Outline of General Chemistry

Contents

Note: To look up references, see the Consciousness Bibliography, listing 10,000 books and articles, with full journal and author names, available in text and PDF file formats at http://www.outline-of-knowledge.info/Consciousness_Bibliography/index.html.

PHYS>Chemistry>General

General Chemistry covers inorganic chemicals, chemical bonding, chemical reactions, and chemical phases.

PHYS>Chemistry>General>Chemicals

elements

Have one atom type.

allotrope

Elements, such as carbon and sulfur, can have several physical forms {allotrope}|.

atomic size

In periodic-table rows, right-most atom is half left-most-atom diameter. Last-row atom diameter is much greater than first-row atom diameter.

electron affinity

In gases, atoms have potential energies, which range from -0.9 eV to +3.6 eV, to attract additional electrons. Atom electronegativity directly correlates with electron affinity. Elements in top-right periodic table have highest electron affinity and highest ionization energy, because they have relatively small diameter and relatively large proton number. Atom diameter and proton number determine electrical properties. Diameter changes have more effect than proton-number changes. Noble gases have lowest electron affinity and highest ionization energy, because electron shells are full.

ionization potential

To remove outermost electron, gas atoms require energy, which ranges from 4 eV to 24 eV. Elements in top-right periodic table have highest electron affinity and highest ionization energy, because they have relatively small diameter and relatively large proton number. Atom diameter and proton number determine electrical properties. Diameter changes have more effect than proton-number changes. Noble gases have lowest electron affinity and highest ionization energy, because electron shells are full.

hydrogen

atomic number 1. H_2 can form polar covalent bonds with non-metals, making clear liquid acids. Hydrogen forms hydrides with strongly reducing metals. Hydrogen gas is mild reducing agent and reacts slowly. Chemical reactions involving hydrogen ion and hydride ion are fast.

inert gases or noble gases

helium - atomic number 2.
neon - atomic number 10
argon - atomic number 18

krypton - atomic number 36
xenon - atomic number 54
radon - atomic number 86.

alkali metals

lithium - atomic number 3
sodium - atomic number 11
potassium - atomic number 19

alkaline earth metals

beryllium - atomic number 4
magnesium - atomic number 12
calcium - atomic number 20

boron - atomic number 5. Boron is solid, can polymerize, form rings, and be in white borates.
aluminum - atomic number 13.

carbon - atomic number 6.
silicon - atomic number 14. Can form silicates with oxygen to make sand, asbestos, mica, glass, and quartz. Etching it makes semiconductor circuits.

pnictogens

nitrogen - atomic number 7. N_2 can be inert gas. Ammonia is in basic fertilizer. Nitrous oxide is anesthetic. Nitrogen-nitrogen double-bond diazo compounds are solid dyes.

phosphorus - atomic number 15, with 15 protons and 15 electrons. non-metal, waxy, white or red, and solid. Atomic weight is 30.97, so phosphorus has 16 neutrons. reactive.

chalcogens

oxygen gas - atomic number 8. O_2 can be gas and be oxidizer but be unreactive at low temperature. Ozone is powerful oxidizer. Two oxygens make the oxidizer peroxide ion. Oxygen combines with hydrogen to make water.

sulfur - atomic number 16. Elements {sulfur, element} can make sulfates and sulfites with oxygen, as well as sulfoxides for detergents. Hydrogen-sulfide gas has rotten-egg smell, as does carbon disulfide.

halogens

fluorine - atomic number 9. F_2 can be reactive gas that forms strong polar covalent bonds with non-metals and forms ionic bonds to metals.

chlorine - atomic number 17. Cl_2 can be gas, be reactive, and make hydrochloric acid. Chlorine makes chlorates, such as bleach, with oxygen.

bromine - atomic number 35. Br_2
iodine - atomic number 53. I_2

transition metals

Columns 3 to 12, rows 4 through 7

lanthanides

58 to 71

actinides

90 to 103

metal

Most elements are solid at room temperature, melt at high temperature, are gray to white in color, shine if polished, conduct electricity and heat, are malleable, are ductile, are dense, and tend to lose electrons in chemical reactions. The most-metallic elements are in periodic-table lower left. Elemental metals can bind to themselves in pure metals or alloys. They can bind to non-metals to make salts.

magic mountain - Lead is stable at 82 protons and 126 neutrons.

metalloid or semimetal

Few elements are soft and crumbly solids or hard and brittle solids, have low melting temperature, are fairly shiny, are gray or colorless, are semiconducting, are not malleable, are not ductile, are rocklike, and have medium density.

non-metal element

Elements can be colorless gases or colored soft solids, have low melting point, have no shine, have no conductivity, have low density, and tend to gain electrons in chemical reactions. The most-non-metallic elements are in periodic-table top right.

transuranium element

Nuclei can be heavier than uranium.

radioactive element

Radium compounds glow in the dark {radioactive element}|. Uranium and plutonium compounds are fuels for nuclear reactors.

ion

Atoms can donate electrons to or accept electrons from other atoms, so atom becomes positive or negative.

anion - Atoms with strong electric forces can gain several electrons to fill shell.

cation - Atoms with strong electric forces can lose several electrons to make empty shell.

compounds

Substances can have different atoms bound together.

molecule

Molecules have diameters from 10^-8 centimeters to 10^-5 centimeter.

binary compound - two different elements.

triatomic compound - three atoms.

polyatomic compound - more than three atoms.

metal complex

Molecules can have metal ions with d orbitals, which chemically bind other molecules at smaller energy levels than s or p orbitals, lowering energy levels to visible-light range from ultraviolet-light range. Iron compounds are red. Cobalt compounds are blue. Nickel compounds are green. Copper compounds are blue or green. Lead compounds are white. Silver compounds are black.

radical

Molecules can have no charge but have only one unpaired electron in outer orbital. Peroxides have oxygen free radicals. When peroxide or double-bonded carbon binds to carbon, carbon atom can have free radical.

molecular weight

In molecules, atom atomic-weight sum is molecule mass, in atomic mass units.

percent composition

In compounds, element percentage {percent composition} is atomic mass, in atomic mass units, multiplied by number of atoms, divided by molecular weight, in atomic mass units.

mole

Moles equals mass in grams divided by molecular weight in atomic mass units.

Substance mass, in grams, is moles multiplied by substance molecular weight, in atomic mass units. One mole = 6 * 10^23 molecules (Avogadro's number).

Avogadro hypothesis

Equal gas volumes at same temperature and pressure contain same number of molecules. At standard temperature and pressure, number is one mole, which is 6*10^23 molecules.

chemical naming

name for each symbol in formula in same sequence as in formula, in order of increasing electronegativity.

simple ion naming

If cation has one atom, use atom name followed by the word ion. For example, Na+ is sodium ion.

If anion has one atom, use atom root followed by "-ide". For example, O2- is oxide ion.

compound naming

first atom or ion name. If molecule is ionic, and metal ion can have more than one valence number, write metal-ion valence in roman numerals in parentheses.

For covalent molecules, if number of attached oxygens or other atoms is one, write "mono-". If two, write "di-". If three, write "tri-". If four, write "tetra-". If five, write "penta-". If six, write "hexa-".

If molecule is ionic, write second-ion root.

If molecule is covalent, write root of atom with attached oxygens or other atoms. Always add "-ide" to root.

For example, the ionic compound FeCl2 is iron (II) chloride. The covalent compound SO2 is sulfur dioxide.

acid naming

If anion name ends in "-ide", start with "hydro-", add anion root, and then add "-ic acid", as in hydrochloric acid.

If anion name ends in "-ate", start with anion root and then add "-ic acid", as in sulfuric acid.

If anion name ends in "-ite", start with anion root and then add "-ous acid", as in sulfurous acid.

complex ion naming

If ion has one hydrogen, begin name with "hydrogen". For two hydrogens, begin with "dihydrogen". For three hydrogens, begin with "trihydrogen".

If ion central atom can attach oxygen in more than two ways, use prefix "per-" for ion with the most oxygen atoms or prefix "hypo-" for ion with the least oxygen atoms.

central atom root. Root for C is carbon-. Root for N is nitr-. Root for O is ox-. Root for P is phosph-. Root for S is sulf-. Root for Cl is chlor-. Root for Mn is mangan-.

If ion central atom can attach oxygen in at least two ways, add "-ite" to root for ion with fewer oxygens or add "-ate" to root for ion with more oxygens.

If ion, add the word "ion".

ClO4- is perchlorate ion, ClO3- is chlorate ion, ClO2- is chlorite ion, and ClO- is hypochlorite ion.

CO3-- is carbonate ion.

NO3- is nitrate ion and NO2- is nitrite ion.

O2-- is peroxide ion.

PO4--- is phosphate ion.

SO3-- is sulfite ion. SO4-- is sulfate ion.

MnO4- is permanganate ion.

NH4+ is ammonium ion.

OH- is hydroxide ion.

CN- is cyanide ion.

C2H3O2- is acetate ion.

HCO3- is bicarbonate ion.

chemical formula

For each element, divide percent composition by atomic mass units, to find number of elements per molecular weight.

Then divide smallest number into others.

If all answers are close to whole numbers, use whole numbers as subscripts in chemical formula. If answers are not all whole numbers, multiply answers by two, then three, then four, and so on, until answers are whole numbers.

formula from name

Chemical formula comes from chemical name.

Because molecule has zero total charge, sum of ion charges and atom valences must equal zero.

First, write atom or complex-ion symbols in same sequence as in name. Remember or look up ion or atom charges or valence.

For atoms or complex ions, assign number subscripts so sum, of charge or valence times subscript, adds to zero.

chemical properties

hardness, strength, color, melting temperature, vaporizing temperature, tensile strength.

adhesiveness - attraction to other substances

cohesiveness - attraction to itself

ductility - wire-forming ease

elasticity - ability to return to original shape after stretching, compressing, or twisting

malleability - rolling-flat ease

compound color

Metal complexes can have metal ions with d orbitals, which chemically bind other molecules at smaller energy levels than s or p orbitals, lowering energy levels to visible light range, from ultraviolet for s and p orbitals. Iron compounds are red. Cobalt compounds are blue. Nickel compounds are green. Copper compounds are blue or green. Lead compounds are white. Silver compounds are black.

Organic molecules can have conjugated double bonds, which spread electron-orbital energies and lower energy levels to visible-light range, from ultraviolet for single bonds. Organic dyes and indicators have long carbon sequences and have lowest light frequencies, from red to blue.

PHYS>Chemistry>General>Chemical Bond

bonding of atoms

Molecule atoms have stable electrical attractions.

Only outermost electrons participate in chemical bonds, because they can contact another atom.

Atoms have potential energy. Bonding lowers potential energy. Because radius is less, filled outer-electron shells have lower potential energy than unfilled.

Bonding types result from competition between weak bonds with low activation energy and strong bonds with large activation energy. Single covalent bond overlaps sigma orbitals. Double covalent bond overlaps sigma orbitals and pi orbitals. Triple covalent bond overlaps sigma orbitals and two pi orbitals.

Bonding order is 1sigma bonding, 1sigma antibonding, 2sigma bonding, 2sigma antibonding, 2sigma bonding, 2pi bonding, 2pi antibonding, 2sigma antibonding, 3sigma bonding, 3sigma antibonding, 3sigma bonding, 3sigma antibonding, 3pi bonding, 3pi antibonding, and so on.

Bonded atom pairs always have same average distance between nuclei. Bond length has lowest potential energy.

Non-polar covalent bonds are stronger than polar, ionic, or metallic bonds, because atoms share electrons more equally and attractive force between electrons and nuclei is higher, making bond length less. Multiple bonds have shorter bond lengths, because nucleus shielding is more. Ionic bonds are stronger if ions have more charges and larger sizes. Metal alloys make stronger metallic bonds, because rapid electron transfer maximizes shell filling.

Average angle between two bonds is always the same.

electronegativity

Atomic nuclei attract electrons in shared orbitals. Electronegativity is proportional to sum of ionization potential and electron affinity. For two bonded atoms, electronegativity difference is proportional to square root of bond-energy ionic bonding part {partial ionic character}, which ranges from 0.8 eV to 4.0 eV.

If both atoms have high electronegativity, they have covalent bonding. If both atoms have low electronegativity, they have metallic bonding. If one atom has high electronegativity and one atom has low electronegativity, they have ionic bonding.

Atom with higher electronegativity has higher probability of containing bonding electrons, and lower probability of containing antibonding electrons, than other atom.

ionic bonding

Atoms can donate electrons to or accept electrons from other atoms, so one atom becomes positively charged and other atom becomes negatively charged, and opposite ion charges attract. Anions with strong electric forces can gain electrons to fill shell. Cations with strong electric forces can lose electrons to empty shell.

metallic bonding

Metal atoms exchange outer electrons to try to fill outer shell. Mercury is liquid, because it has weak metallic bonds.

covalent bond

Two atoms can share two electrons, which spend most time between the atomic nuclei and share a bonding orbital. Molecular electron orbitals fill with electrons using same rules as for filling atomic electron orbitals. Bonding orbitals fill before antibonding orbitals. Covalent bonding fills both atoms' outer shells.

Shared electrons can spend most time outside the atomic nuclei on line between nuclei, in antibonding orbitals. Outside electrons pull nuclei apart and so oppose covalent chemical bonding. Net bond number equals (bonding electrons - antibonding electrons) / 2.

Atoms with weak electric forces make covalent bonds. Atoms with weak electric forces can gain electrons to complete shell. Atoms with weak electric forces can lose electrons to empty shell.

catenation - Similar-size non-metals can form covalent bonds.

conjugation of bonds - Compounds with alternating single and double bonds have electron resonance.

effective atomic radius - In covalent compounds, ionization potential is inversely proportional to half the distance between two covalently bound nuclei. Effective atomic radius ranges from 0.037 nanometers to 0.3 nanometers.

polar bond

Two different-electronegativity atoms can bind by sharing electrons, and one atom attracts shared electrons more {polar bond}| {polar covalent bonding}.

sigma bond

Covalent chemical bonds can overlap atom s orbital and other-atom 1s or 2s orbital, 1p or 2p orbital, or s-p hybrid orbital, with constructive interference and electrons between nuclei. Sigma bonding orbitals are symmetric around line joining the atomic nuclei. Atom s orbital can overlap other-atom 1s or 2s orbital, 1p or 2p orbital, or s-p hybrid orbital, with destructive interference and electrons not between nuclei.

pi bond

Covalent chemical bonds can overlap atom 2p orbitals, so shared electrons are between nuclei but in two regions, one above and one below line between atomic nuclei. Atom 2p orbital can overlap other-atom 2p orbital, with destructive interference and electrons not between nuclei.

chelation

Metals can bind to anion and be soluble.

combining capacity

Atoms can make numbers of covalent bonds or can become ions.

delocalization

Molecule electrons can move among connected p orbitals and spread electron orbitals, which lowers potential energy by minimizing electron repulsions.

hybridization of bonds

Repulsions among electrons in different adjacent orbitals can change orbital shapes, to make same-shape orbitals. Orbital hybridization can happen if excited states are available.

Hybrid orbitals have ellipsoid shapes. Shape is between s and p orbital shape. Electrons are mostly on one atom side.

Four adjacent orbitals make four hybrid orbitals and tetrahedron shape {sp3 hybridization}. Three adjacent orbitals make three hybrid orbitals and triangle shape {sp2 hybridization}, with one unaffected p orbital in pi bond. Two nearby orbitals make two hybrid orbitals and line shape {sp hybridization}, with two unaffected p orbitals in pi bonds.

ligand of bond

Atoms or molecules can covalently or ionically bond to central atom, in different configurations depending on orbitals. Central atom can bind coordination number of ligands. Central atom can have six ligands, at octahedron corners {octahedral}, at two d, one s, and three p orbitals. Cubic ligand arrangement has six bonds, at cube corners, at two d, one s, and three p orbitals. Central atom can have four ligands {square planar}, at square corners, at two d, one s, and three p orbitals. Tetrahedral ligand arrangement has four ligands, at tetrahedron corners, at one s and three p orbitals.

Metal ions have five d orbitals. z^2 and x^2 - y^2 orbitals point along axes and have higher energy. Electrostatic repulsion causes xy, yz, and xz orbitals to point between axes and have lower energy. If field is weak, energy difference is small, and all electrons go to all five d orbitals, with parallel spins, as in Hund's rule. If field is strong, metal d-orbital energy differences are large, spins pair, and electrons stay in the three lower-energy d orbitals.

resonance of bonds

Hybrid bonding orbitals can represent alternative electron-arrangement averages, such as single and double bond interchanges.

Molecule electron orbitals spread when electrons move between connected p orbitals, lowering potential energy by minimizing electron repulsions.

Benzene and other aromatic compounds, which have five-member or six-member carbon and nitrogen rings with alternating single and double bonds, have electron resonance.

Compounds with alternating single and double bonds have electron resonance.

If an atom single-bonds to an atom type and double-bonds to same atom type, so bonds can interchange, as in carboxyl ion -COO- or guanidium ion -CNN-, compound has electron resonance.

shielding

Electrons between two atomic nuclei reduce electric repulsion, by reducing nucleus apparent positive charge.

van der Waals force

Weak electrical attractions {van der Waals force} are between charges induced on neighboring molecules by electronegative atoms. Van der Waals forces are at distances less than 0.25 nanometers. Higher-atomic-weight atoms make stronger van der Waals forces.

molecular orbital theory

Bonding theories can use molecular electron orbitals. Bonding orbitals have electrons between nuclei, which causes shielding. Antibonding orbitals have electrons beyond nuclei, with no shielding. Only these two wave-interference types result in net amplitude and so are the only bonding types.

Number of occupied bonding orbitals compared to number of occupied antibonding orbitals gives total bond number. If antibonding equals bonding, no bond forms.

Bond strength depends on atomic-orbital overlap, which is greatest for identical orbitals. If electronegativity difference between atoms is great, bonding and antibonding orbitals are similar in energy, because shielding is minimal.

Molecular orbital theory weights ionic effects too heavily.

valence-bond theory

Bonding theory {valence-bond theory} involves electric interactions between atoms. Molecular orbitals only form between valence electrons, because only valence electrons can contact outside world and valence electrons are least bound. Other electrons are too tightly bound. However, valence-bond theory weights ionic effects too lightly.

bond energy

Enough energy {bond energy} {energy, bond} can break bonds. Bond breaks at 40,000 to 260,000 calories per mole. Because shell is full, atoms in chemical bonds have lower energy and are more stable than isolated atoms. Symmetries in covalent chemical bonds make energy lower by minimizing electron repulsions.

London energy

Electrons can induce electric dipoles and cause electrical attractions {London energy} between atoms. London energy is at distance less than 0.6 nanometers.

bond types

single bond - One covalent sigma bond {single bond}| can be between two atomic nuclei.
double bond - One sigma and one pi bond {double bond}| can be between two atomic nuclei.
triple bond - One sigma and two pi bonds {triple bond}| can be between two atomic nuclei.

antibonding orbital

Molecular orbitals {antibonding orbital} can be difference between two atomic orbitals that have quantum-wave destructive interference. Antibonding orbitals have electrons beyond nuclei. No shielding makes positively charged nuclei repel. Outside electrons pull positively charged nuclei apart.

bonding orbital - Molecular orbitals {bonding orbital} can be sums of atomic orbitals that have quantum-wave constructive interference. Bonding orbitals have electrons between nuclei. Shielding reduces repulsion of positively charged nuclei. Between electrons pull positively charged nuclei together.

non-bonding orbital - Most atomic and molecular orbitals {non-bonding orbital} have no sharing, overlap, or quantum-wave interference.

hydrogen bond

Chemically bonded hydrogen atom near chemically bonded nitrogen, oxygen, or fluorine atom forms electric dipole {hydrogen bond}|. Nitrogen, oxygen, or fluorine unshared-electron pairs attract hydrogen nuclei. Only nitrogen, oxygen, and fluorine atoms are small enough for unshared-electron pairs to get close enough to hydrogen nucleus.

Electric attraction is one-tenth covalent-bond strength.

Hydrogen bonds break and reform in 10^-11 seconds.

Hydrogen bonds can have two configurations. Oxygen atom, hydrogen atom, and atom bonded to hydrogen can be in straight line. This hydrogen-bond type is stronger. Water has many strong-type hydrogen bonds. At hydrogen, angle of nitrogen, oxygen, or fluorine atom and atom bonded to hydrogen angle can be 109 degrees, as in tetrahedral configuration. This hydrogen-bond type is weaker.

Hydrogen bond polarizes atom bonded to hydrogen atom. Polarization aligns other atoms and causes more hydrogen bonding.

octet rule

Minimum potential energy is when molecule atoms, except hydrogen, have eight electrons, in four orbitals, in outer shell {octet rule}.

Lewis structure

Dots can represent electrons in molecule electron structures. Hydrogen has two dots. Other atoms have eight dots, to represent electrons in outer shell. Two dots are at atom right, bottom, left, or top. If two atoms bond, two dots are between them. If two atoms have double bond, four dots are between them.

valence of molecule

Atoms have number of outer-shell electrons, or missing outer-shell electrons, needed to complete outer shell.

unshared pair

Two non-metal-atom electrons can be in non-bonding outer-shell orbital.

rotation around bond

Molecular groups can rotate around single bonds. Double bonds, triple bonds, and bonds with resonance have no rotation.

Because spherical molecules are symmetric in all three space directions, spherical molecules have no net rotation. Spin around axis leaves molecule the same. Spherical molecules cannot rotate around axis that does not go through center.

Linear molecules can spin around axis perpendicular to chemical bond, so linear molecules can have net rotation. Because linear molecules are symmetric in one space direction, linear molecules have no net rotation around line between nuclei, because spin around that axis leaves molecule the same.

Molecules that are not spherical or linear have no symmetry axis and can rotate around three mutually perpendicular space dimensions.

vibration of bond

Molecule bonds can have different vibration types.

Vibrations can stretch and compress chemical bonds along line between nuclei. Vibrations can widen and narrow angle between two bonds.

Molecules with no bonds cannot vibrate.

Molecules with one bond have one vibration type, bond compressing and stretching.

Molecules with two bonds can have four vibration modes. One bond can stretch, as the other compresses. Both bonds can stretch and compress at same time. Angle between bonds can narrow and widen. One bond can move downward perpendicular to bond plane, while one moves upward perpendicular to bond plane.

Molecule symmetries can make two vibration modes indistinguishable and decrease total number of vibration modes.

energy partition

Atom rotational, vibrational, and translational energy modes have same energy. If energy is different, rotation or vibration gains or loses energy to neighboring rotations and vibrations and returns to equilibrium.

water

Water can exist in 13 crystalline phases and five amorphous phases. Water can be high-density amorphous ice at 10 K to 65 K or low-density amorphous ice at 65 K to 125 K. Amorphous ice is in interstellar space. Space amorphous ice can flow with UV light and allows carbon dioxide, carbon monoxide, methanol, and ammonia formation. Cubic ice forms at 135 K to 200 K. Hexagonal ice forms from 200 K to 273 K.

Ice has liquid water at surface, several molecules thick, with less structure than solid, because water interacts with air. Ice with impurities has thicker layer. Ice has a liquid surface layer even if it is tens of degrees below freezing. Water-surface-layer charge separation, on ice crystals moving upward and hail falling downward, causes lightning.

hard water - Water {hard water}| can have mostly calcium and magnesium ions.

soft water - Water {soft water}| can have mostly sodium and potassium ions.

compound properties

anhydrous - Most substances have no adhering water {anhydrous}|.

deliquescence - Solid can absorb water from air and become solution {deliquescence}|.

desiccant - Chemicals {desiccant}| that can take up water can keep other chemicals dry.

efflorescence of water - Hydrates can give water to air {efflorescence, water}|.

hydrate - Substances {hydrate}|, with dissolved ions, can adhere to water.

hygroscopic - Compounds can absorb water from air {hygroscopic}|.

PHYS>Chemistry>General>Chemical Reaction

chemical reaction

Molecule atoms can make and break chemical bonds. Chemical reactions make reactant molecules into product molecules. Chemical reactions typically release energy as heat.

Molecules have energy levels, with Boltzmann energy distribution. Reactants have higher ground-state potential energy and/or more widely spaced energy levels. If there is reaction path, molecules tend to become products, which have lower ground-state potential energy and/or less widely spaced energy levels. Potential-energy difference becomes kinetic energy and so heat. Products and heat have higher entropy than reactants.

Chemical-reaction rate depends on activation energy to go from reactants to products. Reaction rate depends on forward and backward chemical-reaction rates.

Balanced chemical equations allow knowing reactant or product amounts from reactant or product amounts. Ratio between unknown reactant or product coefficient and known reactant or product coefficient equals ratio between unknown reactant or product moles and known reactant or product moles.

Born-Haber cycle

Reaction diagrams can show how molecule chemical properties relate to atomic chemical-property combinations.

activation energy

Reaching chemical-reaction transition state requires energy. Transition state has potential energy that is higher than reactant potential energy and is higher than product potential energy. For drugs, activation energy equals site-atom attached-hydrogen effective activation-energy sum.

catalyst

Chemicals can increase reaction rate, but chemical reaction does not alter them. Reaction needs only small catalyst amount, because reaction reuses catalyst. However, catalysts can break down, have dirt or product coatings, or have surface damage. Catalysts reduce energy needed to start reaction. Catalysts allow transition state with lower activation energy, make molecule easier to attack, allow leaving group to leave easier, make attacking group attack better, orient molecules for optimum bond stretching, provide functional groups for forces or transfer, or line up reactant molecules. Enzymes are protein catalysts. Acids and bases are catalysts {homogeneous catalyst}. Basic catalysts cause isomerization, halogenation, or condensation. Acid catalysts cause tautomerism, solvolysis, or inversion. Neutral catalysts polarize solvent. Solid catalysts {heterogeneous catalyst} provide structured surfaces. Ceramic or metal catalysts are for industrial processes. Surface chemistry is for catalysis, corrosion, membranes, surface tension, and electrodes. If molecule collision energy with surface is same as surface thermal-vibration energy, surface can absorb molecule and collision energy. Molecule-absorption rate depends on collision energy. Electrode surfaces have an ion layer, covered by an opposite-charge ion layer. Catalytic surfaces must not bind too strongly or too weakly. Collision rate is not important, because absorption surface is large. Activation energy is small and not determining factor for surface catalysts. As atoms bind to catalyst, catalyst surfaces orient molecules and dissociate molecular bonds. Then new bond can form by collision or reorientation. Molecules on catalysts can move depending on impurities, defects, and crystal planes. Movement allows reaction atom transfer.

Gas molecules chemisorb on metals, because metal absorption area is much greater than gas collision area, so entropy decreases. How saturated surface is affects absorption. If concentration is high or time on surface is long, absorption is less. Because neighboring sites move, they affect absorption sites. Metals bind oxygen strongest, then acetylene, ethylene, carbon monoxide, hydrogen, carbon dioxide, and nitrogen. Platinum, iron, vanadium, and chromium can adsorb all these substances. Manganese and copper can adsorb some. Magnesium and lithium only absorb oxygen. Iron, nickel, platinum, and silver surfaces are catalysts for hydrogenations and dehydrogenations. Nickel oxide, zinc oxide, and magnesium oxide are catalysts for oxidations and dehydrogenations, because they are semiconducting. Metal sulfides are catalysts for desulfurations, because they are semiconducting. Aluminum oxide, silicon oxide, and magnesium oxide are catalysts for dehydrations, because they are insulators. Phosphoric acid and sulfuric acid are catalysts for polymerizations, isomerizations, alkylations, and dealkylations {cracking, petroleum}.

initiation of reaction

Chemical reaction starts when outside energy stretches, twists, or compresses molecule chemical bonds {initiation, reaction}.

Energy typically comes from heat or light. Light adds electric energy and affects electrons directly. Heat makes molecules move faster with more kinetic energy, causing more and higher-energy molecule collisions.

In large molecules, collision is less likely to disrupt bond, because collision is more likely to hit other bonds.

Molecule shape determines if collision affects bond. If collision is along bond line, bond disruption is more than if collision is from side.

Bond disruption is greater if colliding atoms have opposite electric charges. Bond disruption is greater if colliding atoms have same electric-charge absolute value.

photoactivation

Light can cause chemical reaction, as in photosynthesis.

transition state

Chemical bond is stable state with relatively low potential energy. Collision, heat, or radiation can stretch, twist, or compress chemical bond to maximum extent, as molecule electrical attractions resist chemical-bond disruption. Transition state has greatest disruption, highest potential energy, and maximum separation. If it can become new conformation or molecule, transition state is hybrid of stable chemical states before and after chemical reaction. From transition state, molecules can go back to original states or become new conformations or molecules, with equal probability.

virtual work principle - After displacement from equilibrium, system returns to equilibrium and sum of all work done by forces during displacement and return equals zero.

reaction rate

Chemical reactions proceed over time.

Reaction goes in two directions at once, from reactants to products {forward reaction} and products to reactants {reverse reaction}. Backward reaction rate divides into forward reaction rate to find overall rate.

Reactant amount eventually reaches half original amount {half-life, reactant}: half-life = C * (1 / c^(n - 1)), where C is constant, c is concentration, and n is reaction order.

Reaction rate depends on temperature, pressure, reactant concentrations, catalysts, states, and reactant physical forms: rate constant = (collision frequency) * e^(-E / (R*T)), where R is gas constant, T is temperature, and E is activation energy. If reactant concentration is in excess, concentration stays constant during reaction.

Reactants and products have initial, intermediate, and final concentrations. Reactions destroy reactants and makes products.

Reaction rate depends on reaction mechanism. Reaction mechanism can depend on zeroth, first, second, or third reactant-concentration power (reaction order). Reaction rate can be constant {zero-order reaction}. Reaction rate can depend on one reactant concentration or pressure {first-order reaction}. First-order reaction uses linear equation: rate = dC / dt = k * C0 where dC is concentration change, dt is time change, k is rate constant, and C0 is concentration. ln(C / C0) = -k*t, where C is concentration, C0 is initial concentration, k is rate constant, and t is time. Find final and intermediate product or reactant concentrations from initial concentration, rate constant, and time: Cf = Ci * e^(k*t), where Cf is final concentration, Ci is initial concentration, k is rate constant, and t is time. Reaction rate can depend on two reactant concentrations or pressures {second-order reaction}. Second-order reaction uses quadratic equation.

if rate constant = 3 and concentration is half or roriginal, time is -ln(1/2)/3

Reaction rate depends directly on temperature. Reaction rate is faster with higher temperature. 10-K increase doubles reaction rate.

Reactant physical form affects reaction rate. Greater surface area, lower viscosity, and higher solvent polarity increase reaction rate. If surfaces must touch for reaction, rate depends on contact area.

Reactant gas, liquid, or solid physical state affects reaction rate.

Physical factors that affect reaction rate are temperature, catalyst, physical form, and physical state. All physical factors are in one constant {rate constant}. People know rate constants for many chemical reactions.

In chemical-reaction series, in which previous-reaction products are next-reaction reactants, one reaction {rate-limiting reaction} is slowest.

Ionic reactions are fast if both reactants have opposite charge. Large ions and high-charge ions increase reaction rate. Increased ionic strength increases rate, if ions have opposite charge, but otherwise slows reaction rate. Solvents with high dielectric constants, like water, reduce repulsions and attractions between reactants and slow reaction rates. Acid-base reactions are ionic, and reaction rate increases with more acid or base. Ions can modify reaction by forming weak acids and bases. In ionic solutions, higher ionic strength, more polar solvent, and greater ion charge causes high collision rate and short contact time, so reaction rate is higher.

In non-polar solutions, higher viscosity makes contact longer and collision rate lower, so reaction rate is lower.

chemical activity

Substances have chemical reactivity. Chemical activity expresses true concentration or pressure. Substance concentration relative to other concentrations depends on chemical potential. Solids and pure liquids, including water, have chemical activity one. Metal activities, in decreasing order, are Li, K, Ba, Sr, Ca, Na, Mg, Al, Mn, Zn, Cr, Fe, Cd, Co, Ni, Sn, Pb, H, Cu, Ag, Pd, Hg, Pt, and Au. Non-metal activities, in decreasing order, are F, Cl, Br, and I.

affinity in chemistry

Chemical potential difference from reactants to products is chemical-reaction driving force.

fugacity

Substance partial pressure, relative to other partial pressures, depends on chemical potential.

reaction control

Methods can control reaction {reaction control}. In non-polar solution, if activation energy is low, diffusion controls reaction. In ionic solutions, if activation energy is high and is late in reaction, use vibration at frequency similar to rotation frequencies to control reaction, because bonds are short. In ionic solutions, if activation energy is high and is early in reaction, use translational energy to control reaction, because bonds are long.

flash photolysis - Lasers can initiate photolytic reactions {flash photolysis}.

flow technique - Mixing chambers and controlled reactant flows control reaction {flow technique}.

molecular beam - Molecule streams {molecular beam} can hit other molecules at precise speeds and orientations.

relaxation method chemistry - Temperature can change equilibrium {relaxation method, chemistry}, if reaction requires heat.

PHYS>Chemistry>General>Chemical Equation

chemical equation

Chemical-reaction equation uses molecule chemical formulas and special symbols. Reactant formulas are on left, and product formulas are on right.

Horizontal arrow pointing right separates reactants from products. Delta symbol means to add heat. hv symbol means to add light.

Plus signs separate molecules.

Up arrow (^) at formula right indicates that reaction produces gas. Down arrow at formula right indicates that reaction precipitates solid. The letter s at formula right means that reagent is solid. The letter l at formula right means that reagent is liquid. The letter g at formula right means that reagent is gas. The letters aq at formula right mean that reagent is aqueous.

Atoms on chemical-reaction left must also be on right, so both sides have same atom numbers and types.

reactant

Before chemical reaction, chemicals {reactant}| {reagent} exist.

limiting reagent

If chemical reaction has more than one reactant, one reactant depletes first as reaction proceeds. Find limiting reagent from balanced chemical reaction, using the following rule. If first-reactant coefficient to second-reactant coefficient ratio is larger than first-reactant moles to second-reactant moles ratio, first reactant is limiting reagent.

$2A + 3 B \rightarrow$ ... and 3 moles of both A and B. $2/3 < 3/3$ so first reactant is limiting.

product of reaction

After chemical reaction, new chemicals exist.

stoichiometry

Relative reactant and product masses have relations.

balancing equation

If written chemical reaction has one product or reactant missing, calculations can find missing product or reactant. If written chemical reaction has one coefficient missing, calculations can find missing coefficient.

First, find all missing atoms, because each atom on left must also be on right. Using found atoms, write positively charged atom symbol first and negatively charged atom symbol second.

Use naming-formula rules to find candidate molecule, using number subscripts for symbols if necessary.

Write equation using candidate molecule.

Add coefficients to reactants and products to make atom numbers equal on both sides. To find coefficients, first balance metal-atom coefficients, then balance non-metal-atom coefficients, except H and O, then balance hydrogen coefficients, and finally balance oxygen coefficients. If chemical equation is not yet balanced, double metal-atom coefficients, then balance non-metal-atom coefficients, except H and O, then balance hydrogen coefficients, and finally balance oxygen coefficients.

conservation of mass

In chemical reactions, total mass, total charge, and total energy stay constant. Sum of reactant charges equals sum of product charges. Total reactant mass equals total product mass. Reactant energy equals product energy plus heat.

Guy-Lussac law

In chemical reactions, formed or used gas volumes relate by whole-number ratios.

Hess law

In reaction series, in which previous-reaction products are next-reaction reactants, total change over series equals sum of reaction changes.

yield of reaction

Chemical-reaction product amount never equals maximum theoretical product amount, because reactions are inefficient. Calculating reaction efficiency {percent yield} uses the balanced chemical reaction. Percent yield equals ratio between product moles and limiting-reagent moles, expressed as percentage.

limiting reagent used 2 moles, and product was 1.5 moles, so yield = 75%.

equilibrium of reaction

After reaction, reactant and product amounts stay constant. At equilibrium, total-energy change is zero, free-energy change is zero, substance change is zero, all chemical potentials are equal, and all forces are equal. Product concentrations and reactant concentrations have equilibrium-constant ratio.

At equilibrium, forward and backward reaction rates are equal. Product-formation rate equals reactant-formation rate. Amounts do not change, so reaction is complete.

Equilibrium concentrations and amounts do not depend on catalyst or factors affecting reaction rate. Equilibrium concentrations depend only on energies and entropies.

If reaction requires heat, temperature increase makes more product.

If gas is reactant, pressure increase makes more product. If temperature increases, system acts to reduce pressure and so return to equilibrium.

Adding more reactants changes them to products, until equilibrium reestablishes. Adding more products turns them into reactants, until equilibrium reestablishes. Increasing reactant concentration, or removing product, increases product.

equilibrium constant

Knowing chemical equation and reactant and product concentrations at equilibrium allows reaction-constant calculation. Equilibrium constant is product of product concentrations, each raised to power of its chemical-equation coefficient, divided by product of reactant concentrations, each raised to power of its chemical-equation coefficient.

If 2 A + 3 B -> C + 4 D, equilibrium constant K = ([A]^2 * [B]^3) / ([C] * [D]^4).

Chemical reaction aX + bY -> cZ + dW equilibrium constant is K = (X^a * Y^b) / (Z^c * W^d).

People know many reaction equilibrium constants, at specific temperatures. Dissociating acids and bases have equilibrium dissociation constants. Dissolving salt in water or other solvent has equilibrium solubility constant.

Equilibrium constant greater than 10^9 means reaction is irreversible.

Equilibrium constant and initial reactant concentrations result in product concentration at equilibrium. First, use chemical equation to make equilibrium-constant equation with correct exponents. In equilibrium-constant equation, replace product concentration with x if coefficient is 1, replace with 2*x if coefficient is 2, replace with 3*x if coefficient is 3, and so on. If coefficient is 1, replace reactant concentration with its initial concentration minus x. Replace with 2 * (initial concentration minus x) if coefficient is 2. Replace with 3 * (initial concentration minus x) if coefficient is 3, and so on.

If 2 A + 3 B -> C + 4 D, equilibrium constant K = ([A]^2 * [B]^3) / ([C] * [D]^4). To find A concentration: K = ((2*x)^2 * B^3) / (C * D^4). Use equilibrium constant value from table of constants. Solve for x. Product concentration is x times its coefficient in chemical equation. Reactant concentration is (initial concentration minus x) times its coefficient in chemical equation.

Reactant and product partition functions can find chemical-reaction equilibrium constant.

reaction types

- chain reaction in chemistry - Product can be reactant, which can make more product. Reaction rate continually increases, until system physically disrupts.
- decomposition reaction - One reactant can make two products. Decomposition includes hydrolysis and dehydration reactions.
- electrophilic reaction - Chemical can attack negatively charged group.
- metathesis reaction - Two compounds can make two new compounds {double replacement reaction}. Acid-base reactions have metathesis. Metal compounds can catalyze carbon-carbon double-bond changes.
- molecular rearrangement - One reactant can change to same chemical in different configuration.
- nucleophilic reaction - Chemical can attack positively charged group.
- substitution reaction inorganic - Element and compound can make another element and another compound . Metal-atom to metal-ion oxidation has substitution.
- synthesis reaction - Two reactants can make one product. Synthesis includes polymerization, hydration, and oxidation reactions, like rusting and combustion.

reaction energies

- irreversible reaction - Energy transfer can involve permanent change that cannot reverse, because heat is made.

- reversible reaction - Energy transfer can have no friction or other opposing changes. In reversible reactions, external and internal temperatures and pressures are approximately the same. In reversible processes, system and surroundings are always in equilibrium. Reversible processes approximate slow energy transfer with small force and minimal resistance.
- spontaneous reaction - In reactions, activation energy can be less than difference in potential energy between transition state and products.

reaction kinds

- monomolecular reaction - Reactions can have one reactant, as in SN1 and E1 reactions. Molecule vibrations and rotations can cause molecule to decay to new state, as in gas decays, Type I nucleophilic substitutions, Type I eliminations, dissolution, and state changes.
- bimolecular reaction - Reactions can have two reactants, as in SN2 and E2 reactions. Molecule collisions can form transition states and can transfer energy or functional groups, as in isomerizations, Type II nucleophilic substitutions, Type II eliminations, enzyme reactions, syntheses, and dimerizations.
- termolecular reaction - Reactions can have three reactants, as in enzymatic reactions.

cracking

Phosphoric acid and sulfuric acid are catalysts for carbon-chain dealkylations. Petroleum separation uses phosphoric acid, sulfuric acid, silicon oxide, and aluminum oxide. Silicon oxide and aluminum oxide build branched hydrocarbons. Olefins form on platinum with silicon oxide, followed by isomerization, ring formation, splitting, and hydrogenation.

hypergolic

Two chemicals can bind to make something with different properties than original chemicals. For example, hydrazine and nitrogen tetroxide react when in contact to make nitrous oxide and water: N2H2 + NO4 -> 3 NO + H2O.

triboluminescence

Pressure can cause luminescence.

heat of reaction

Chemical reactions can release or absorb thermal energy.

endothermic reaction - Chemical reactions absorb energy if product potential energy is higher than reactant potential energy. Endothermic reactions make complex molecules and require high temperature or strong light at specific frequency.

exothermic reaction - Chemical reactions release energy if reactant potential energy is higher than product potential energy.

enthalpy - internal energy and thermodynamic potential.

entropy - thermal energy not available for work.

Gibbs free energy - reaction energy available for work.

PHYS>Chemistry>General>Acid-Base Reactions

acid

Chemicals can contribute hydrogen ions to water (Arrhenius acid). Acids donate protons {Brönsted acid}, accept electron pairs {Lewis acid}, or add hydrogen ions to water when they dissolve. Acids {polyprotic acid} can donate more than one proton.

Acids taste sour, are colorless, and are corrosive.

Dissolving non-metallic oxide in water makes acid.

For diatomic acids, acidity increases with negative-ion atomic weight. Acidity increases with increasing number of no-hydrogen oxygens around central atom.

Common acids are nitric acid, sulfuric acid, hydrochloric acid, hydrofluoric acid, carbonic acid, phosphoric acid, formic acid, acetic acid, and other carboxylic acids.

corrosive - Acids can react with metals and inorganic materials.

hydronium ion - Hydrogen ions in water bind to water molecules electrically to make positively charged ion: H+ + H2O -> H3O+.

base in chemistry

Chemicals can contribute hydroxide ions to water (Arrhenius base). Bases accept protons {Brönsted base}, donate electron pairs {Lewis base}, or donate hydroxide ions to water when they dissolve. Bases taste bitter, are colorless, are slippery, and are caustic. Dissolving metal oxide in water makes base. Bases include sodium hydroxide, potassium hydroxide, ammonium hydroxide, magnesium hydroxide, calcium hydroxide, and aluminum hydroxide.

caustic as base - Bases can react with organic matter.

acidity

Solution acidity is negative logarithm of hydrogen-ion concentration: pH = -log(H+). pH can range from 0 to 14. Pure water has dissociation constant K = 10^-7, so pK is 7, and pH is 7. Pure water is neither acid nor base. 1 M hydrochloric acid has pH 0. Lemon juice has pH 2. Soda water has pH 4. Coffee has pH 5. Urine and rain have pH 6. Water has pH 7. Bicarbonate of soda has pH 8. Milk of magnesia has pH 10. Cleaning ammonia has pH 11. 1 M sodium hydroxide has pH 14.

amphiprotic

Chemicals can either donate or accept proton.

amphoteric

Molecules can have both acidic and basic groups.

buffer

To keep solution pH constant, add weak acid or base and soluble salt with same anion. Weak-acid anion acts as weak base. Weak-base anion acts as weak acid. Adding acid or base to solution causes weak base or acid to neutralize added acid or base. However, adding too much acid or base can overwhelm weak acid or base. Weak-acid or base concentration to soluble-salt concentration ratio, and anion dissociation constant, determine buffer pH. Citrate buffer has pH near 5. Bicarbonate buffer has pH near 6. Phosphate buffer has pH near 7. Tris buffer has pH from 4 to 8.

conjugates

conjugate acid - After base accepts proton, it becomes weak acid.

conjugate base - After acid donates proton, it becomes weak base.

Ostwald dilution law - As concentration decreases, ionized-acid percentage increases.

hydrolysis

Weak acids have hydrogen ion and anion. Salts with anion react with water to associate some hydrogen ion and form weak bases. Weak bases have hydroxide ion and cation. Salts with cation react with water to associate some hydroxide ion and form weak acids.

indicator

Weak acids or bases with conjugated double bonds can change electronic structure and color at different pH. At pH 1, malachite green changes from yellow to green. At pH 2, thymol blue changes from red to yellow. At pH 4, bromphenol blue changes from yellow to blue. At pH 4, methyl orange changes from red to yellow. At pH 4.5, bromcresol green changes from yellow to blue. At pH 5, methyl red changes from red to yellow. At pH 7, bromthymol blue changes from yellow to purple. At pH 7.4, phenol red changes from yellow to red. At pH 9, phenolphthalein changes from clear to red. At pH 9, thymol blue changes from yellow to blue. At pH 10, thymolphthalein changes from clear to blue. At pH 11, alizarin yellow R changes from yellow to red.

titration

Add acid or base to neutralize base or acid.

moles acid = moles base.

neutralization

Acid and base reactions make water, metal anions, and non-metal cations. Neutralization reactions involve proton transfer. Acid and base neutralize each other, because metal anions and non-metal cations are not very acidic or basic.

dissociation in chemistry

Solutes dissolve in solvent. Buffer, weak-acid, or weak-base solution has low dissociation. Dissociation constant equals hydrogen ion concentration times anion concentration divided by acid concentration. Water dissociation constant = 10^-14, so hydrogen ion = 10^-7 M. Water ionization is more if temperature is more.

salt

In acid-base reactions, anion and cation can attract electrically to form compounds.

PHYS>Chemistry>General>Oxidation-Reduction

oxidation-reduction

Oxidizing agent and reducing agent react to make reduced molecule and oxidized molecule, respectively, by transferring one or more electrons.

Find redox-reaction equilibrium constant from balanced equation. From balanced equation, separate reduction and oxidation half-reactions. Find half-reaction standard potentials and total potential. Note number of electrons transferred. Equilibrium constant equals exponential of standard potential V times number of electrons transferred n times one faraday F, divided by gas constant k times temperature T: exp(V*n*F/k*T).

oxidation

Molecules can lose electrons. Losing electrons increases positive charge.

reduction

Molecules can gain electrons. Gaining electrons decreases positive charge.

oxidizing agent

Molecules can gain electrons from another molecule. Oxygen, halogens, permanganates, and chromates are oxidizing agents.

reducing agent

Molecules can lose electrons to another molecule. Small, light metals are reducing agents.

oxidation number

If all electrons shared in chemical bonds go to the more-electronegative atom, each atom has resultant charge. Oxidation number is number of electrons added to, or subtracted from, outer shell to make full shell. In molecules, atoms with higher electronegativity tend to gain electrons from atoms with lower electronegativity.

Metals have positive oxidation numbers, because they lose electrons and empty outer shell. Metals are reducers, because they themselves oxidize.

Oxygen and fluorine have negative oxidation numbers, because they gain electrons to fill outer shell. Oxygen and fluorine are oxidizers, because they themselves reduce.

Hydrogen can gain or lose one electron, making oxidation number +1 or -1.

Atoms can have several oxidation numbers, because they can fill or empty outer shell in different ways, through orbital hybridization.

Functional-group oxidation number is atom oxidation-number sum.

corrosion

Slow metal ionization can use oxidation, from oxygen in air or acidic water. Metal impurities make circuit from impurity to metal. Corrosion rate depends on exposed area. Aluminum oxide covers aluminum and prevents further corrosion.

anodized - Zinc coatings protect metal, because zinc corrodes first.

cathode protection - Because it is electron source, magnesium can prevent corrosion by returning electrons to metal.

cell in chemistry

Redox reactions can be in solutions. Conducting plates can be in two connected half-cells. Current from one electrode goes through wire to other electrode and then through solution. Metal and metal ion have potential difference, because electrons and ions separate. If potential > 0, metal favors reduction. If potential < 0, metal favors metal oxidation.

If electrode voltage is > 0, reaction is spontaneous. At equilibrium, voltage = 0 and current = 0.

Tables can show reduction-half-reaction potentials. Hydrogen electrode has standard potential 0 V.

In fast reactions, diffusion controls reaction rate. Moving electrode or stirring solution minimizes diffusion effects.

Agar and potassium-chloride salt bridge can connect half-cells.

Membranes that block and allow ion flows can have potential difference, because membrane sides have different ion concentrations. Membrane-permeable ion diffuses through until electrostatic repulsion from higher-concentration side stops ion flow.

half-reaction

Redox reaction has two parts that interact, reducing-agent oxidation and oxidizing-agent reduction. Cell redox reactions oxidize reducing agent and reduce oxidizing agent. For example, half-reaction for hydrogen electrode is 2 H+ + 2 e- <-> 1 H2 .

potential of half-cell

In oxidation-reduction reactions, electrons transfer. Electron transfer requires voltage. Oxidation-reduction half-reactions have potentials. Total oxidation-reduction reaction potential is sum of half-reaction potentials.

Electronegative atoms have high reduction potentials.

Cell standard potential depends on half-reaction potentials and balanced equation. From balanced equation, separate reduction and oxidation half-reactions. Tables have half-reaction reduction potentials at 25 C. Subtract reduction potential for oxidation half-reaction from reduction potential for reduction half-reaction to find chemical-reaction potential.

If oxidation-reduction chemical-reaction potential is greater than zero, chemical reaction is spontaneous. If reduction potential is more than oxidation potential, ionization potential is higher, hydration energy is lower, and sublimation energy is more.

When oxidation-reduction reaction is complete, potential is zero.

Nernst equation

Half-cell reduction-reaction voltage depends on oxidized and reduced concentrations, temperature, and number of electrons transferred. V = V0 - R * T / (n * ln([Co] - [Cr])), where V is reaction potential, V0 equals standard unit cell potential, R is gas constant, T is temperature, n is number of electrons transferred, Co is oxidized-ion concentration, and Cr is reduced-ion concentration.

Activation energy comes from electric field and from temperature.

Electrode voltage is V = V0 + R * T * ln(K / (z * F)), where V is voltage, V0 is standard potential, R is gas constant, T is temperature, K is equilibrium constant, z is absolute value of transferred charge {valence charge}, and F is 1 Faraday. Therefore, exp(-V / (R*T)) = K / (z * F) and K = z * F * exp(-V / (R*T)). Standard potential is at concentration 1 M, pressure 1 atmosphere, and temperature 25 C. Solids have concentration = 1. Voltage is always positive. If V > 0, reaction is spontaneous. At equilibrium, voltage = 0 and current = 0.

galvanic cell

Spontaneous redox reaction in cells can make current. In galvanic cells, anode oxidizes and is negative, while cathode reduces and is positive. Metal or metal-oxide electrode can be in solution {wet cell} or paste {dry cell}. Different metal or metal oxide can be in same solution or paste, or another solution or paste connected to first by conductor, to make two coupled cells and a battery. Metal reacts with solution to make ions. Metal anode loses electrons and becomes positive. Metal cathode gains electrons and becomes negative. Electrodes have potential difference. Batteries can have nickel and cadmium in acid solution. Edison cells have nickel oxide and iron electrodes in alkaline solution. Batteries can have lead and lead oxide in acid solution.

fuel cell - Galvanic cells can have continuous fuel supply. Fuel cells make electric current by oxidizing hydrides or other substances. Fuel cells are efficient, cool, and clean.

battery

Galvanic cells can connect to make electric current by solution chemical reactions. Batteries {dry cell battery} can use paste. Batteries {wet cell battery} can use liquid solution.

Metals used for electrodes depend on battery solution. Metal combinations make different battery voltages across electrodes. Typical voltage is one volt to three volts.

Electrochemical cells can connect in series, so cell voltages add. Automobile batteries use lead plates and sulfuric acid solution.

Some batteries can recharge by applied electric voltage and current.

primary cell - Some batteries cannot recharge.
secondary cell - Some batteries can recharge.

electromotive force

Potential difference produced by open circuit is voltage that can make electricity.

back electromotive force - Devices that make electric current have resistance force to current. Maximum battery power has battery back electromotive force equal to circuit resistance.

load on circuit

Current flows from battery negative terminal through circuit and power-using device to positive terminal and then solution.

electrolytic cell

Voltage applied to cells can force current through cell and cause reverse redox reaction. In electrolytic cells, cathode is electron source, and anode is electron sink. Applying voltage and current can split molecules by electrolysis. For example, water can form hydrogen and oxygen. Aluminum salts can make aluminum.

anode - In electrolytic cells, oxidation is at positively charged electrode. In galvanic cells, reduction is at positively charged electrode.

cathode electrode - In electrolytic cells, reduction is at negatively charged electrode. In galvanic cells, oxidation is at negatively charged electrode.

electrolysis

Electric voltage can electrolyze or electroplate material. Electrolysis uses potential to drive electrochemical reaction. Electric current can split solute molecules into two ions. Mass deposited in electroplating, or split in electrolysis, is proportional to total charge transfer. Mass deposited in electroplating, or split in electrolysis, is proportional to atomic weight to ion charge ratio {chemical equivalent}.

Find material electrolyzed or electroplated, or charge needed to electrolyze or electroplate material amount, from balanced equation. From balanced equation, separate reduction and oxidation half-reactions. Note number of electrons transferred. Moles of electrons used are coulombs divided by 96,500 Coulombs. Coulombs used is current amperes times seconds. Ratio of electrolyzed or electroplated product coefficient to transferred-electron number is ratio of electrolyzed or electroplated product moles to electrons-used moles.

Nernst potential is minimum voltage needed to reverse spontaneous reaction at given conditions. Concentration gradient at electrode surface can make overpotential. Total needed potential {decomposition potential} includes Nernst potential, overpotential, and electrical-resistance potential.

Temperature, current-to-area ratio, electrode surface, and electrode type affect reaction activation energy.

Constant-current electrolysis is for metals with reduction potential greater than hydrogen and for potential greater than hydrogen decomposition potential. Hydrogen ions in high-acidity solution carry constant current, because they are much more numerous than metal ions. Substrate keeps hydrogen gas low, so gas does not cover hydrogen electrode. Constant current times time makes total charge.

Constant-voltage electrolysis keeps potential high enough to lower metal-ion concentration to optimum level but low enough to stop hydrogen-gas evolution or other-metal deposition. In this method, current decreases over time.

Controlled-potential electrolysis uses third electrode (SCE) as reference to keep oxidation potential at cathode constant, keep current high enough, and prevent unwanted reactions. Current decreases over time exponentially. Cathode potential determines decomposition potential, so, as metal deposits, ion concentration goes down, and decomposition potential goes up.

equivalent

In electrolysis, material moles formed or reacted is electron moles divided by metal or other ion charge.

electroplating

Electrolysis can put metal on conducting surfaces. In silver plating, gold plating, and zinc plating, metal derives from salt solution. Electric current adds electrons to change ion to metal at electrode.

PHYS>Chemistry>General>Phase

Phase

Substance can be solid at low temperature, liquid at intermediate temperature, and gas at high temperature. States or phases have different relations between material volume, temperature, and potential energy.

Current phase has lowest free energy for temperature. Potential energy depends on distance between molecules. Entropy depends on molecule number and temperature.

Solid phase has lowest volume and smallest potential energy, because average distance between molecules is smallest. Solids have low chemical potential. Low temperature makes material solid, because decrease in potential energy is higher than decrease in entropy. Solid phase has most order, because it has patterned crystal structure and temperature is lowest. Solid phase has least randomness and lowest entropy. Solids can have several crystal forms and so different solid phases.

Gas phase has highest volume and greatest potential energy, because average distance between molecules is greatest. Gases have high chemical potential. High temperature makes material gas, because increase in entropy is higher than increase in potential energy. Gas phase has least order, because all molecules are independent, with no physical structure, and temperature is highest. Gas phase has greatest randomness and highest entropy.

Compared to solid, liquid phase has more volume and more potential energy, as distance between molecules becomes more. Intermediate temperature makes material liquid, because entropy change is similar to potential energy change. Liquid phase has less order than solid phase, because crystal structure breaks down into fluid structure, and temperature is more.

Temperature, pressure, phase number, and substance amounts, concentrations, and pressures affect chemical systems.

Some factors relate to others, and some are independent. Available phases are independent, because substances can go to all phases.

Temperature can freely vary, must be the same throughout system, and does not depend on other factors.

Pressure can freely vary, must be the same throughout system, and does not depend on other factors.

Number of independent components is number of substances minus one. Because total percentage must be 100%, because sum of mole fractions must equal one, one substance's percentage is dependent.

With no equilibria in system, number of independent factors is $(c - 1) * p + 2$, where c is component number, and p is phase number.

Substance amounts in two phases in contact at phase boundary typically are in equilibrium. Chemical potentials of both phases must be equal. Number of equilibria is $c * (p - 1)$, where c is component number, and p is phase number.

Factor number and equilibria number determine how many factors {degrees of freedom, equilibrium} can freely change in chemical systems. Degrees of freedom equal free-variable number v minus equilibria number e {phase rule, components}: $v - e$. Degrees of freedom total: $c - p + 2 = 1 + 1 + (c - 1) * p - c * (p - 1)$, where c is component number, and p is phase number.

If chemical systems have ions, electrical neutrality requires adding one more equilibrium to chemical system.

Knowing chemical-system initial state fixes one more factor and adds one more equilibrium.

If force field is present in chemical system, it adds one more phase to chemical system.

Two immiscible substances can make two heterogeneous phases, even if both are liquids, because they have boundary. Both components are in equilibrium. Heterogeneous phase is mixture and has several components, several phases, and several equilibria. Two miscible substances can combine to make only one homogeneous phase, with no boundary, because they mix and so are not in equilibrium. Homogeneous phase counts as one component and one phase.

Physical systems tend to go to lowest potential energy and greatest entropy. Electrical forces push or pull atoms and change potential energy to kinetic energy. Friction opposes electrical forces, so kinetic energy tends to become heat. Physical systems tend toward most randomness and lowest physical order and so greatest entropy. Number of molecules freed from order relates to average random kinetic energy and so temperature.

Total available energy is free energy and is energy from order breakdown plus potential energy. Physical systems tend to go to lowest free energy. Lowest free energy is optimum between lowering potential energy and raising entropy.

System has same average random translational kinetic energy throughout. Temperature stays constant during phase change from gas to liquid, or liquid to solid, because all heat energy removed is potential energy, which kept molecules apart, not kinetic energy. Temperature stays constant during phase change from solid to liquid or from liquid to gas, because added kinetic energy from heat becomes potential energy that makes molecules farther apart.

phase rule

Degrees of freedom d are free-variable number n minus equilibria number e: $d = n - e$.

critical temperature

Substance solid, liquid, and gas phases can all be present at triple point at high temperature and high pressure.

criticality - At critical temperature and pressure, substance solid, liquid, and gas phases are all present {triple point}. At high pressure and temperature, gas and liquid can be one phase, dense gas or gaseous liquid.

opalescence - Vapor at critical point has white glow.

equilibrium of phase

Two phases can have no net flow between them. At boiling-point or melting-point temperature, both phases, liquid/gas or solid/liquid, respectively, have same free energy and are in equilibrium. If temperature decreases, substance becomes denser, because entropy change is small but potential-energy change is large. If temperature increases, substance becomes less dense, because entropy change is large but potential-energy change is small.

fluids

Gases and liquids are similar in flow properties. Van der Waals forces between molecules can cause cohesion, adhesion, adsorption, and surface tension. More polarized materials have increased molecule van-der-Waals forces, so cohesion and surface tension are greater, and adhesion is smaller.

Brownian motion

Particles in suspension move randomly in all directions with many velocities. Brownian motion depends on time.

Fluid-molecule collisions cause fluid random motions and microscopic-particle oscillations. Particles travel short average distance, through mean free path, before next collision. Collision frequency varies inversely with mean free path.

Telephone errors have bursts, with random intervals between errors. Brownian-motion zerosets, or random Cantor sets with fractal dimension between zero and one, can model them.

Processes {random walk} can take same-length steps in all directions. Average distance from origin is square root of step number. Return to origin is probable. Random walk can be along a line or have more dimensions.

Brownian motion is neither fractal nor self-similar.

diffusion

If substance releases into another substance, released substance tends to flow throughout other substance, until released substance evenly distributes.

Random molecule elastic collisions cause diffusion. Diffusion decreases system order. Gases tend to expand as molecules collide. Expansion rate depends on molecule velocity and size and on barriers to motion.

In media, introduced foreign substance diffuses outward over time depending on number {spectral dimension} of medium-particle nearest neighbors. For ink in water, ink volume V increases as time t to 3/2 power: V = t^1.5.

effusion

Molecular collisions cause gas to flow through container holes.

fractionation

Different pure fluids can come from fluid mixture by controlling pressure, vaporizing mixture at increasing temperatures, and cooling and condensing each vapor.

solids

Matter can be firm, be incompressible, have local and global order, have no flow, and have no expansion to fill volume. Most solids are crystals. Crystals can be cubes, prisms, rhomboids, parallelepipeds, hexagons, or diamonds. Elements {molecular elemental solid} can have molecules bound by weak van der Waals forces. Metal elements {metallic elemental solid} have atoms bound with metallic bonds. Non-metal elements {non-metallic network elemental solid} have each atom covalently bound to four or less atoms, in covalent-bond crystal lattices. Salts {ionic solid} can have ionic bonds between ions. Salts {polar solid} can have dipole-dipole bonds between polar molecules.

Solid molecules vibrate and rotate but do not translate to new positions. Solids can transmit pressure in force direction but cannot cause pressure.

adsorption to surface

Substances can bind onto solid surfaces.

conductivity

Materials have ability to conduct electricity.

Conductor molecules can have half-filled electron energy level, so electrons can jump to same energy level in neighboring molecules. Metallic crystals are conductors.

Semiconductor molecules can have energy level filled with electrons and slightly higher empty energy level, so electrons can jump into neighboring molecules at normal temperatures. Covalently bonded crystals are semiconductors.

Insulator molecules can have energy level filled with electrons and much higher empty energy level, so electrons cannot jump into neighboring molecules at normal temperatures. Ionic crystals and hydrogen-bonded crystals like ice, hydrogen fluoride, and proteins are insulators. Insulators are typically transparent.

Dulong and Petit law

For most solids, heat capacity per mole is constant.

amorphous solid

Solids can have no regular structure, such as glass.

crystals

Almost all solids have regular molecule, ion, or atom arrays.

Crystals {cubic crystal} can have eight atoms around small atom, three perpendicular four-fold same-length axes, and five crystal classes.

Crystals {rhomboid crystal} can have twelve atoms around similar size atom with every third layer directly above another.

Crystals {rhombohedral crystal} can have one three-fold axis, with one axis perpendicular to the other two but with different length, and two axes with same length at 120-degree angle to perpendicular axis, and make five crystal classes.

Crystals {hexagonal crystal} {hexagon crystal} can have twelve atoms around similar size atom with alternate layers directly above each other.

Crystals {tetrahedral crystal} {tetrahedron crystal} can have four large ions around each small ion.

Crystals {octahedral crystal} {octahedron crystal} can have six large ions around each small ion.

Crystals {triangular crystal} can have three large ions around each small ion.

Crystals {planar crystal} can have two large ions around each small ion.

Only one, two, three, four, or six rotational symmetries can fill all space with no gaps or overlaps, so only seven crystal types are common.

The more similar in size atoms or ions are, the more atoms or ions can surround one atom or ion. Number of atoms or ions surrounding one atom also depends on ion charge or covalent-bond number.

Crystals have lattice structure. Including crystal type and lattice type, 32 crystal classes exist. 14 unit cell and 32 crystal class translations and transformations make 234 possible crystal shapes.

Crystals grow at dislocations, because binding molecules can contact two surface atoms. Impurities, long bond lengths during fast growth, and screw dislocations can cause dislocations and irregularities. Small crystal faces grow fastest by deposition. Large crystal faces can adsorb other materials.Crystal surfaces are never flat but are lumpy. Perfect crystals cannot grow.

Miller index

Diagrams can show crystal planes, using three perpendicular axes. Crystal vertices are one unit length or less apart. Coordinate reciprocals indicate planes through vertices. Putting origin at one vertex and using coordinates for other vertices indicates edges. If plane is parallel to axis, coordinate reciprocal is zero.

dendritic growth

Crystal growth can be along one axis, because heat leaves best at tips, so deposition is easiest there.

glide plane

Unit cells can undergo translation and reflection.

screw axis

Unit cells can undergo translation and turn around axis 0, 180, 120, 90, 72, or 60 degrees.

clathrate

Crystals can be so open that they can hold small molecules inside, without bonding. Examples are very-cold-water forms and very-cold methane and hydrogen gas mixtures.

metal crystal

Metal crystals can have hexagonal close packing, face-centered cubic close packing, or body-centered cubic close packing.

polymorphic solid

Substances can have more than one crystal form.

liquid crystal

Crystals can have regularity in only one or two dimensions, allowing unit cells to slide past each other in third dimension. Liquid crystals are anisotropic, flow in sheets or steps, and are asymmetric molecules. Numbers can display electronically using reflected or transmitted light, as electric field makes crystal tinted, and no electric field makes crystal transparent.

nematic crystal - Liquid crystals can have linear crystals with regularity in only one dimension, are like threads with no planes, orient, and are not periodic.

smectic crystal - Liquid crystals can have planar crystals with regularity in two dimensions, orient, and are not periodic.

crystal defect

Missing atoms, extra atoms, or different atom types alter regular crystal structure.

dislocation in crystal

Crystal defects can displace unit cells from usual positions. Inserted atoms wedged into lattice can cause dislocations {edge dislocation}. Dislocations {screw dislocation} can be around axis to make helical unit-cell arrangements. Dislocations in crystals affect brittleness, ductility, and other mechanical crystal properties. Alloys lack dislocations and so do not slide, because odd atoms move to lowest free-energy positions.

Frenkel defect

Crystal ions can move to interstitial places and so leave vacancies.

impurity

Different molecules or atoms can be in crystals.

interstitial ion

Extra ions can be in ionic crystals.

Schottky defect

In ionic crystal, cations and anions can be missing.

vacancy

Ions can be missing from ionic crystals.

unit cell

Crystals have repeating atom groups. Crystals have lattice structure.

Bravais lattice

Only 14 possible arrangements of identical spheres can make unit cells {space lattice}. Three are cubic, two monoclinic, four orthorhombic, two tetragonal, one triclinic, one hexagonal, and one rhombohedral.

primitive lattice

Lattices can have atoms at unit-cell corners.

body-centered lattice

Lattices can have one or two atoms at unit-cell centers. Lattices {body-centered cubic close packing} can have atoms in cube centers, with identical atoms at cube corners.

face-centered lattice

Lattices can have one atom in unit-cell face. Unit cells with different atoms {face-centered cubic close packing} can have atoms in cube centers and in cubic-unit-cell face centers.

cubic close packing

Unit-cells can be cubic. Twelve identical atoms surround each atom, and every third layer is directly above another.

hexagonal close packing

Unit cells can be hexagonal. Twelve identical atoms surround each atom, and alternate layers are directly above each other.

symmetry in crystal

Unit crystals can have same structure after rotation around axis, reflection across axis, inversion through central point, translation along axis, or any combination. Nature has six symmetry groups: isometric, hexagonal, tetragonal, orthorhombic, monoclinic, and triclinic.

hexagonal crystal

Symmetry groups can have rotation by 60 degrees. Hexagonal crystals have one six-fold axis, with one axis perpendicular to the other two axes but with different length, and two axes with same length at 60-degree angle to perpendicular axis, and makes seven crystal classes.

isometric crystal

Symmetry groups can have rotation by 90 degrees, reflection, and inversion.

monoclinic crystal

Symmetry groups can have rotation by 90 degrees. Crystals can have one two-fold axis, two perpendicular same-length axes, and one non-perpendicular different-length axis, to make three crystal classes.

orthorhombic crystal

Symmetry groups can have rotation by 180 degrees and reflection. Crystals can have three two-fold axes, which are all perpendicular but have different lengths, to make three crystal classes.

tetragonal crystal

Symmetry groups can have rotation by 90 degrees. Crystals can have one four-fold axis and three perpendicular axes but only two with same length, to make seven crystal classes.

triclinic crystal

Symmetry groups can have rotation by 120 degrees. Crystals can have three axes, all not perpendicular but all of same length, to make two crystal classes.

coordination number

In inorganic ionic crystals, a number of other atoms surround each atom. Eight opposite-charge atoms surround each atom, if atoms have equal size. Six opposite-charge atoms surround each atom, if atoms have unequal size. Four opposite-charge atoms surround each atom, if atoms have very unequal size.

zone melting

Melting solids along a moving band can make impurities remain in melted part and pure part solidify behind moving band.

gases

Matter can have low density, flow, be compressible, have no local or global order, and expand to fill volume.

Gases have high kinetic energy. Molecules have speed 500 meters per second. Electric forces between molecules have little effect.

Molecules are 20 nanometers apart. Gases have distances between molecules that are five to ten times more than distances in solids and liquids. Gases are like spaces of vacancies with some positions filled.

Gas-molecule diameter is 0.5 nanometers.

If gas volume increases while temperature stays the same, pressure decreases, and entropy and potential energy increase. If gas temperature increases while pressure stays the same, volume increases, and entropy and potential energy increase. Temperature and pressure depend on kinetic energy, and volume and entropy depend on potential energy. If total energy is constant, P/V = S/T, where P is pressure change, V is volume change, S is entropy change, and T is temperature change. Potential energy increases when entropy increases, and entropy increases when potential energy increases.

gas law

Work that gas can do equals heat energy in gas: P*V = n*R*T, where P is pressure, V is volume, n is moles, R is gas constant, and T is absolute temperature. Pressure times gas volume is work that gas can do. Product of gas moles and absolute temperature and gas constant is gas heat energy.

Volume {molar volume} of one mole of ideal gas at standard temperature of 25 C and standard pressure of one atmosphere is 22.4 liters.

Ideal gas law assumes that molecules have elastic collisions, have no volume, and have no forces among them. At low pressure, real gas has less pressure than ideal gas, because molecules attract each other. At high pressure, real gas has higher pressure than ideal gas, because molecules repulse each other.

Ideal-gas-law modifications {van der Waals equation} {virial equation} account for molecule sizes and interactions.

Graham law

Gas diffusion rate is inversely proportional to gas-density square root.

Joule-Thompson inversion

Expanding gas can cool if temperature is below maximum temperature.

liquids

Matter can be dense, flow, have local order several atoms wide, have long-range disorder between local-order regions, have no expansion to fill volume, have less than 3% compression, have translational kinetic energy, and have molecule interactions.

Liquids can be molten salts, metallic liquids, hydrogen-bonded liquids, or van der Waals bonded liquids.

Cohesive forces can be ionic, dipoles, ions and dipoles, dipoles and induced dipoles, or London forces between neutral atoms. Internal forces can be repulsive if compression is high.

Melting is similar to reaching elastic limit in solid under stress.

plasma

plasma phase - At high temperature, materials can exist as ionized gas, as substance vaporizes and loses electrons.

Wigner solid - High-density proton plasmas can act like liquids.

mixture

Substances can mix with other substances. Mixed-substance chemical potential is less than pure-substance chemical potential, because mixtures have more disorder.

aerosol suspension - Liquid can be suspension in gas, like dust and fog.

emulsion of fluid - Liquid can be suspension in liquid, like mayonnaise, cheese, and shaken salad dressing.

sol mixture - Liquid can be colloid in another liquid, like india ink.

foam - Gas can be suspension in liquid, like whipped cream and foam rubber.

gel - Solid can be colloid in liquid, like jelly.

homogeneous phase - Mixing two substances can make one phase or solution, with no boundaries.

heterogeneous phase - Two substances can mix but have boundaries between different regions.

colloid

Mixtures with particle diameters 1 to 100 nanometers are translucent or opaque, separable by fine membranes, and settle slowly. Solutions can have many small particles, which attract an opposite-charge ion layer, which then attract an ion layer. Layers prevent good precipitation. Hydrophilic colloids are viscous, hard to coagulate, and gel-like. Hydrophobic colloids are sols, make curds, and are easy to coagulate.

suspension mixture - Mixtures can have particles with diameter greater than 100 nanometers, be translucent or opaque, be separable by coarse membranes, and settle quickly.

osmosis

Membrane allows solvent molecules to pass but not solute molecules. If membrane separates solution from another solution, solvent passes into solution with higher solute concentration, because more solvent molecules hit membrane on side with less solute and pass through to other side. For example, solvent can be water, with many solute molecules inside membrane bag.

Osmosis increases solvent amount on membrane side with higher solute concentration, causing extra pressure on that membrane side. The osmotic pressure resists further osmosis, because number of solvent molecules hitting both membrane sides becomes equal. For example, water passes into bag, making bag bigger and stretching it. Membrane is under pressure. The extra water inside causes higher pressure inside, meaning more water molecules hit membrane inside.

Mixtures have higher chemical potential than pure liquids, so pure liquid goes from pure-liquid membrane side to mixture membrane side, raises liquid level on mixture side, and lowers chemical potential. Chemical-potential decrease generates osmotic pressure, which tries to bring system into equilibrium.

Membrane can allow small solute molecules and ions to pass through, but not large solute molecules. Small molecules diffuse through membrane, tending to make small-solute molecule concentrations equal on both membrane sides.

osmotic pressure - Osmosis increases solvent amount on membrane side with higher concentration, causing extra pressure {osmotic pressure}| on membrane from that side. The extra pressure resists further osmosis, as number of solvent molecules hitting both sides becomes same.

solution

Solvent molecules can surround solute molecules, to make one phase. Liquid solutions are transparent, because they are one phase with no surfaces for light reflection. To find solution substance concentration, divide substance moles by volume in liters of solution, not just solvent.

immiscible liquids - Two liquids can be unable to dissolve in each other, like benzene and water.

miscible liquids - Two liquids can dissolve in each other, like alcohol and water.

solubility

At temperature, a number of solute grams can dissolve in 100 milliliters of solvent. If solubility is greater than one percent, solute can dissolve well in solvent {soluble}.

Compounds have solubility in water {aqueous solubility}. Nitrates, acetates, chlorides, bromides, most iodides, most sulfates, sodium salts, potassium salts, and ammonium salts are soluble in water. Hydroxides except sodium hydroxide and potassium hydroxide; sulfides except sodium sulfide, magnesium sulfide, and aluminum sulfide; arsenates except sodium arsenate and potassium arsenate; carbonates except sodium carbonate and potassium carbonate; and phosphates except sodium phosphate and potassium phosphate are insoluble in water.

Solubility is maximum concentration before precipitation from solvent. For precipitation, concentration product {solubility product} must be greater than equilibrium constant.

Solute solubility in solvent depends on solute polarity, size, and surroundableness and solvent polarity, size, and surroundability.

Higher temperature increases solubility, because increased random motion breaks up and mixes solute and solvent more.

Low concentration increases solubility, because solvent molecules can better surround solute molecules, with less solute molecules near other solute molecules.

Stirring increases solubility, because more motion mixes solute and solvent more.

More hydrogen ion increases solubility in polar solvent, by increasing polarity. More ions increase solubility in polar solvent, by increasing polarity. Ionic solutes with large ions and large charges are harder to dissolve. Ionic solutes with small ions and charges of +1 or -1 are easier to dissolve. Large ions with charge +1 or -1 dissolve better than small ions with charge +2, -2, +3, -3, or greater. Salts with small volume, especially hydrogen ions and acids, increase solubility by allowing more shielding. Higher-charge salts increase solubility more, because they shield better. Salts that form metal-ion complexes increase solubility, by more shielding.

Low concentration of ion common to two solutes increases solubility, because solvent molecules can better surround solute molecules, with less solute molecules near other solute molecules. Common ion provides more molecules for collision. Solubility decreases if common ion is present in large amounts, because the other ion must then be at low concentration to equal solubility product.

Solubility increases if diverse ion is present, because charges have more shielding and polarity increases.

Similar-polarity molecules dissolve each other best, because electrical attractions for similar molecules are stronger.

Small molecules dissolve better, because solvent molecules can better surround solute molecules.

Spherical molecules dissolve better than elongated ones, because solvent molecules can better surround spherical molecules.

hydration energy

When water molecules surround ions, energy releases. Small atoms have more hydration than large ones, because water surrounds them better.

solute

Solvent molecules can surround other-substance molecules.

common ion effect

Solubility can decrease by adding salt with common ion and higher solubility, to make higher concentration and force solute out of solution.

Debye-Huckel law

Ions in solution increase polarity and increase solubility.

solvent of solution

Substance molecules can surround solute molecules.

aqueous solution - Solvent is usually water.

tincture - Solvent can be alcohol.

alloy

Solid can dissolve in another solid, as in steel, bronze, and brass. Steel is carbon in iron. Bronze is tin in copper. Brass is zinc in copper.

amalgam - Liquid can dissolve solid, as metals can dissolve in mercury.

concentration of solution

Solutions have ratio of solute mass or volume to total solution or solvent mass or volume.

molal - Concentration can measure number of solute moles dissolved in one solution kilogram.

molar concentration - Concentration can measure number of solute moles dissolved in one solution liter.

mole fraction - Concentration can measure number of solute moles dissolved in one solution mole.

normal concentration - Concentration can measure number of solute-ion equivalents dissolved in one solution liter.

parts per million - Concentration can measure number of solute milligrams dissolved in one solution liter.

percent solution - Concentration can measure number of solute grams or milliliters dissolved in 100 solution grams or milliliters.

concentration amount

dilute solution - Solutions can have low concentration.

concentrated solution - Solutions can have high concentration.

saturated solution - Solutions can have maximum concentration, at temperature, if dissolved solute is in contact with solid solute.

supersaturated - At a temperature, if no solid is present and solution has no crystallization, solution can have concentration higher than saturated concentration.

polarity

Solvents can have electronegative atoms.

Polar solvents include water, ethyl alcohol, and methyl alcohol. Acetone is slightly polar. Benzene is non-polar.

Like dissolves like. Polar solvents can dissolve in each other. Non-polar solvents can dissolve in each other. Sticky non-polar solids are harder to dissolve in non-polar solvents, because they do not break up. Acetone, methyl alcohol, ethyl ether, and ethyl alcohol are soluble in water. Benzene, carbon tetrachloride, chloroform, hexane, cyclohexane, methylene chloride, toluene, and xylene are insoluble in water.

electrolyte - Solid can dissolve in water to make ion solutes.

partial molar quantity

Material true concentration or partial pressure depends on other-substance concentrations or partial pressures, because having different substances contributes more disorder to system. Substances interact, because system has total pressure, temperature, and concentration.

chemical potential

Material true concentration or partial pressure depends on partial molar quantity, because having different substances contributes more disorder to system. Substances interact, because system has total pressure, temperature, and concentration that distribute among substances. Substance partial molar Gibbs free energy is partial derivative of free energy with substance moles, if temperature, pressure, and other-substance amounts are constant.

Henry law

Gas solubility in liquid is proportional to partial pressure of gas in contact with liquid.

partial pressure law

Gases in mixtures independently contribute pressure {partial pressure} to total gas pressure.

Raoult law

Solute-vapor partial pressure above solution equals solute mole fraction times pure-solute vapor pressure.

colligative property

Pure substances have molar quantities, but mixtures have partial molar quantities, which depend on material moles divided by total moles, the mole fraction. Solution properties, such as partial pressures, boiling point elevation, freezing point depression, osmotic pressure, solubility, volatility, and surface tension, can depend on solute mole fraction. Partial molar quantities are interdependent, because mole fraction total must be one.

boiling point elevation

Solutions have higher boiling point than pure solvent, because solute molecules are heavier than solvent molecules and have lower volatility. If liquid includes impurities that are less volatile than liquid, liquid boils only at higher temperature. Salt in water raises boiling temperature. Mixture boiling point is higher, because mixtures are more random, so difference between liquid and gas is less. Immiscible substances lower boiling point, because both vapor pressures add to increase pressure. In boiling-point elevation, temperature change dT equals constant k times molality M: $dT = K*M$. People know constants for solutes and solvents.

freezing point depression

Solutions have lower freezing point than pure solvent, because solute molecules are impurities in solvent crystals and so make crystals harder to form. In freezing-point depression, temperature change dT equals constant k times molality M: $dT = K*M$. People know constants for solutes and solvents.

precipitation of solute

Solid solute molecules can crystallize and separate from liquid solvent molecules.

Decreasing solubility causes precipitation. Solubility decreases by cooling. Solubility decreases by adding organic non-polar solvents, such as acetone and ethanol, to water solution.

Solubility decreases by neutralizing solution to reduce acidity and polarity. Adding concentrated ammonium sulfate usually causes precipitation from aqueous solution. Molecules precipitate best at isoelectric points, because they have least polarization there.

Precipitates {lyophobic}, like sulfur and metal salts, can be small pellet-like precipitates, with molecules that reject water and adsorb ions. Precipitates {lyophilic}, like starch and gelatin, can be large curd-like precipitates, with molecules that adsorb water.

fractional crystallization

Crystals can precipitate from solution, by evaporating solvent, by cooling solution, or by adding another solvent to solution.

PHYS>Chemistry>General>Chemical Tests

chemical tests

- Color.
- Odor.
- For state, find solid crystal group, liquid viscosity, or gas nature.
- melting or boiling point.
- For solubility, check in water, organic solvent, base, bicarbonate, hydrochloric acid, and sulfuric acid, which protonates O, N, and S.
- For combustion, use Beilstein test for halides, ignition test for highly unsaturated or aromatic organic chemicals, and flame test for metals.
- Chemical tests:
- Sodium iodide in acetone detects halides.
- Ferrous hydroxide detects nitro- groups.
- Bayer test or bromine in carbon tetrachloride detects double bonds.
- Tollen's test detects aldehydes.
- Iodoform, dinitrophenylhydrazine, and chromic-acid tests detect aldehydes and ketones.
- Sodium bicarbonate, silver nitrate, and neutralization with base, on pH paper or in meter, detect carboxylic acid.
- Sodium hydroxide, ferric chloride, and bromine water detect phenols.
- Hinsberg test and nitrous acid detect amines.
- Acetyl chloride, Lucas test, and chromic acid detect alcohols.
- Ferric hydroamate or hydrolysis with base detects esters.
- Spectroscopy detects cyano- groups.
- Infrared spectroscopy detects chemical bonds.
- Ultraviolet-visible spectroscopy detects aromatic chemical groups.
- Nuclear magnetic resonance (NMR) detects electron densities.
- Mass spectroscopy detects elements.