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 | Hydragyrum | Mercury |
| Sn | Stannum | Tin |
| Sb | Stibium | Antimony |
| Pb | Plumbum | Lead |
| W | Wolfgram | 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

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: 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.