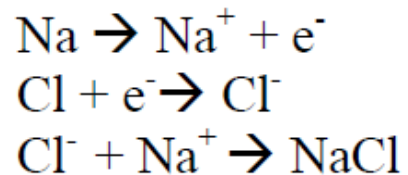
BONDING FORCES AND ENERGIES

An understanding of many of the physical properties of materials is enhanced by a knowledge of the interatomic forces that bind the atoms together. Perhaps the principles of atomic bonding are best illustrated by considering how two isolated atoms interact as they are brought close together from an infinite separation at large distances, interactions are negligible because the atoms are too far apart to have an influence on each other; however, at small separation distances, each atom exerts forces on the others. These forces are of two types, attractive (FA) and repulsive (FR), and the magnitude of each depends on the separation or interatomic distance (r);

PRIMARY INTERATOMIC BONDS

Ionic Bonding

Atoms have a tendency to achieve a completely filled valence shell like the inert gases. Metals like sodium tend to lose electrons to achieve full valence shells, and nonmetals like chlorine tend to gain electrons. The process of gaining or losing electrons creates ions, and electrostatic forces bring the ions together to form compounds. For example, sodium may lose an electron to chlorine to create Na^+ and Cl^- , and the Coulombic attraction between the resulting ions makes NaCl , or table salt. The attraction between the oppositely charged ions constitutes the ionic bond.

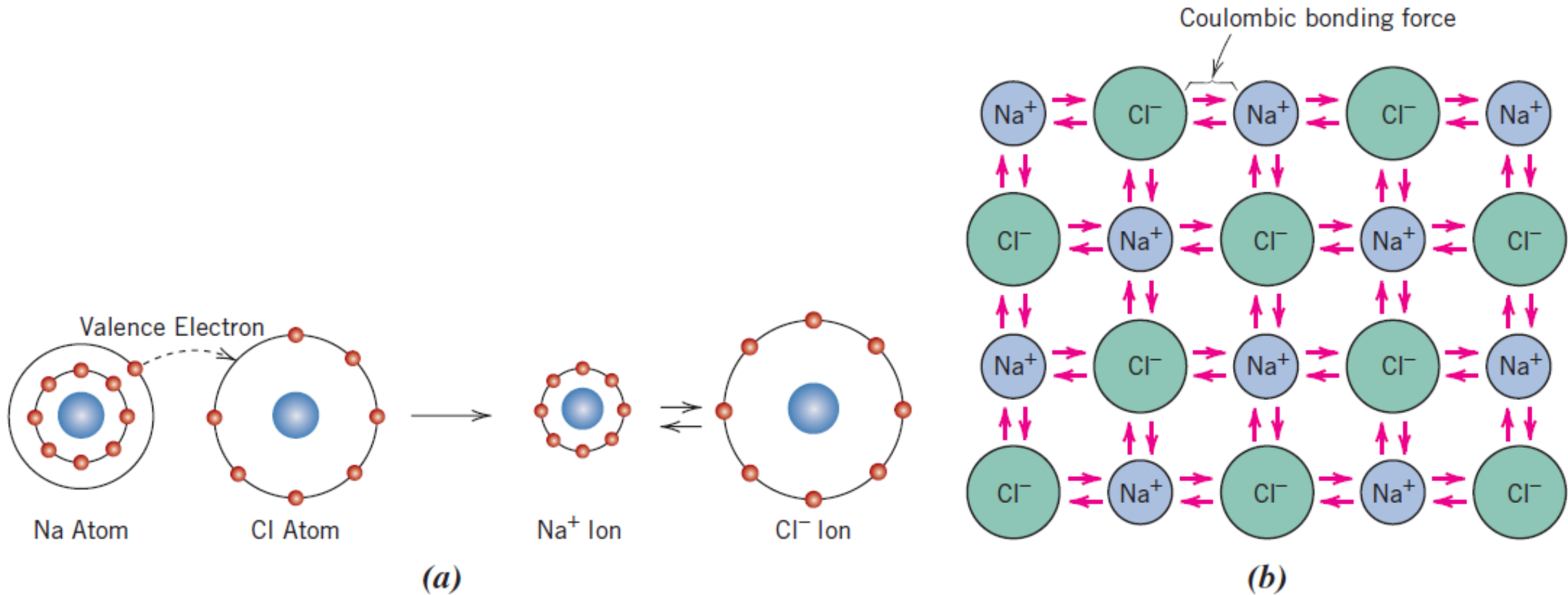


The strength of an ionic bond depends on the distance between ions (r) and the product of their charges (q). The mathematical expression for ionic bond strength is very similar to the expression for the Coulomb force:

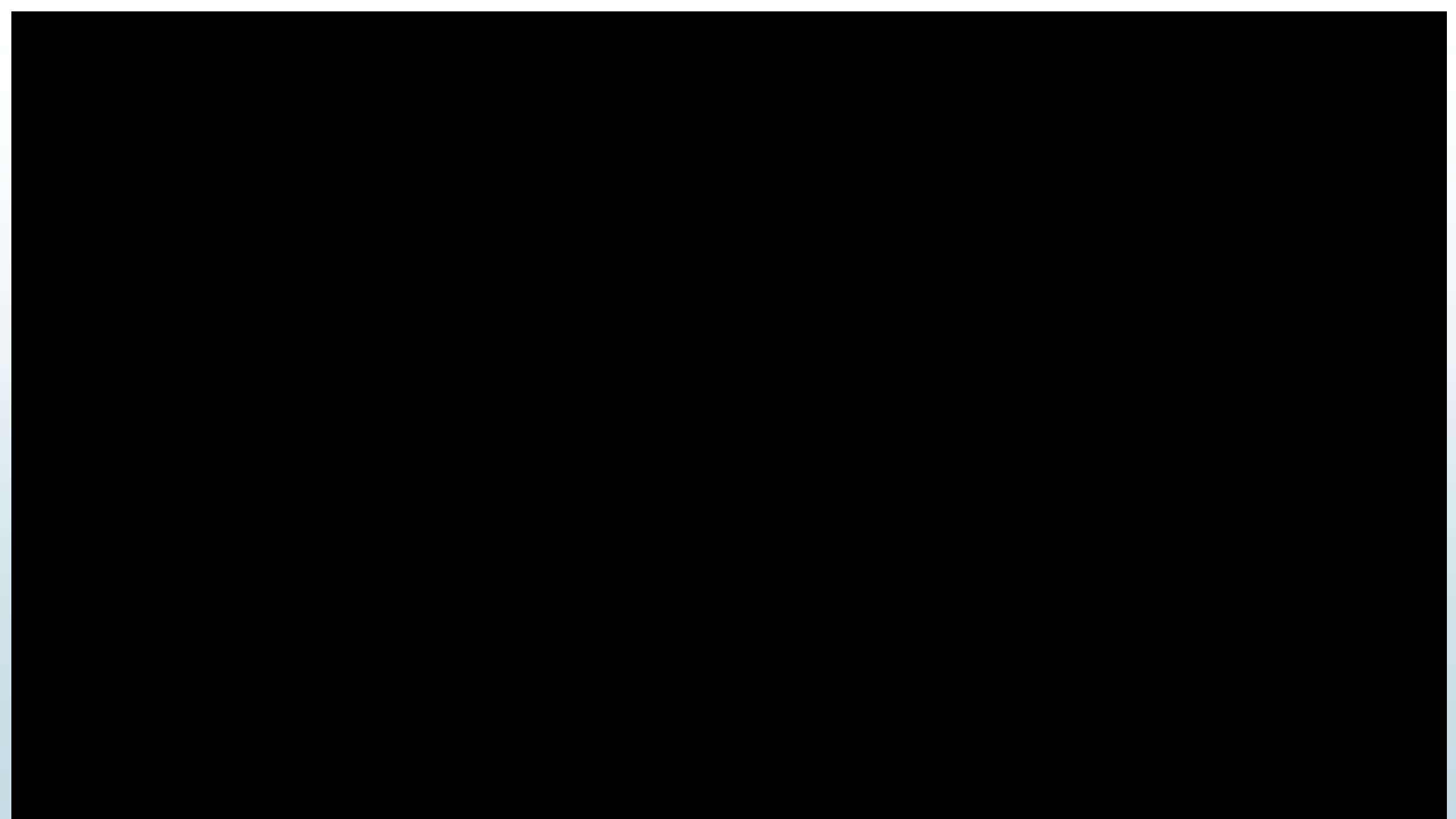
$$\text{Bond strength} \propto (q_1 q_2) / r$$

Properties of Ionic Materials

Ionically bonded materials usually have moderate hardness and fairly high melting points. They are generally soluble and are poor conductors of electricity because their constituent ions are fairly stable and neither lose nor gain electrons easily. They are usually highly symmetric and nondirectional (isotropic).

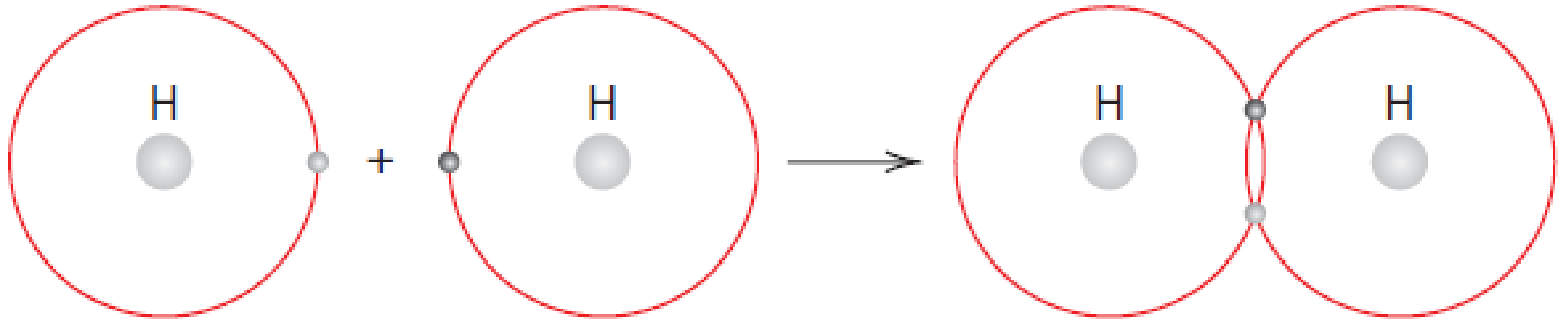


Schematic representations of (a) the formation of Na⁺ and Cl⁻ ions and (b) ionic bonding in sodium chloride (NaCl).



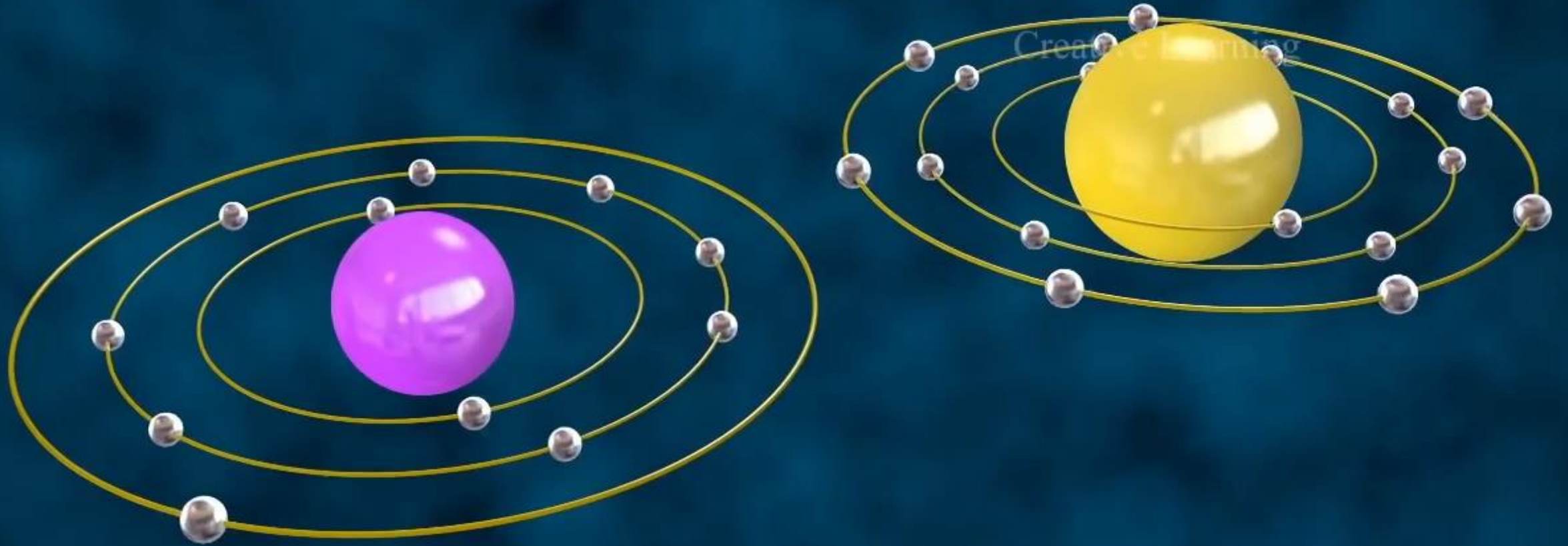
Covalent Bonding

found in materials whose atoms have small differences in electronegativity. For these materials, stable electron configurations are assumed by the sharing of electrons between adjacent atoms. Two covalently bonded atoms will each contribute at least one electron to the bond, and the shared electrons may be considered to belong to both atoms. Covalent bonding is schematically illustrated in Figure 2 for a molecule of hydrogen (H_2). The hydrogen atom has a single $1s$ electron. Each of the atoms can acquire a helium electron configuration (two $1s$ valence electrons) when they share their single electron. Furthermore, there is an overlapping of electron orbitals in the region between the two bonding atoms.




Schematic representation of covalent bonding in a molecule of hydrogen

Ionic Bonding



The bonding which results due to the complete transfer of valence electrons between atoms is called ionic bonding.



Many nonmetallic elemental molecules (e.g., Cl_2 , F_2), as well as molecules containing dissimilar atoms, such as CH_4 , H_2O , HNO_3 , and HF , are covalently bonded. Furthermore, this type of bonding is found in elemental solids such as diamond (carbon), silicon, and germanium and other solid compounds composed of elements that are located on the right side of the periodic table, such as gallium arsenide (GaAs), indium antimonide (InSb), and silicon carbide (SiC).

Properties of Covalent Materials

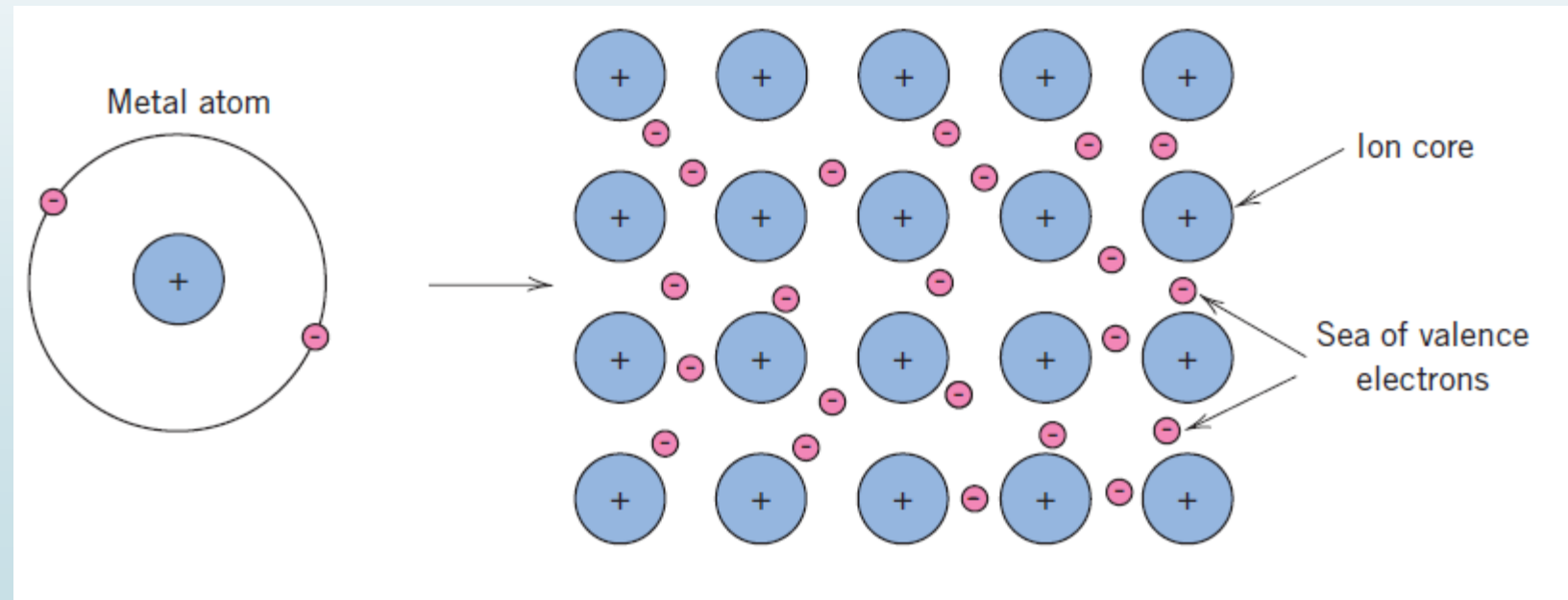
Covalent bonds are the **strongest of the chemical bonds**. Covalently bonded materials generally **have very high melting points** and **are generally insoluble**.

They typically do not conduct electricity as solids or in solution. Because covalent bonds are highly directional, covalently bonded materials tend to have less symmetry than ionically bonded materials.

Metallic Bonding

Metallic bonding, the final primary bonding type, is found in metals and their alloys. A relatively simple model has been proposed that very nearly approximates the bonding scheme. **With this model, these valence electrons are not bound to any particular atom in the solid and are more or less free to drift throughout the entire metal.** They may be thought of as belonging to the metal as a whole, or forming a “**sea of electrons**” or an “**electron cloud**.” The remaining nonvalence electrons and atomic nuclei form what are called *ion cores*, which possess a net positive charge equal in magnitude to the total valence electron charge per atom

Figure 3 illustrates metallic bonding. The free electrons shield the positively charged ion cores from the mutually repulsive electrostatic forces that they would otherwise exert upon one another; consequently, the metallic bond is nondirectional in character. In addition, these free electrons act as a “glue” to hold the ion cores together.



Schematic illustration of metallic bonding.




Properties of Metals

Metals are generally ductile and malleable. They are conductive and not very hard. They are highly symmetric because metallic bonds are nondirectional.

SECONDARY BONDING OR VAN DER WAALS BONDING

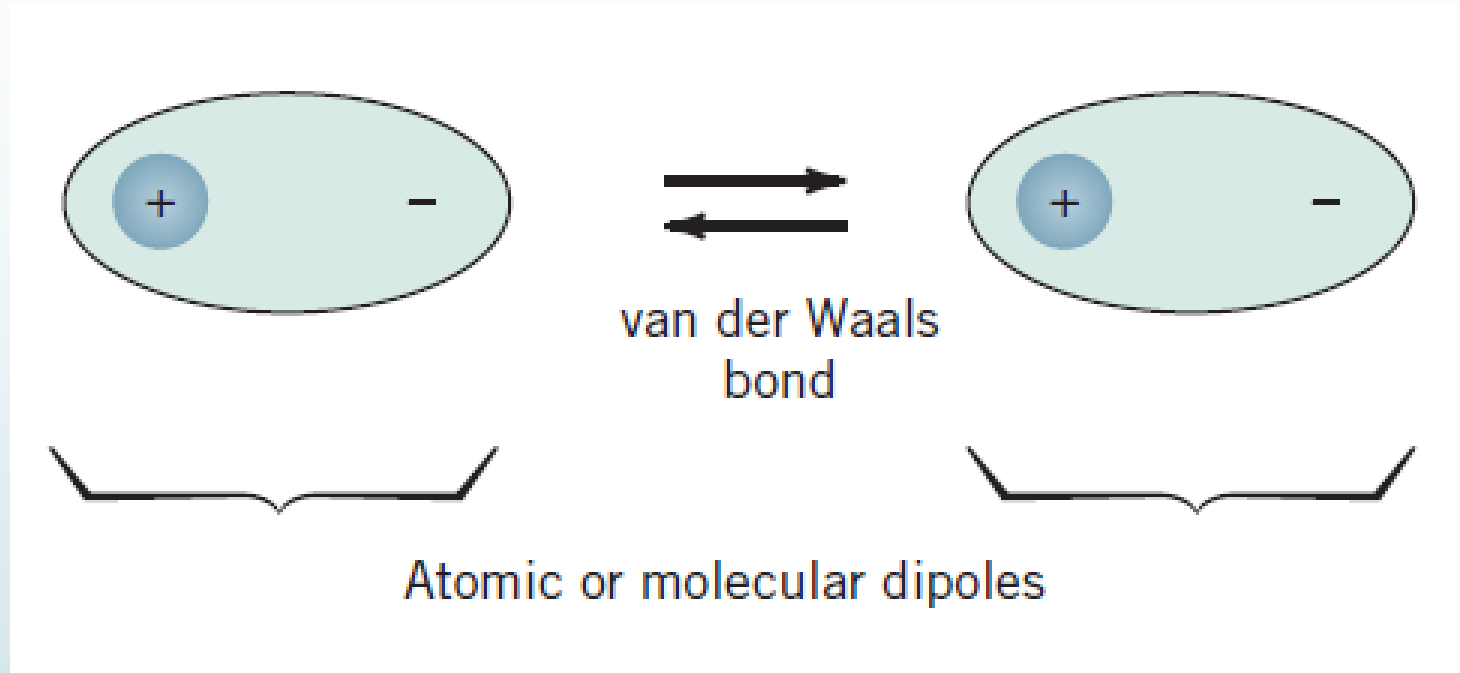
Secondary bonds, or **van der Waals** (physical) **bonds**, are weak in comparison to the primary or chemical bonds; bonding energies range between about 4 and 30 kJ/mol. Secondary bonding exists between virtually all atoms or molecules, but its presence may be obscured if any of the three primary bonding types is present. Secondary bonding is evidenced for the **inert gases**, which have stable electron structures. In addition, secondary bonds are possible between atoms or groups of atoms, which themselves are joined together by primary ionic or covalent bonds.



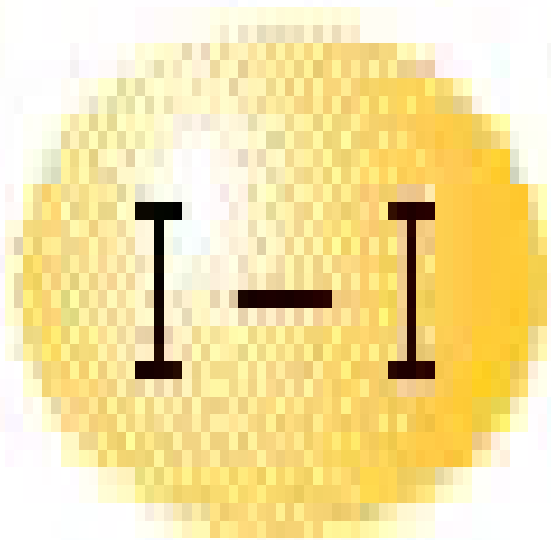
Secondary bonding forces arise from atomic or molecular **dipoles**. In essence, an electric dipole exists whenever there is some separation of positive and negative portions of an atom or molecule. The bonding results from the coulombic attraction between the positive end of one dipole and the negative region of an adjacent one, as indicated in Figure 4. Dipole interactions occur between induced dipoles, between induced dipoles and polar molecules, and between polar molecules. **Hydrogen bonding**, a special type of secondary bonding, is found to exist between some molecules that have hydrogen as one of the constituents.

<i>Substance</i>	<i>Bonding Energy (kJ/mol)</i>	<i>Melting Temperature (°C)</i>
Ionic		
NaCl	640	801
LiF	850	848
MgO	1000	2800
CaF ₂	1548	1418
Covalent		
Cl ₂	121	−102
Si	450	1410
InSb	523	942
C (diamond)	713	>3550
SiC	1230	2830
Metallic		
Hg	62	−39
Al	330	660
Ag	285	962
W	850	3414
van der Waals^a		
Ar	7.7	−189 (@ 69 kPa)
Kr	11.7	−158 (@ 73.2 kPa)
CH ₄	18	−182
Cl ₂	31	−101
Hydrogen^a		
HF	29	−83
NH ₃	35	−78
H ₂ O	51	0

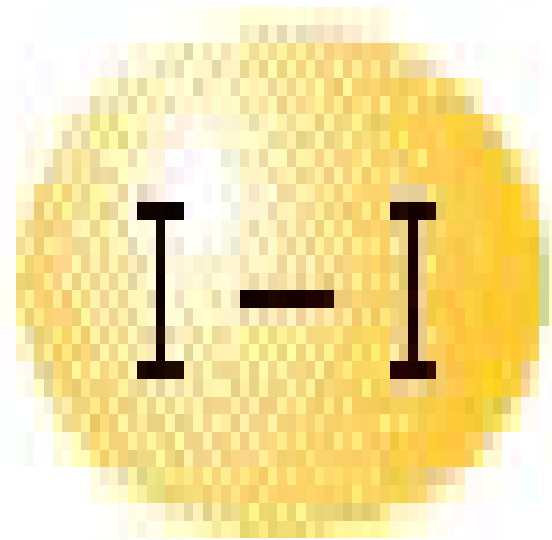
^aValues for van der Waals and hydrogen bonds are energies *between* molecules or atoms (*intermolecular*), not between atoms within a molecule (*intramolecular*).



Schematic illustration of van der Waals bonding between two dipoles.



No Dipole



No Dipole



Why do atoms combine?

Atoms combine together to lose their energy. This would make them stable.

- **Why do certain atoms combine while others do not?**

This is mainly because a compound forms only when there is an attractive force leading to the lowering of energy. On the other hand, in case of a repulsive force, we find an increase in overall energy of the system. Thus, we do not see the formation of any compounds.

Thank you for listening