The 4 Quantum Numbers

Principle Quantum Number	Secondary Quantum Number	Magnetic Quantum Number	Spin Quantum Number
n	l	m_l	m_s

Electron Configurations

Electrons fill orbitals from **lowest to highest** energy. Therefore, orbital 1s fills before 2s and 2p. However, an orbit does not necessarily fill completely before the next begins.

Types of Bonds

IONIC

- metals give electrons to non-metals
- metals form cations (**+**)
- non-metals form anions (**-**)
- this gives both atoms a stable electron configuration
- the energy level of each atom is decreased

If attraction outweighs repulsion, then a bond will form

Characteristics
- conductive in the dissolved or molten state
- solid, hard, brittle
- **high** melting point, **low** boiling point

COVALENT BONDING

1. Non-polar ▸ equal sharing of electrons for bonds
2. Polar ▸ unequal sharing of electrons, atom with higher ▲ EN is slightly **+**, lower ▲ EN is slightly -
3. Coordinate Covalent Bonds ▸ - both electrons forming the bond come from the same atom

Characteristics
- generally low boiling points
- solid, liquid, gas
- do not conduct electricity
- dull
- don't dissolve in water

Intermolecular Forces

Intermolecular Forces (cont)

3. **Dipole-Dipole➞** the attraction between oppositely charged dipoles of 2 polar molecules
- strength depends on the polarity of the molecule (more polar=stronger dipole force)
- H-bonding is a special type which is the strongest (5% of covalent bond strength)
- H bonded to N, O, F
- a lone pair of electrons must be on the neighbouring molecule for the H to bond with
- strength depends on the number of H bonds

4. **Dipole-Induced Dipole➞**- nonpolar molecule forced into polarity

4. **Induced Dipole-Induced Dipole➞** a.k.a. London Dispersion Forces
- the random motion of electrons creates a temporary dipole in one nonpolar molecule. This induces polarity in the neighbouring molecule. Strength depends on # of electrons (and protons) in a molecule.

Types of Solids

Metallic Crystals (Metallic Bonding)	- valence electrons from a mobile sea of electrons which comprise the metallic bond - high melting and boiling points
Ionic Crystals (Ionic Bonding)	- attraction of charged ions for one another. Lattice energy is a measure of ionic strength - high melting and boiling points
Covalent Crystals (Network Covalent Bonding)	- network solids are extremely hard compounds with very high melting and boiling points due to their endless 3-D network of covalent bonds

Types of Solids (cont)

Molecular Crystals a)H-bonding b)LDF c)Dipole-Dipole Forces	a) H-bonds are weaker than covalent bonds, but stronger than b) or c) below b) universal force of attraction between instantaneous dipoles. These forces are weak for small, low-molecular weight molecules, but large for heavy, long/highly polarizable molecules. They are stronger than c) below c) these forces act between **polar** molecules. They are much weaker than H-bonding
Atomic Crystals (Dispersion Forces)	- see section b) above

Physical properties depend on these forces. The **stronger** the forces between particles,
- the 🡅the melting and boiling point
- the 🡇the vapour pressure
- the 🡅 the viscosity
- the **greater** the surface tension

An attraction holding neighbouring molecules or ions together. **These are not bonds**

1. **Ion-Ion➞** whole charges attract

2. **Ion-Dipole➞** an ion is attracted to a polar molecule. The cation is attracted to the slightly negative portion of polar molecules and the anion to the slightly positive end

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