

	Statement	Guidance
4.1.U1	Positive ions (cations) form by metals losing valence electrons.	
4.1.U2	Negative ions (anions) form by non-metals gaining electrons	Students should be familiar with the names of these polyatomic ions: NH_4^+ , OH^- , NO_3^- , HCO_3^- , CO_3^{2-} , SO_4^{2-} and PO_4^{3-}
4.1.U3	The number of electrons lost or gained is determined by the electron configuration of the atom	
4.1.U4	The ionic bond is due to electrostatic attraction between oppositely charged ions	
4.1.U5	Under normal conditions, ionic compounds are usually solids with lattice structures	
4.2.U6	A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei	
4.2.U7	Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively	
4.2.U8	Bond length decreases and bond strength increases as the number of shared electrons increases	
4.2.U9	Bond polarity results from the difference in electronegativities of the bonded atom	Bond polarity can be shown either with partial charges, dipoles or vectors
4.3.U10	Lewis (electron dot) structures show all the valence electrons in a covalently bonded species	
4.3.U11	The "octet rule" refers to the tendency of atoms to gain a valence shell with a total of 8 electrons.	The term "electron domain" should be used in place of "negative charge centre"
4.3.U12	Some atoms, like Be and B, might form stable compounds with incomplete octets of electrons.	
4.3.U13	Resonance structures occur when there is more than one possible position for a double bond in a molecule	
4.3.U14	Shapes of species are determined by the repulsion of electron pairs according to VSEPR theory	
4.3.U15	Carbon and silicon form giant covalent/network covalent structures	Allotropes of carbon (diamond, graphite, graphene, C_{60} buckminsterfullerene) and SiO_2 should be covered. Coordinate covalent bonds should be covered
4.4.U16	Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding	The term "London (dispersion) forces" refers to instantaneous induced dipole- induced dipole forces that exist between any atoms or groups of atoms and should be used for non-polar entities. The term "van der Waals" is an inclusive term, which includes dipole-dipole, dipole-induced dipole and London (dispersion) forces
4.4.U17	The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds	
4.5.U18	A metallic bond is the electrostatic attraction between a lattice of positive ions and delocalized electrons.	
4.5.U19	The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion.	
4.5.U20	Alloys usually contain more than one metal and have enhanced properties	
14.1.U21	Covalent bonds result from the overlap of atomic orbitals. A sigma bond (σ) is formed by the direct head-on/end-to-end overlap of atomic orbitals, resulting in electron density concentrated between the nuclei of the bonding atoms. A pi bond (π) is formed by the sideways overlap of atomic orbitals, resulting in electron density above and below the plane of the nuclei of the bonding atoms.	The linear combination of atomic orbitals to form molecular orbitals should be covered in the context of the formation of sigma (σ) and pi (π) bonds

14.1.U22	Formal charge (FC) can be used to decide which Lewis (electron dot) structure is preferred from several. The FC is the charge an atom would have if all atoms in the molecule had the same electronegativity. $FC = (\text{Number of valence electrons}) - \frac{1}{2}(\text{Number of bonding electrons}) - (\text{Number of non-bonding electrons})$. The Lewis (electron dot) structure with the atoms having FC values closest to zero is preferred.	
14.1.U23	Exceptions to the octet rule include some species having incomplete octets and expanded octets.	Molecular polarities of geometries corresponding to five and six electron domains should also be covered.
14.1.U24	Delocalization involves electrons that are shared by/between more than one pair in a molecule or ion as opposed to being localized between a pair of atoms.	
14.1.U25	Resonance involves using two or more Lewis (electron dot) structures to represent a particular molecule or ion. A resonance structure is one of two or more alternative Lewis (electron dot) structures for a molecule or ion that cannot be described fully with one Lewis (electron dot) structure alone	
14.2.U26	A hybrid orbital results from the mixing of different types of atomic orbitals on the same atom.	Students need only consider species with sp^3 , sp^2 and sp hybridization.