Elements; Numbers of: There are 115 known elements, with 3 extras under dispute. In addition, there are 90 natural elements, numbers 1-92 with the exclusion of 43 Technetium and 61 Promethium.

Chemical Symbols; Naming of: Most elements have chemical symbols that resemble the name of the elements that they represent, such as oxygen having the chemical symbol O, however, there are several elements listed below in figure I that have their names derived from their Latin or in tungsten’s case German name.

|  |  |  |
| --- | --- | --- |
| Symbol | Latin Name | Name |
| Na | Natrium | Sodium |
| K | Kalium | Potassium |
| Fe | Ferrum | Iron |
| Cu | Cuprum | Copper |
| Ag | Argentum | Silver |
| Au | Aurum | Gold |
| Hg | Hydrargyrum | Mercury |
| Sn | Stannum | Tin |
| Sb | Stibium | Antimony |
| Pb | Plumbum | Lead |
| W | Wolfram | Tungsten |

Figure I-A table containing the symbols, Latin/Germanic names of elements, and their real names.

Amphoteric line: The line that runs diagonally from below boron down to the corner between Polonium and Astatine. The elements that run along this line are metalloids and to its left are metals (except for hydrogen) and to its right are non-metals.

Properties of Metals: Metals are shiny, malleable, ductile, have a positive charge and conduct electricity. In addition, their names, with few exceptions end with -ium.

Noble Gases: The column marked 18/VIII on the far right of the periodic table are called the noble gases and do not react chemically with any elements except in the rarest of circumstances, they do not react due to their full external electron shells.

Binary Compounds: Chemical compounds that contain only two elements, it may contain more than two atoms, however, there can only have two elements, therefore, H2O and NaCl are both examples of binary compounds. You can identify a compound as a binary compound by the -ide ending.

Rules for Naming Binary Compounds:

1. Name the metal first as it appears on the left of the periodic table.
2. Name the non-metal second, using an -ide ending.

Charge of Non-metals: Non-metals have a single negative charge, i.e. Oxygen can only have the negative charge O2-, but may also have multiple positive charges, i.e. O2+, O4+, O6+.

Ion: An element that has a positive or negative charge, but does not have a neutral charge.

Charge of Metals: Metals always have a positive charge.

Metals with more than One Oxidation State; Naming of: When using metals with more than one oxidation state in a compound they should be denoted in the form-

Metal, (oxidation state in roman numerals), Non-metal

Ternary Compounds: A chemical compound consisting of three or more elements.

Naming Binary Compounds with Two Non-metals:

1. Name the first non-metal without changing its name.
2. Name the second non-metal using an -ide ending.
3. Use Latin prefixes for each element to indicate quantity. Below in figure II is a table of these prefixes. In the case of the first element only being one you can drop the mono. Note that 2 is di- and not bi- as bi- is used to indicate hydrogen. Also note that 4 is tetra- and not quad-.

Figure II

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 |
| mono- | di- | tri- | tetra- | penta- | hexa- | hepta- | octa- | nona- | deca- |

Complex Inorganic Compounds: An ion that contains a metal and non-metal as well as a central atom.

Ligand: Attached ions or compounds in a complex ion.

|  |  |  |
| --- | --- | --- |
| List of Common Ions | | |
| F- | Fluoride | Fluoro |
| Cl- | Chloride | Chloro |
| Br- | Bromide | Bromo |
| I- | Iodide | Iodo |
| OH- | Hydroxide | Hydroxo |
| CN‑ | Cyanide | Cyano |
| H2O | Water | Aquo |
| NH3 | Ammonia | Ammine |

Converting to Ligand Names: When taking a name of something such as cyanide, to convert it to a ligand name change its suffix to -o. In the case of cyanide, you get cyano.

|  |  |
| --- | --- |
| Negative Complex Ion Name | |
| Chromium | Chromate |
| Cobalt | Cobalate |
| Copper | Cuprate |
| Gold | Aurate |
| Iron | Ferrate |
| Lead | Plumbate |
| Mercury | Mercurate |
| Nickel | Nickelate |
| Platinum | Platinate |
| Silver | Argentate |
| Tin | Stannate |
| Zinc | Zincate |

Figure III-A list that covers many of the common ions with their chemical symbol followed by charge in the left column, their common name in the second column and their ligand name in the third.

Figure IV-A list that covers the negative complex ion names for many elements and features the element name on the left and the element negative complex ion name on the right.

Positively Charge Ion Nomenclature:

1. Name a metal and place it at the ‘back’ of the name, use a roman numeral if necessary.
2. Name the ligand with a Latin prefix indicating quantity in the front
3. Add the work ion at the end of the name

Negatively Charged Ion Nomenclature:

1. Name the metal in its Latin form (if applicable) and add an -ate suffix, proceed to put it in the back and add a roman numeral is necessary.
2. Name the ligand with a Latin prefix indicating quantity and place it in the front
3. Add the word ion at the end of each name

Organic Compounds: Compounds that are carbon-based, as carbon can create an extraordinary number of bonds. They may include side-chains and may form rings of carbon atoms.

|  |  |
| --- | --- |
| Names for Main Carbon Chains | |
| 1 | meth- |
| 2 | eth- |
| 3 | prop- |
| 4 | but- |
| 5 | pent- |
| 6 | hex- |
| 7 | hept- |
| 8 | oct- |
| 9 | non- |
| 10 | dec- |

|  |  |
| --- | --- |
| Types of Organic Bonding | |
| all single bonding | -ane |
| at least one double bond | -ene |
| at least one triple bond | -yne |

Figure V: The endings of the names of organic compounds by types of bonding.

Figure VI: The beginning of the names of organic compounds based off the number of carbon atoms in the main carbon chain.

Alkanes: Organic compounds that are all single bonded.

Alkenes: Organic compounds that have at least one double bond.

Alkynes: Organic compounds that have at least one triple bond.

Saturated: An organic compound that has all single bonds.

Unsaturated: An organic compound that has all double bonds.

Lines Between Carbon Atoms: The lines between carbon atoms are used to indicate the number of bonds that have, if they have no or one line such as C-C it is single bonded, if they have two lines then they are double bonded (C=C), and if they have three lines they are triple bonded C≡C.

Structural Isomers: Molecules which have the same chemical formula, but have different structures.

Nomenclature for Side Chains: In organic molecules, a side chain can be denoted by adding -yl to the end of the stem name.

Providing a Chemical Name given a Structural Formula or Vice Versa:

1. Name or draw the main chain first. When writing, the main chain is determined first, but written last. If you have a double or triple bond present, indicate it in the beginning of the name of the main chain with a number indicating its location, if there are multiple double or triple bonds in the chain indicate so with the numbers.
2. Name or draw the side chains. When writing, this is placed in front of the main chain name. The side chains are ordered by bond type going from triple bonds to double to single bonds.
3. Provide the address or location of the side chain, addresses are determined from the least length, i.e. if there are two possible ways to name the chain, use the one which starts with the smallest number.

Cycloalkanes: Several molecules that wrap themselves into a ring, the specific values are usually not shown when drawn as they are assumed to have a carbon and as many hydrogens as necessary to fill carbon’s four bonds. They may include a number of side chains and are named in the same way as normal hydrocarbons, with the exception of the added prefix cyclo- to the main chain.

|  |  |  |
| --- | --- | --- |
| Figure VI-SI Prefixes | | |
| Tera- | T | 12 |
| Giga- | G | 9 |
| Mega- | M | 6 |
| kilo- | k\* | 3 |
| -- | -- | 0 |
| deci- | d | -1 |
| centi- | c | -2 |
| milli- | m | -3 |
| micro- | µ | -6 |
| nano- | n | -9 |
| pico- | p | -12 |

Rules for Significant Digits:

1. All digits 1-9 are significant
2. Zero is significant if it is between significant digits
3. Leading zeros are never significant
4. Trailing zeros in the presence of a decimal are significant
5. Trailing zeros in the absence of a decimal are not significant
6. In exponential notation, the number of significant digits is the number of digits preceding the exponent.

Avogadro’s Number: 6.022

Solute: Substance being dissolved.

Solvent: Substance dissolving another substance.

Molarity: Moles per litre of solution, can be expressed as a mass divided by molecular weight, over litres or (m/MW)(L)-1. Molarity may be notated mol/L or simply M.

Hydrated Crystals: Crystals that have trapped water within them when formed. Their chemical formula is given in the form A • xH2O with A denoting the chemical formula of the crystal, • denoting physical as opposed to chemical relationship and x being the multiplier of H2O.

Diatomic Molecules: Molecules of elements that bond with themselves, these elements are very reactive and are hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine.

Synthesis Equation: A chemical equation of the form A + B → C.

Decomposition Equation: A chemical equation of the form A → B + C.

Single Displacement Equation: A chemical equation of the form A + BC → AC + B.

Double Displacement Equation: A chemical equation of the form AB + CD → AD + BC.

Combustion Equation: A chemical equation of the form A + O2 → AO.

Enthalpy(▲H°): Heat content, measured in joules.

Endothermic Reaction: A reaction that requires energy, the reactants start with less energy than the products, ▲H° > 0.

Exothermic Reaction: A reaction that gives off energy, the reactants start with more energy than the products, ▲H° < 0.

Enthalpy of a Reaction: The difference between the product enthalpy and the reactant enthalpy at constant pressure.

Enthalpy of Formation: The enthalpy change, when one mole of a compound is formed from its constituent elements.

State Function: A function defined for a system relating several state variables/quantities that only depends on the current equilibrium. They do not depend on the path by which the system arrives at its present value.

Hess’s Law: The value ▲H for a reaction is the same whether it occurred in one step or in a series of steps.

Charge of an Electron: 1.6 • 10-19 Coulombs

Alpha Particle (α): A helium nucleus discharged in radioactive decay.

Beta -Minus Particle (ß-): An electron discharged when a neutron decays into a proton and electron through the weak nuclear force.

Beta -Plus Particle (ß+): A positron discharged in radioactive decay.

Positron: The anti-particle of an electron, has a charge of +1.

Half Life: The amount of time it takes for one half of a mass to decay,

Wave Equation: Let λ (lambda) denote the wavelength in metres, let v(nu) denote frequency in Hertz (s-1). The equation for a wave is given by: c = λv.

Plank’s Constant: 6.63 • 10-34 J•s (Joule-seconds).

Energy of a Wave: Let v(nu) denote frequency and E denote energy, the energy of a wave is given by the equation: E=hv.

Particles as a Wave (de Broglie Equation): Particles may exhibit the characteristics of a wave according the equation shown where m is mass and v is velocity: λ =h/mv.

Energy Levels for Electrons: Electrons are only allowed in certain defined orbitals where the higher the orbital the higher the energy. Moving down an orbital causes energy to be released. These orbitals may be divided into subshells, namely the S, P, D, and F shells. The S shell has a spherical shape, the P shell may be divided into 3 bell orientations, the D shell 5, and F shell 7. The D shell may hold one pair of electrons, the P shell 3 pairs, the D shell 5 pairs and the F shell 7 pairs.

Electron Configurations: Describing each orbital by notating each subshell, its orbital and the number of electrons in the subshell in the form below:

1s22s22p63s23p4 (silicon)

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Figure VII – Orbital Diagram | | | | | | | | |
| 1s | 2s | 2p | | | 3s | 3p | | |
| ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ |

Orbital Diagrams: A depiction of electrons as up and down arrows in orbital boxes. The s sublevel will have 1 box, p sublevel 3, d sublevel 5 and f sublevel 7. An example is below:

Core Notation: The description of electrons of atoms by only showing the outermost orbital.

D and F Shell Orbital Numbers: The orbital number of a D shell is one less than the preceding S shell and the orbital number of an F shell is 2 less than the preceding S shell. The reason for this I know not of.

Orbital Exceptions: 5 important exceptions to the rules of orbitals are listed below:

I-Cr is [Ar] 4s13d5, this is to attain a half-full d-shell.

II-Mo is [Kr] 5s14d5, this is to attain a half-full d-shell.

III-Cu is [Ar]4s13d10, this is to attain a full d-shell.

IV-Ag is [Kr] 5s14d10, this is to attain a full d-shell.

V-Au is [Xe} 6s14f145d10, this is to attain a full d-shell.

Note that all of these exceptions relate to having a half-full or full d-shell.

Isoelectronic: When different elements share the same electron structure due to one or both being ionized.

Order of Removing Electrons (Ionizing) Metals: When removing electrons while ionizing an element removes the s-shell first from transition metals or p-shell first, s-shell second for post-transition metals.

Quantum Numbers: Each electron in an atom or ion may be described by four numbers, these quantum numbers are n, l, m1 and m2. n represents the major orbital regions and has a value 1-7, l is called the angular momentum quantum number and describes the sublevel (s, p, d, f), its value ranges from 0 to n-1. m1 is called the magnetic quantum number and refers to the orientation of the individual orbitals and has a range of ±l. m2 is associated with electron spin with +1/2 being up spin and -1/2 being down spin.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Figure VIII- Quantum Numbers Table | | | | |
| n | l | subshell | m | # of Orbitals |
| 1 | 0 | 1s | 0 | 1 |
| 2 | 0 | 2s | 0 | 1 |
|  | 1 | 2p | -1 0 1 | 3 |
| 3 | 0 | 3s | 0 | 1 |
|  | 1 | 3p | -1 0 1 | 3 |
|  | 2 | 3d | -2 -1 0 1 2 | 5 |
| 4 | 0 | 4s | 0 | 1 |
|  | 1 | 4p | -1 0 1 | 3 |
|  | 2 | 4d | -2 -1 0 1 2 | 5 |
|  | 3 | 4f | -3 -2 -1 0 1 2 3 | 7 |

Alkali Metals: The elements that reside in the first column of the periodic table excluding hydrogen, they are softer than most metals and are the most reactive metals.

Alkaline-Earth Metals: The elements that reside in the second column of the periodic table, like alkali metals they are very reactive,

Transition Metals: The elemnts that reside in columns 3-12 of the periodic table.

Inner Transition Metals (Rare Earths): The elements that reside between columns 3-4 and are placed below the periodic table. The first row is the Lanthanide Series and the second row the Actinide Series.

Post-Transition Metals: The elements excluding Aluminium that are between the transition metals and the Amphoteric line.

Non-metals: The elements that reside on the right of the Amphoteric line.

Halogens: The elements that reside in column 17 of the periodic table. They are highly reactive.

Noble Gases: The elements that reside in column 18 of the periodic table. They do not react easily due to having a full valence shell.

Metalloids: The elements that run along the Amphoteric line which have properties of both metals and non-metals.

Atomic Radius: The distance from the nucleus to the outermost electron.

Cation: A positive ion.

Anion: A negative ion.

Predicting Atomic and Ionic Radii: The general size of atomic and ionic radii is given by the following four rules:

I-Down any column is larger.

II-Left on any row is larger.

III-The atom is always larger than the cation of the same element.

IV-The anion is always larger than the atom of the same element.

Ionization Energy: The energy required to remove an outside electron from an atom. This energy increases as you move up and right on the periodic table. In addition, half and full shells require a higher ionization energy.

Electron Affinity: The attraction of atoms for additional electrons. It increases as you move up and right on the periodic table.

Electronegativity: A value assigned on a scale from 0-4 which measures the attraction of atoms for shared electrons and combines both ionization energy and electron affinity. High electronegativity indicates an atom wants to acquire additional electrons.

Ionic Bonds: Chemical bonds formed when one atom loses an electron and another atom gains an electron. These bonds occur when one atom has a low electronegativity and another atom a high electronegativity. These most often involve a metal bonding with a non-metal.

|  |
| --- |
| Figure IX- Lewis Dot Structure |
| A picture containing clipart  Description generated with high confidence |

Covalent Bonds: Chemical bonds formed when atoms share electrons so that part of the time they have full orbitals. These bonds are likely to be found when atoms have similar electronegativities meaning diatomic atoms will always be covalent.

Bond Length: The distance between nuclei in a bond.

Valence Electrons: The electrons that occupy the s and p sublevels.

Lewis Dot Structure: A way of representing the valence electrons of an atoms by arraigning them around an element’s chemical symbol as demonstrated in figure IX.

Valence Shell electron Pair Repulsion (VSEPR): A theory used to predict molecular geometries. In these geometries ‘A’ represents the central atoms, ‘X’ represents atoms attached to the central atom by covalent bond, ‘E’ represents unshared electron pairs attached to the central atom and a line represents covalent bonds between the central atom and attached atoms. Below are some geometries created by VSEPR.

A screenshot of a cell phone

Description generated with high confidence

Figure X- VSEPR Geometries

Polar molecules: Molecules that have a covalent bond, but the elements involved have different electronegativities such that there is some ‘stealing’ and electrons spend more time with the element with the higher electronegativity. All diatomic molecules are not polar.

Hydrogen Bonding: A weak bond between two molecules resulting from the polarity of both.

Rules for Problems in Stoichiometry:

I-If you have a chemical formula, the molecular mass can be calculated.

II-If you have grams in a problem you will have to use molecular mass.

III-If you have atoms, ions, molecules, isotopes or formula units in a problem you must use Avogadro’s number.

IV-Molar volume of a gas = 22.4 L/mol at STP.

Coulombic Attraction: The electrical attraction between oppositely charged particles. The potential energy is inversely proportional to distance and is given by , with d representing distance, q1 representing one charge and q2 representing the other. K is a positively-valued proportional constant. Due to this V will always be a negative value, since V will only equal 0 when d is equal to infinity, and as d decreases then V will move towards negative infinity.

Core Charge (Effective Nuclear Charge): The charge of the atom’s core, its nucleus and inner shells. It is calculated by subtracting the charge of inner shells from the nucleus.

Emission of Energy: The amount of energy emitted from a system (▲E) is given by ▲E = nhv, with n being an integer and nu being the frequency of radiation.

Photoelectric Effect: The phenomenon related to the emission of electrons from a metal when light strikes it characterized by the following rules-

I-No electrons are emitted below a threshold frequency v0.

II-No electrons are emitted despite intensity of light below v0.

III-Electrons emitted increased with intensity while above v0.

IV-The kinetic energy of emitted electrons increases with frequency as long as its frequency is greater than v0.

Diffraction: Diffraction occurs due to light being scattered by an array of points and results in colours due to varying wavelengths.

Atomic Spectrum of Hydrogen: Hydrogen emits a few discrete wavelengths of light; the wavelengths indicate that only certain energies are allowed for electrons in the hydrogen atom.

Photoelectron Spectroscopy: A method used to determine the ionization energy of the electrons in an atom bombarding a gaseous sample of atoms with x-rays or ultraviolet light of known energy. The energy that the ejected photoelectrons have is measured and used to determine ionization energy.

|  |  |
| --- | --- |
| Figure XI-Regions of the Electromagnetic Spectrum | |
| Region | Wavelength |
| Radio wave | 3\*10-2 ↔1\*103 m |
| Microwave | 1\*10-3↔ 3\*10-2 |
| Infrared | 8\*10-7↔ 1\*10-3 |
| Visible | 4\*10-7↔ 8\*10-7 |
| Ultraviolet | 1\*10-8↔ 4\*10-7 |
| X-Ray | 1\*10-10↔1\*10-8 |
| γ-Ray | <10-10 |

Mass Spectroscopy: A method used to determine the isotopes of elements by ionizing them, accelerating them through a magnetic field and deflecting them towards a detector. The amount they are deflected by the magnetic field is dependant on charge and mass and can be used to calculate the mass of the original atom.

Fragmentation: When a molecule breaks apart due to collision with high speed particles.

Electron Spin: Electrons have an intrinsic angular momentum which contributes to their magnetic moment called its spin. Though this may seem to imply electrons spin, it rather imitates spinning without actually doing so.

Pauli Exclusion Principle: In a given atom no two electrons may have the same quantum numbers. Note as a result electron in the same orbital have opposite spins.

Aufbau Principle: In the ground state of an atom or ion electrons fill atomic orbitals of the lowest available energy levels before occupying higher ones. This results in the most stable electron configuration possible.

Hund’s Rule: The lowest electron configuration for an atom is the one having the maximum number of unpaired electrons allowed by the Pauli exclusion principle in a particular set of degenerate orbitals.

Bohr Model: A model in which electrons travelled in pre-defined circular orbits around an atom. Though it proved accurate for only hydrogen it served as a basis for later models.

Precipitation Reaction: A reaction where a solid called precipitate is produced from the mixture of two solutions.

Rules for Solubility of Salts in Water:

I-Most nitrate (NO3-) salts are soluble.

II-Most salts containing the alkali metal ions and ammonium (NH4+) are soluble.

III-Most chloride, bromide, and iodide salts are soluble. Notable exceptions include Ag+, Pb2+ and Hg2+.

IV-Most sulfate salts are soluble exceptions include BaSO4, PBSO4, Hg2SO4 and CaSO4.

V-Most hydroxides are slightly soluble. The important soluble hydroxides are NaOH and KOH; Ba(OH)2, Sr(OH)2 and Ca(OH)2 are marginally soluble.

VI-Most sulphide (S2-), carbonate (CO32-), chromate (CrO42-) and phosphate (PO43-) salts are only slightly soluble with the cations in rule II being the exceptions.

Slightly Soluble Terminology: The tiny amount of solid that dissolves are barely noticeable, making it almost insoluble and both terms are used interchangeably.

Formula Equation: A chemical equation that gives the overall reaction stoichiometry, but not necessarily the actual forms of the reactants and products in a solution.

Spectator Ions: Ions that do not participate directly in a reaction.

Complete Ionic Equation: A chemical equation which represents as ions all reactants and products that are strong electrolytes.

Net Ionic Equation: A chemical equation that includes only the solution compounds that undergo a change. Spectator ions are not included.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Figure XII- Average Bond Energies (kJ/mol) | | | | |
| H-H: 432 | C-H: 413 | C-Cl: 339 | N-N: 160 | O-H: 467 |
| H-F: 565 | C-C: 347 | C-Br: 276 | N-F: 272 | O-O: 146 |
| H-Cl: 427 | C-N: 305 | C-I: 240 | N-Cl: 200 | O-F: 190 |
| H-Br: 363 | C-O: 358 | C-S: 259 | N-Br: 243 | O-Cl: 203 |
| H-I: 295 | C-F: 485 | N-H: 391 | N-O: 201 | O-I: 234 |
| F-F\*: 154 | F-Cl: 253 | F-Br: 237 | Cl-Cl: 239 | Cl-Br: 218 |
| Br-Br: 193 | I-I: 149 | I-Cl: 208 | I-Br: 175 | S-H: 347 |
| S-F: 327 | S-Cl: 253 | S-Br: 218 | S-S: 266 | Si-Si: 340 |
| Si-H: 393 | Si-C: 360 | Si-O: 452 | C=C: 614 | C≡C: 839 |
| O=O: 495 | C=O: 745 | C=O (CO2): 799 | C≡O: 1072 | N=O: 607 |
| N=N: 418 | N≡N: 941 | C≡N: 891 | C=N: 615 | At-At: 116 |
| \*due to electron-electron repulsion F-F’s value is lower than expected by intuition. | | | | |

Diploe Moments: A trait shared by most polar molecules, diploe moments occur when a molecule has a negative centre of charge and a positive centre of charge and align themselves along those centres in a magnetic field.

Lattice Energy: The change in energy that takes place when separated gaseous ions are packed together to form an ionic solid. The energy associated can be broken down as seen below in figure XIII-

Figure XIII- Lattice Energy

Li (s) + ½F2 (g) 🡪 LiF (Total Reaction)

Li (s) 🡪 Li (g) (Heat of Sublimation [161 kJ/mol])

Li (g) 🡪 Li+ (g) + e- (First Ionization [520 kJ/mol])

½ F2 (g) 🡪 F (g) (F2 Bond Energy [72 kJ/mol])

F (g) + e- 🡪 F- (g) (Fluorine Electron Affinity [-328 kJ/mol])

Li+ (g) + F- (g) 🡪 LiF (s) (Lattice Energy [-1047 kJ/mol])

Exceptions to the Octet Rule: Molecules often form in ways that defy convention and the octet rule. Known as extended or reduced octets they follow the below rules-

I-Only third row or heaver elements may have an extended octet by using their d-shell.

II-Elements such as B or Be often form a reduced octet due to their reactivity.

Resonance: The electron structure of an atom may not always be represented by a single Lewis structure, when multiple valid Lewis structures exist for a molecule they are called resonance structures and represent the element’s delocalisation of electrons.

A picture containing clock, object

Description generated with very high confidence

Figure XIV-Resonance Structures

Formal Charge: The formal charge of an atom in a molecule is given by subtracting the number of associated electrons (with bonds split evenly) from the number of protons the atom has. The formal charge can be used to determine the optimal structure of a molecules as molecules tend towards minimizing formal charge.

Hybridization: The combination of native orbitals to form molecular orbitals. The type of hybridization is denoted by the associated orbitals such as sp or sp3.

π Bonds: The bond type associated with p orbitals found in sp2 and sp3, typically weak and relatively easily broken.

σ Bonds: The bond type associated with sp orbitals. Typically, strong and not easily broken.

Dipole-Diploe Attraction: Attraction between molecules aligning with diploe moments. ~1% as powerful as covalent bonds.

Hydrogen Bonding: Strong dipole-dipole attraction between highly polar molecules containing H.

London Dispersion Force: The movement of electrons within nuclei causes temporary asymmetries which can induce asymmetries in other atoms. Known as instantaneous diploes, they can cause a temporary and weak attraction primarily seen in nonpolar molecules and noble gases, but present in all atoms. In very large atoms they may be more influential than hydrogen bonds.

Alloy: A substance that contains a mixture of elements and has metallic properties.

Substitutional Alloy: An alloy where host atoms are replaced by similarly sized atoms.

Interstitial Alloy: An alloy formed when interstices (holes) are occupied by smaller atoms. Ex: Steel.

Network Solids: Atomic solids with strong directional covalent bonds. Properties include being brittle and a poor conductor of heat or electricity.

Ceramics: Non-metallic materials that are strong, brittle, and resistant to heat and chemicals.

Semiconductors: Elements with a relatively low energy gap between the highest valence orbital and the lowest unoccupied orbital. They typically have an average of 4 electrons per atom in the molecule.

n-type Semiconductors: A semiconductor which is doped with atoms with a lower higher number of valence electrons. The extra electrons are easily excited into the conduction band and increase general conductivity.

p-type Semiconductor: A semiconductor which is doped with atoms with a lower number of valence electrons. The deficit in electrons creates more holes in the valence band for electrons to jump into and increases conductivity.

Oxidation-Reduction Reactions (Redox Reactions): Reactions where one or more electrons are transferred.

Oxidation States: A method used to determine the location of electrons in an oxidation reduction reaction. In general, for a covalent bond between diatomic molecules the electrons are shared evenly, for polar molecules the electrons are assigned to the more electronegative atom and neutral atoms have an oxidation state of 0.

Oxidation: An increase in oxidation state.

Reduction: A decrease in oxidation state.

Oxidizing Agent: Electron acceptor in a redox reaction.

Reduction Agent: Electron donor in a redox reaction.

Electrolysis: The process of forcing a current through a cell to produce a chemical change. In electrolysis cell potential is negative.

Electrolytic Cell: A cell which uses electrical energy to produce chemical change.

Ampere (A): A unit of measurement representing 1 coulomb of charge per second.

Half Reaction Method: A method to solve redox reactions by dividing them into the oxidation reaction and reduction reaction and then combining. In the process fist charges, then atoms must be balanced with the addition of H2O or OH- sometimes involved depending on if the reaction is acidic or basic respectively.

Galvanic Call: A device that converts chemical energy to electrical energy and consists of two compartments connected by a wire and salt bridge/porous disk with the wire connected to metallic bars. The electricity is generated from redox reactions between the compartments with the solution in each interacting with the metallic bars. The site where oxidation occurs is the anode and the site where reduction occurs is the cathode.

Cell Potential (Electromotive Force [emf]): The driving force on electrons measured in volts (joules per coulomb). It is determined for a galvanic cell by standard reduction potential which assigns a reduction potential for half reactions based off the premise that 2H+ + 2e- 🡪 H2 has a 0V reduction potential. This is done as half reactions cannot be determined in strength. Th cell potential of a cell is determined by adding the reduction potential of the cathode and subtracting that of the anode. The result must be positive for a galvanic call. Notably reduction potential isn’t affected by quantity of reaction as it is intensive.

Line Notation of a Galvanic Cell: A notation of a galvanic cell in the form X(s) | Xa+(aq) || Yb+(aq) | Y(s).

Faraday: A constant which is equal to the charge on 1 mole of electrons (96,485 coulombs per mole).

Maximum Cell Potential: The maximum cell potential of a galvanic cell is denoted by ▲G° = -nFξ°, withξ° being cell potential, F being Faraday’s constant and n being the number of moles of electrons. The negative sign indicated the spontaneity of the process.

Concentration and Cell Potential: An imbalance in concentration of a cell favours the side of a reaction with the higher concentration. This is quantified by the Nernst Equation ξ = ξ° - or the equation .

Equilibrium: A state where the net change in a reaction is nothing and the reaction proceeds back and forth at equal rates.

Law of Mass Action: Given an equilibrium condition of the form jA + kB ⇌ lC + mD there is an equilibrium constant k = . The constant is the same at a given temperature regardless of amount mixed and each set of equilibrium concentrations which satisfy k are equilibrium positions.

Heterogeneous Equilibria: With equilibrium involving pure solids or liquids neither ae included in the equilibrium equation for k because as long as they are pure their concentrations cannot change.

Reaction Quotient: The value calculated when applying the law of mass action to initial concentrations which can be used to predict the direction of a reaction with the reaction at equilibrium if Q = K, shifting left if Q > K and shifting right if Q < K.

Guidelines for Solving Equilibrium Problems

I-Write the balanced equation for the reaction.

II-Write the equilibrium expression using the law of mass action.

III-List the initial concentrations.

IV-Calculate Q and determine the direction of shift.

V-Define the change needed to reach equilibrium and define the equilibrium concentrations by applying the change to the initial concentrations.

VI-Substitute the equilibrium concentrations into the equilibrium expression and solve for the unknown.

VII-Check the calculated equilibrium concentrations against K.

Le Châtelier’s Principle: If a change is imposed on a system at equilibrium, the position of the equilibrium will shift to reduce the change. If the change is a change in concentration, then the equilibrium position will shift to lower or increase the changed concentration depending if it was increased or decreased respectively. If an inert gas is added total pressure increases, but partial pressure and concentration remain constant. When the volume of the container holding a gaseous system is reduced, the system reduces its volume by shifting to the side with the fewest moles. The converse is also true. When the temperature changes the system responds with a change in K which favours the energy being used.