4.2. Covalent bonding	
Nature of science: Looking for trends and discrepancies—compounds containing non-metals have different properties than compounds that contain non-metals and metals. (2.5) Use theories to explain natural phenomena—Lewis introduced a class of compounds which share electrons. Pauling used the idea of electronegativity to explain unequal sharing of electrons. (2.2)	
Understandings: • A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei. • Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively. • Bond length decreases and bond strength increases as the number of shared electrons increases. • Bond polarity results from the difference in electronegativities of the bonded atoms. **Applications and skills:** • Deduction of the polar nature of a covalent bond from electronegativity values. **Guidance:** • Bond polarity can be shown either with partial charges, dipoles or vectors. • Electronegativity values are given in the data booklet in section 8.	**Utilization:** • Microwaves—cooking with polar molecules. Syllabus and cross-curricular links: Topic 10.1—organic molecules **Aims:** • **Aim 3**: Use naming conventions to name covalently bonded compounds.

UNIT 4.2 – COVALENT BONDING

COVALENT BONDING VS IONIC BONDING

Ionic Bonding	Covalent Bonding
- Atoms lose or gain electrons in order to attain a **noble gas configuration**. - Metal and non-mental	- Atoms share electrons in order to attain a **noble gas configuration**. - Non-metal and non-metal

Ionic bonding	Covalent bonding
Formed between a cation (usually metal) and an anion (usually non-metal). Some cations (such as NH_4^+) can be comprised of non-metals and some anions (such as MnO_4^-) can contain metals.	Usually formed between non-metals.
Formed from atoms either losing electrons (process of oxidation) or gaining electrons (process of reduction) in order to attain a noble gas electron configuration.	Formed from atoms sharing electrons with each other in order to attain a noble gas electron configuration.
Electrostatic attraction between oppositely charged ions, that is, a cation (positive ion) and an anion (negative ion).	Electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
Ionic compounds have lattice structures.	Covalent compounds consist of molecules.*
Ionic compounds have higher melting points and boiling points.	Covalent compounds have lower melting points and boiling points.
Ionic compounds have low volatilities.	Covalent compounds may be volatile.
Ionic compounds tend to be soluble in water.	Covalent compounds typically are insoluble in water.
Ionic compounds conduct electricity because ions are free to move in the molten state. They do not conduct electricity when solid, however, as the ions are not free to move.	Covalent compounds do not conduct electricity because no ions are present to carry the charge.

▲ Table 2 Differences between ionic and covalent bonding

*We shall discuss covalent network structures that involve lattices later.

LEWIS SYMBOLS AND STRUCTURES

LEWIS SYMBOLS

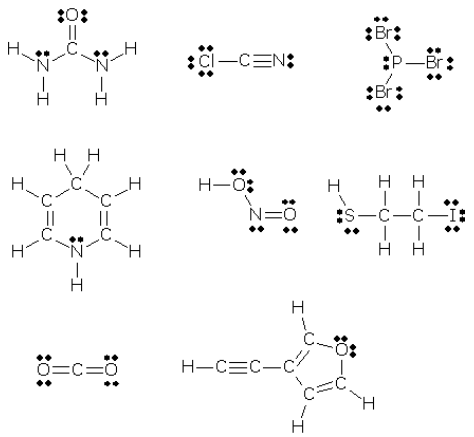
A simple and convenient method of representing the valence electrons in an element.

TABLE 8.1 Lewis Symbols

Element	Electron Configuration	Lewis Symbol
Li	$[He]2s^1$	Li·
Be	$[He]2s^2$	·Be·
B	$[He]2s^22p^1$	·$\dot{B}$·
C	$[He]2s^22p^2$	·$\dot{\underset{\cdot}{C}}$·
N	$[He]2s^22p^3$	·$\dot{\underset{\cdot}{N}}$:
O	$[He]2s^22p^4$	:$\dot{\underset{\cdot}{O}}$:
F	$[He]2s^22p^5$	·$\ddot{\underset{\cdot}{F}}$:
Ne	$[He]2s^22p^6$	:$\ddot{\underset{\cdot}{Ne}}$:

LEWIS STRUCTURES

A representation of elements in covalent compounds, by using dots and line to represent bonds and lone pairs.



Double and triple bond and represented by multiple lines, and mean that there are 2/3 electrons being shared by that particular element.

BOND STRENGTH AND BOND LENGTH

BOND STRENGTH

Highest	Triple bond
Medium	Double bond
Lowest	Single bond

BOND LENGTH

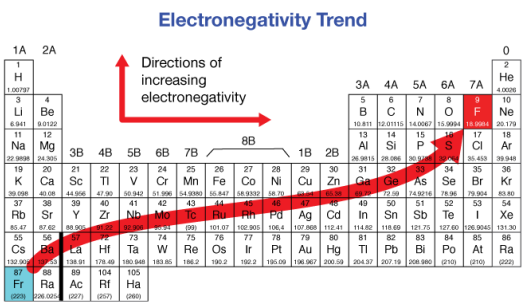
Longest	Single Bond
Medium	Double bond
Shortest	Triple bond

ELECTRONEGATIVITY AND POLARITY

Electronegativity refers to how willing an element is to bond with another element.

An atoms electron pulling power

Electronegativity is highest in the top right corner of the periodic table.



Highest at the top of a group

Valence shell is closer to positively charged nucleus, and hence harder to pull electrons away.

Highest to the right

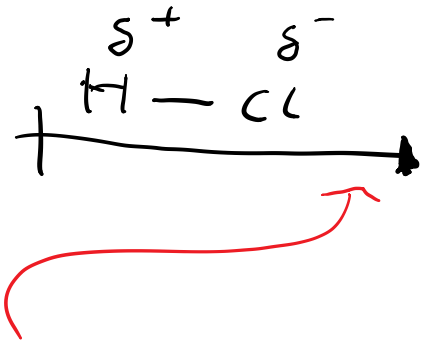
Nuclear charge increases, and so too does the attraction between the nucleus and the valence electrons.

The greater the difference in the electronegativity's of the atoms in a compound, the more uneven the sharing of electrons will be.

Atoms do not share a pair of electrons equally in a covalent bond. When the difference between electronegativity values is between 0 and 1.8, these atoms will most likely form a **polar** covalent bond.

POLAR COVALENT BONDS

If electrons are shared unevenly in a covalent bond, the bond is said to be **polar covalent bond**, or a **permanent dipole**.



- Has a greater electronegativity
- Direction that the electrons are moving towards.

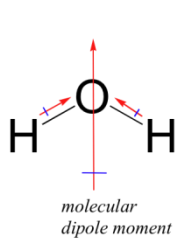
POLAR MOLECULES

If a polar bond occurs in a **diatomic** molecule (with only one covalent bond), one part of the molecule will be more negative than another, due to having a larger share of the bonding electrons.
This is described as a **polar molecule.**

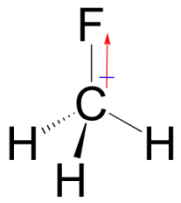
Such us the "H – CL" example above

When there is more than one covalent bond in a molecule, the shape of the molecule must be considered. Permanent dipoles can cancel each other out, meaning that a molecule can have polar bonds, but overall be non-polar.

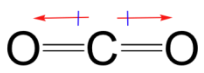
You can look at polarity arrows like vectors, and add them together



Polar



Polar



Non – Polar

VSEPR theory has to be taken into account