Worksheet 4.1

Chapter 4: Bonding – glossary

Allotrope Different forms of an element in the same physical state.

Anion A negatively charged ion. It is attracted to the anode during electrolysis.

Aromatic Structure characteristic of benzene, C₆H₆, and its derivatives. The molecules contain a delocalized ring of pi electrons.

Cation A positively charged ion. It is attracted to the cathode during electrolysis.

Coordination number The number of ions that surround an ion of the opposite charge in an ionic compound.

Covalent bond The electrostatic attraction between a pair of electrons and positively charged nuclei.

Dative bond This forms when both the electrons in the shared pair originate from one of the bonded atoms.

Delocalization The ability of pi bonding electrons to extend over two or three bonding positions within a molecule. Valence electrons in metal atoms are also delocalized within the metallic bond.

Diamond An allotrope of carbon in which each carbon atom is sp³ hybridized and tetrahedrally bonded to four others.

Dipole A pair of separated opposite electric charges.

Dipole–dipole attraction An intermolecular force between opposite ends of permanent dipoles in adjacent molecules.

Double bond A covalent bond that forms when two pairs of electrons are shared between the bonded atoms. It consists of one sigma and one pi bond.

Ductile The property of being able to be drawn out into threads. It is typical of metals.

Electronegativity The ability of an atom to attract the pair of shared electrons in a covalent bond. The difference in the electronegativity of the bonded atoms determines the polarity of the bond.

Expanded octet When the central atom in a molecule has more than 8 valence electrons. This occurs only with elements from the 3rd Period and beyond, e.g. PCl₅.

Fullerene A group of cage-like molecules built from hexagonal and pentagonal arrangements of carbon atoms. The most common example is buckminsterfullerene C₆₀.

Graphite An allotrope of carbon characterized by layers of carbon atoms arranged in hexagonal sheets. Each carbon atom is sp² hybridized.

Hybridization The process by which atomic orbitals within an atom mix to produce hybrid orbitals of intermediate energy. The hybrid atomic orbitals take part in bonding.
Hydrogen bond  A stronger form of dipole-dipole attraction that occurs when hydrogen is directly bonded to an atom of fluorine, nitrogen or oxygen.

Incomplete octet  When the central atom has fewer than 8 valence electrons, e.g. BeCl₂.

Induced dipole  The creation of a dipole in a molecule as the result of a dipole in an adjacent molecule.

Instantaneous dipole  Also known as temporary dipole. This forms when the electron density in an atom is unevenly distributed around the atom at a moment in time. It may cause an induced dipole in a neighbouring molecule.

Ion  A charged particle. It forms when an atom loses or gains one or more electron (s).

Ionic bond  The force of electrostatic attraction between oppositely charged ions.

Ionic compound  A compound which forms between a metal and a non-metal element by the transfer of electrons from the metal to the non-metal.

Lattice enthalpy  A measure of the strength of attraction between the ions in the lattice. It is the amount of energy required to separate one mole of ionic compound into isolated gaseous ions under standard conditions.

Lattice structure  A regularly repeating three-dimensional crystalline structure.

Lewis structure  A representation of the outer (valence) electrons of all the atoms in a molecule. It is sometimes called an electron dot diagram.

Lone pairs  Also known as non-bonding pairs. These are pairs of electrons in the valence shell of an atom which do not take part in bonding within the molecule.

Malleable  The property of being able to be shaped under pressure. It is typical of metals.

Metallic bond  The attraction between metal atoms as a result of the delocalized electrons in the lattice of cations.

Metals  Elements that form positive ions. They have a small number (one, two or three) of electrons in their outer shell.

Molecule  A group of atoms held together by covalent bonds.

Nanotubes  Fullerenes such as C70 based on a more cylindrical than spherical structure. They typically have a diameter of only a few nanometers.

Non-bonding pairs  Also known as lone pairs. These are pairs of electrons in the valence shell of an atom which do not take part in bonding within the molecule.

Non-metals  Elements that have a tendency to gain electrons and so form negative ions or covalent bonds.

Octet rule  The fact that most atoms (except H and He) bond together to achieve a stable octet (8) of valence electrons.
Pauling scale  A scale used to compare the electronegativity of different atoms. The most electronegative element (fluorine) is assigned a value of 4.

Pi (\(\pi\)) bond  A covalent bond in which the electron density is concentrated in two regions: above and below the bond axis. It forms by the lateral overlap of two unhybridized p orbitals, and forms alongside a \(\sigma\) bond in a double bond.

Polar bond  Forms when the shared electrons of the covalent bond are unequally shared. They are attracted more strongly by one of the bonded atoms than the other, due to a difference in the electronegativity of the bonded atoms.

Polyatomic ion  A group of covalently bonded atoms carrying a charge, due to the net loss or gain of one or more electrons.

Resonance structures  This occurs when more than one valid Lewis structure can be drawn for a molecule due to the delocalization of electrons. The true structure is known as a resonance hybrid.

Sigma (\(\sigma\)) bond  A covalent bond in which the electron density is concentrated along the bond axis. All single covalent bonds are \(\sigma\) bonds. They form by the overlap of s orbitals, p orbitals end-on, or hybridized orbitals.

Transition element  An element which has an incomplete d-subshell. Transition elements typically are able to form more than one stable ion.

Triple bond  A covalent bond that forms when three pairs of electrons are shared between the bonded atoms. It consists of one sigma and two pi bonds.

Van der Waals’ forces  Weak intermolecular forces, arising from the induced dipoles between atoms or molecules as a result of their temporary dipoles. They are the only force between non-polar molecules and noble gases. Their strength increases with an increasing number of electrons in the molecule.

VSEPR theory  Valence Shell Electron Pair Repulsion: the theory that enables the shape of a molecule to be predicted from positioning the charge centres around the central atom as far apart as possible.