

IGC	Learning Outcomes	
8.1	<b>Describe characteristics of electromagnetic radiation.</b>	
8.1		Define electromagnetic radiation.
8.1		Distinguish among amplitude, wavelength, and frequency.
8.1		Describe the general trends in the electromagnetic spectrum (frequency, wavelength, and energy).
8.1		State the colors in the visible region along with energy, frequency, and wavelength trends.
8.1		Recall that light in the visible region has wavelengths from 400-750 nm.
8.1		Classify regions of the spectrum as higher or lower in energy than visible region.
8.1		Distinguish between constructive and destructive interference.
8.1		Describe diffraction in waves.
8.1	<b>Complete calculations related to electromagnetic radiation.</b>	
8.1		Recognize that $c$ is the speed of light.
8.1		Calculate among values of energy, frequency, and wavelength.
8.1		Relate energy to the number of photons.
8.1	<b>Describe the photoelectric effect.</b>	
8.1		Explain the concept of threshold frequency and how it affects the ejection of electrons.
8.1		Explain how intensity and wavelength affect the electrons emitted (or not) in the photoelectric effect.
8.1		Recognize that the photoelectric effect led to the understanding of the particle nature (photons) of light.
8.1		Calculate energy/frequency/wavelength associated with photoelectric effect.
8.1	<b>Describe the emission of electromagnetic radiation.</b>	
8.2		Distinguish between emission line spectra and the continuous spectrum of white light.
8.2		Relate the Bohr model to the emission spectrum of hydrogen.
8.2		Relate the energy of the photon emitted or absorbed to the energy change of the electron.
8.2		Complete calculations for hydrogen atom using the Rydberg equation.
8.1	<b>Recognize the wave-particle duality of matter.</b>	
8.1		Explain how the interference pattern from a beam of electrons supports that electrons behave as waves.
8.3		State the deBroglie formula and identify all variables.
8.3		Use the deBroglie formula in calculations.
8.3	<b>Describe Heisenberg's uncertainty principle</b>	
8.3		Explain the inversely proportional relationship between the uncertainty of position and the uncertainty of the velocity.
8.6	<b>Summarize the meaning and relevance of the quantum numbers.</b>	
8.6		Identify each of the four quantum numbers.
8.6		Define the meaning of each of the quantum numbers.
8.6		Give the possible values for each quantum number.
8.6		Identify sets of allowed/disallowed quantum numbers.
8.6		Pair angular momentum quantum numbers with the shape of orbitals.
8.6		Describe the shape of each orbital type (s, p, d, and f).
8.6		Describe the relationship between nodes and orbitals.
8.6		Given select values of quantum numbers, determine the number of orbitals or electrons with those values.
8.4		Define degenerate.
8.4		Rank the energy level of sublevels within a principal level (s, p, d, f)
8.6		Define the Pauli exclusion principle.
8.6		Explain how the Pauli exclusion principle affects the values of quantum numbers.
8.5	<b>Write the electron configuration of an element.</b>	
8.4		Describe Hund's rule and its effect on electron arrangement.

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8.4		Define aufbau principle.
8.4		Draw orbital diagrams of neutral atoms.
8.5		Write electron configuration of neutral atoms.
8.5, 9.1		Distinguish between valence and core electrons.
9.1		Determine the number of valence electrons in an atom.
8.5		Write electron configurations for transition metals (including exceptions for d4 and d9 elements).
8.4		Define diamagnetic and paramagnetic.
8.4		Determine if an element is diamagnetic or paramagnetic.
8.5	<b>Identify patterns in electron configurations based on arrangement in the periodic table.</b>	
8.5		Relate the electron configuration of a group of elements to their arrangement in the periodic table.
8.6		Write the quantum numbers for any electron in an atom based on its orbital diagram or electron configuration.
9.2	<b>Summarize periodic trends for effective nuclear charge and atomic radius.</b>	
9.2		Describe shielding of electrons by other electrons.
		Define effective nuclear charge.
9.2		Identify the trends in effective nuclear charge.
9.2		Define atomic radius for atoms.
9.2		Identify the trends in atomic radii.
9.2		Explain the trends in atomic radii and effective nuclear charge.
9.2		Estimate the value of the effective nuclear charge of an electron.
9.2	<b>Summarize periodic trends of ions and isoelectronic series.</b>	
9.2		Describe the size of an ion relative to its parent atom.
8.6		Define isoelectronic.
8.6		Identify isoelectronic species.
9.2		Rank isoelectronic species according to size.
9.3	<b>Summarize periodic trends for ionization energy.</b>	
9.3		Define ionization energy.
9.3		Write the reaction that represents the first ionization of an atom.
9.3		Describe trends in first ionization energies.
9.3		Describe trends in second and successive ionization energies.
9.3		Identify element based on sequence of ionization energies.
9.3	<b>Summarize periodic trends for electron affinity.</b>	
9.3		Define electron affinity.
9.3		Write the reaction that represents the process associated with the first electron affinity.
9.3		Describe "trends" in electron affinity values.
2.3	<b>Summarize periodic trends for metallic character.</b>	
2.6		Define metallic character.
2.6		Describe trends in metallic character.
9.4	<b>Describe formation of ionic compounds.</b>	
8.6		Write the electron configuration of cations and anions.
9.4, 10.1, 12.6		Distinguish among ionic, covalent, and metallic bonding.
10.3, 10.5	<b>Use electronegativity to describe properties of covalent bonds.</b>	
10.3		Explain electronegativity of an element.
10.3		Describe the trends for electronegativity of elements.
10.5		Relate electronegativity to bond polarity.
10.5		Categorize bonds as ionic, polar, or nonpolar.
10.2-10.4	<b>Draw Lewis structures of compounds.</b>	
10.1		Draw Lewis symbols of atoms.
10.2		Draw the Lewis structure of a binary ionic compound.

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10.2		Describe how electrons form single, double, or triple bonds.
10.2		Draw Lewis structures of molecules and polyatomic ions.
10.4		Name the three types of exceptions to the octet rule.
10.4		Draw Lewis structures of compounds that violate the octet rule.
10.3		Define resonance.
10.3		Identify when resonance is possible in a molecule or ion.
10.3		Define resonance hybrid.
10.3		Calculate formal charge of atoms in a molecule or ion.
10.3		Explain the significance of formal charge values.
10.3		Apply a formal charge explanation to determine the best Lewis structure for a compound.
10.6	<b>Characterize the strength of covalent bonds.</b>	
10.6		Define bond energy.
10.6		Discuss bond strength as a function of bond length.
10.6		Describe how the bond length changes from single to double to triple bond between atoms.
10.6		Calculate the unknown given bond energy values and/or the enthalpy of a reaction.
9.5	<b>Characterize the strength of ionic bonds.</b>	
9.5		Define lattice energy.
9.5		Describe the lattice energy trends related to ion size and charge.
9.5		Describe the steps of the Born-Haber cycle.
11.1	<b>Describe the three-dimensional shape of a molecule.</b>	
11.1		Describe valence shell electron pair repulsion (VSEPR) theory.
11.1		Recognize the electron group geometry of molecules.
11.1		Determine the electron group geometry based on the Lewis structure.
11.1		Determine the effect of lone pairs on the geometry of a molecule.
11.1		Name the possible electron group and molecular geometries.
11.1		Identify the electron group geometry and molecular geometry based on the number of <b>bonding and non-bonding groups in the Lewis structure of a molecule or ion.</b>
11.1		Determine the geometries for molecules with more than one central atom.
11.2	<b>Determine the polarity of a molecule.</b>	
11.2		Define dipole moment.
11.2		Recognize that polarity affects solubility of molecular compounds (like dissolves like).
11.3	<b>Describe valence bond theory in terms of orbital overlap to form bonds.</b>	
11.3	<b>Describe hybridization of atomic orbitals.</b>	
11.3		Pair hybridization schemes with the appropriate electron group geometry.
11.3		Compare the energy of the hybrid orbitals to the atomic orbitals from which they were formed.
11.3		Use the overlap of atomic and hybrid orbitals to explain the bonding in a molecule.
11.3		Define sigma and pi bonding according to the location of the electron density with respect to <b>the nuclei of the atoms in the bond.</b>
11.3		List the types of orbitals which form a sigma bond.
11.3		List the types of orbitals which form a pi bond.
11.3		Classify covalent bonds in a molecule or polyatomic ion as sigma or pi.
11.5	<b>Describe molecular orbital theory for homonuclear diatomic elements or ions.</b>	
11.5		Distinguish between bonding and antibonding orbitals.
11.5		Draw the MO diagram for homonuclear diatomics.
11.5		Calculate the bond order based on the MO diagram.
11.5		Relate the bond order to the stability of the bond.
11.5		Determine if a diatomic molecule is diamagnetic or paramagnetic based on the MO diagram.

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11.5		Relate the concept of delocalized molecular orbitals to the concept of resonance.
11.5		Describe the linear combination of atomic orbitals that form bonding and anti-bonding molecular orbitals.
11.5		Describe how destructive and constructive interference affects molecular orbitals.
11.5		Label the components of the molecular orbital diagram of H <sub>2</sub> .
7.1	<b>Identify and use units of pressure.</b>	
7.1		Memorize the common pressure units and the conversion factors between pressure units.
7.1		Convert between pressure units.
	<b>Describe the simple gas laws qualitatively and quantitatively.</b>	
7.2		Calculate values of pressure or volume using Boyle's Law.
7.3		Calculate values of volume or temperature using Charles' Law.
7.5		Calculate values of moles or volume using Avogadro's Law.
7.6		Use the ideal gas law equation to derive the simple gas laws.
	<b>Use the ideal gas law in calculations.</b>	
7.6		Calculate values of P, V, n or T if given the other three using the ideal gas law.
7.4		Complete calculations with the combined gas law
7.1		Define standard temperature and pressure (STP).
7.8		Combine density calculations and molar mass calculations with the ideal gas law to determine the density or molar mass of a gas.
	<b>Define and derive molar volume.</b>	
7.1		Use molar volume at STP as a conversion factor.
7.9		Use Avogadro's Law to do stoichiometry conversions in reactions involving gases.
7.7	<b>Use Dalton's Law of partial pressure to calculate total pressure of a gas if given individual gas pressures and vice versa.</b>	
7.7		State Dalton's law of partial pressure and define all variables.
7.7		Define mole fraction
7.7		Calculate the mole fraction of a substance in a mixture.
7.7		Calculate partial pressure in terms of mole fraction and total pressure.
7.11	<b>Describe the processes of diffusion and effusion of gases.</b>	
7.11		Define diffusion and effusion.
7.11		Compare relative speeds of molecules and rates of effusion as a function of molar mass.
7.10	<b>Use the kinetic molecular theory of a gas to explain the simple gas laws on the molecular level.</b>	
7.10		Retell the statements of the Kinetic Molecular Theory of a Gas.
7.10		Recognize the equation for kinetic energy ( $\frac{1}{2}mv^2$ ) and identify all variables.
7.11		Describe the qualitative relationship between variables in the rms velocity equation.
7.11		Describe the qualitative relationship between molar mass and average kinetic energy of gases.
7.11		Describe the qualitative relationship between temperature and average kinetic energy of gases.
7.11		Describe the relationship between rate of effusion and molar mass as explained by Graham's law.
7.12	<b>Distinguish between real and ideal gases.</b>	
7.12		Describe the conditions of pressure and temperature that distinguish a real gas from an ideal gas.
7.12		Explain qualitatively how the terms of the van der Waals equation account for the properties of a real gas.