Masters for transparencies

5.1
Matter and materials

KNOWLEDGE AREA: MATTER AND MATERIALS

- Matter and materials
  - Atomic bonds
  - Intermolecular forces
  - Ideal gases and thermal properties

UNIT 1 ATOMIC BONDS: MOLECULAR STRUCTURE

- Atomic bonds
  - Chemical bonds
  - Bonding models
  - Valence electrons
  - Valency
  - Lewis structure
  - Formation of bonds
  - Single covalent bonds
  - Molecules with multiple bonds
  - Dative covalent bond
  - VSEPR shapes
  - Electronegativity
  - Electrical potential energy changes in the formation of a molecule
  - Bond energy and length

1.1 Chemical bonds

A **chemical bond** occurs when atoms bond together to form a new substance with new properties. In doing so it has a noble gas electron structure and a lower potential energy.

Quick facts

Chemical bonds occur to increase the stability of a substance, by decreasing the potential energy and increasing the entropy.

1.2 Bonding models

Covalent bond

- Bond between non-metals
- Electrons in half-filled orbitals overlap and are shared.
- Smallest particle is a molecule.
- Atoms must have half-filled orbitals.
- The half-filled orbitals overlap to form a new filled orbital.
- The atom's electronegativity must be the same or the difference must be less than 1,9.
- Polar or non-polar bonds form.

Non-polar bonds: atoms attract the shared pair of electrons equally.
Example:

$$O + O \rightarrow O_2$$

Polar bond: one atom has a greater pull on the shared pair than the other.
Example:

$$H_2 + C\ell_2 \rightarrow 2\overset{\delta+}{H}\overset{\delta-}{C\ell}$$

A covalent bond is where two non-metal atoms in a molecule share an electron pair during the overlapping of orbitals.

A non-polar covalent bond is where the shared electron pair is distributed evenly between the two atoms in the region where the orbitals overlap.

A polar covalent bond is where the shared electron pair is distributed unevenly between the two atoms in the region where the orbitals overlap.

Ionic bond

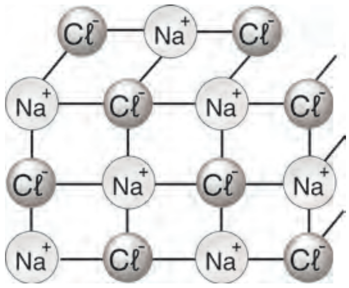
- Bond between metals and non-metals
- Electron transfer occurs.
- Positive ions (cations) and negative ions (anions) attract each other with strong electrostatic or Coulombic forces.
- Atoms' electronegativity must differ by more than 2,1.
- One atom must have a low ionisation energy so that it gives electrons away easily.
- Metals tend to donate electrons and become positive ions (cations).
- One of the atoms must have a high electron affinity to accept electrons.
- Smallest particle is an ion.

- Examples of ionic bonds:

  $Na \rightarrow Na^+ + e^-$

  $C\ell + e^- \rightarrow C\ell^-$

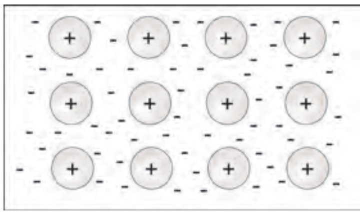
  $Na^+ + C\ell^- \rightarrow Na^+C\ell^-$



Ionic bonds take place between metal atoms and non-metal atoms when electrons of a metal atom are transferred to the non-metal atom. The ions that form are attracted to each other with strong electrostatic forces.

Metallic bond

- Bond within metals
- Crystal lattice of positive atomic core with a sea of delocalized electrons.
- Closely-packed crystal lattice.
- Smallest particle is a positive core ion.
- For a metallic bond to form:
  - Low ionisation energy
  - Atoms must have empty valence orbitals



Metallic bonds are when metal atoms of a metal bind through the attraction force between delocalised electrons and the crystal lattice of a positively charged atomic core.

1.3 Valence electrons

Electrons found in the outermost energy level.
Corresponds to the group number.
Valence electrons of some elements

Element	Group	Period	Valence electrons	Number of valence electrons
Na	1 (I)	3	$3s^1$	1
Mg	2 (II)	3	$3s^2$	2
N	15 (V)	2	$2s^22p^3$	5
O	16 (VI)	2	$2s^22p^4$	6
Cℓ	17 (VII)	3	$3s^23p^5$	7

To avoid confusion the groups on the Periodic Table are indicated with numbers 1 – 18. The old numbers (Roman numbers) are indicated in brackets.

1.4 Valency

Number of electrons that an atom will donate, accept or share.

1.5 Lewis structure

Examples of Lewis diagrams for period 2

Group 1 (I)	Group 2 (II)	Group 13 (III)	Group 14 (IV)	Group 15 (V)	Group 16 (VI)	Group 17 (VII)	Group 18 (VIII)
Li	Be	B	C	N	O	F	Ne

Covalent bonds
Example:

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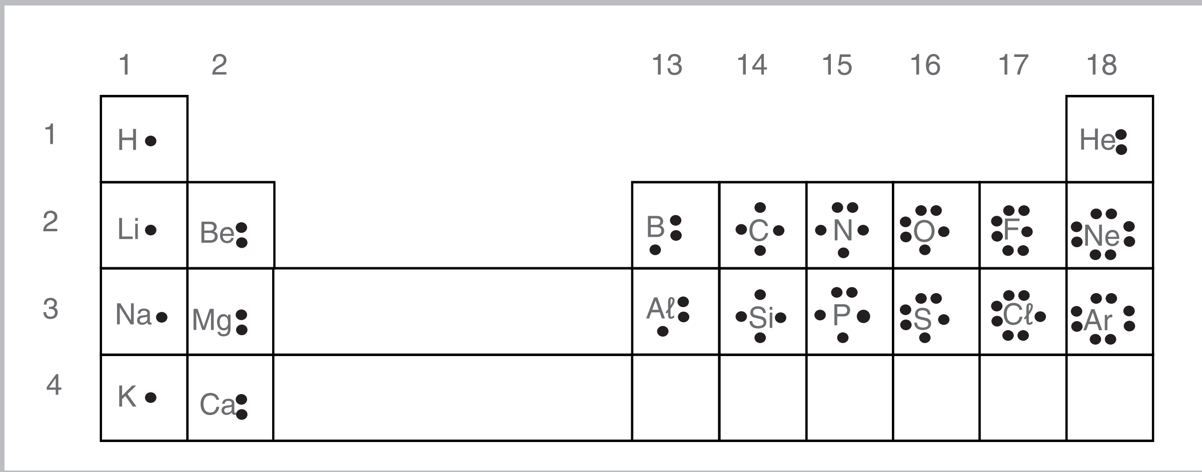
1. Draw the Lewis notation for the first twenty elements. Write the correct symbol in the block and then draw the notation.



2. If we look at the first twenty elements on the Periodic Table, how many bonds can occur in the following groups?

Group 1	**1**	Group 15	**3**
Group 2	**2**	Group 16	**2**
Group 13	**3**	Group 17	**1**
Group 14	**4**	Group 18	**0**

1.6 Formation of bonds

A	Different atoms, each containing an unpaired electron.	Sharing of electrons.	Simple molecule; covalent bond	E.g. HCℓ; HBr; HI; CH_4
B	The same atoms, each containing an unpaired electron.	Sharing of electrons.	Simple molecule; covalent bond	E.g. H_2; $Cℓ_2$; I_2
C	Different atoms containing only paired electrons (lone pairs).	No sharing of electrons.	No bond	E.g. He and Ar
D	The same atoms containing more than one unpaired valence electron.	More than one shared pair of electrons.	Multiple covalent bonds	E.g. O_2 and N_2
E	Different atoms containing more than one unpaired valence electron.	More than one shared pair of electrons.	Multiple covalent bonds	E.g. HCN and CO_2