SAT Chemistry

Atomic structure and periodicity

- Dalton's theory of atoms:
  - matter made up of small discrete atoms
  - atoms of the same element weight the same
  - cannot be subdivided, created or destroyed
  - they combine in simple ratios to form compounds
  - in chemical reactions, they are combined, separated or rearranged.
- Robert Milikan determined electron mass
- Chadwick discovered neutrons, whose existence as predicted by Rutherford
- Bohr proposed a quantised planetary model:
  - Max number of electrons in an energy level is $2n^2$.
- Henry Mosely determined atomic numbers with X-ray diffraction
- Electrons move discretely and absorb or emit quanta:
  - drop to n=1 - Lyman (UV)
  - drop to n=2 - Balmer (visible)
  - drop to n=3 - Paschen (IR)
- Light spectroscopy:
  - prism or diffraction grating disperses light into spectra
  - spectral lines are unique to an element, can be found with $E = hf$
- Mass spectroscopy separates isotopes based on mass difference: acceleration, deflection, detection

Quantum model

- Louis de Broglie extended Planck's wave particle duality to all matter
- Heisenberg and Schrödinger: electrons exist in orbital clouds
- Electrons can be described by four quantum numbers:
  1. Principal quantum number (n):
     - main energy levels
     - n = 1,2,3,4…
  2. Angular momentum /azimuthal (l)
     - orbital shape
     - l = 0,1,2 …, n-1
     - l = 0: s orbital
     - l = 1: p orbital
     - l = 2: d orbital
  3. Magnetic quantum number (m_l)
     - number of orientations of an orbital
     - s =1 , p =3, d = 5, f = 7… (the odd numbers)
  4. Spin quantum number (m_s)
     - m_s = +/- 1/2
     - each orbital can be filled with two electrons of opposite spins (Hund's rule)
- Pauli exclusion principle: in a given atom, no two electrons can have the same set of quantum numbers
- Aufbau principle: each electron occupies the lowest orbital
- Nodes are the regions of zero probability

Periodicity

- Chromium: [Ar] 4s1 3d5.
- Copper: [Ar] 4s1 3d10.
- Lowest electronegativity = highest desire to lose electron

Bonding and Chemical Formulae

- If E.N > 1.7, ionic.
- If 0.4 < E.N < 1.6, then polar covalent
- When a covalent bond forms, there is minimum potential energy between atoms
- Hybridisation and VSEPR are two theories to predict shapes.
- Molecules with eight electron pairs have sp3d2 hybridisation.
- For ions with oxygen:
  - hypo- = one oxygen (or fewest number of oxygens)
  - -ite = fewer oxygens
  - -ate = more oxygens
  - per- = most oxygens
- For naming acids:
  - If there is no oxygen, then it is hydro-
  - If the acid anion is an -ate, then the acid will be -ic e.g sulphuric acid
  - If the acid anion is an -ite, then it will be -ous e.g sulphurous acid
- An **acid anhydride** is a nonmetal oxide which reacts with water to form an acidic solution
- Law of Multiple Proportions - two elements can sometimes combine in different small-whole-number ratios
- The formal charge formula is: $\text{Formal charge} = V - \frac{1}{2}N_{bonding} - N_{nonbonding}$
- where V is the number of valence electrons.

Gases and the gas laws

- To prepare oxygen, we heat potassium chlorate with a manganese dioxide catalyst:

  $2KClO_3 \rightarrow 2KCl + 3O_2$
- Hydrogen is normally produced by electrolysis, passing steam over coke, or cracking hydrocarbons
- 1 atm = 760mmHg = 760 torr = 101.325kPa
- Effusion - passage of gas through small hole into empty chamber.

Gas laws

- Graham's Law of Effusion:
  - The rate of effusion: $rate \propto \sqrt{M_R}$
- Gay-Lussac's law = pressure law.
- Dalton's law of partial pressures:
  - When a gas is made up of a mixture of different gases, the total pressure is the sum of the partial pressures, where the partial pressure is the pressure that the gas would make if it were alone:

$$P_{total} = P_{gas1} + P_{gas2} + P_{gas3} + ...$$

- When a gas is collected over a volatile liquid, you must subtract the vapour pressure: $P_{gas} = P_{atm} - P_{H_2O}$
-
- Since mercury is 13.6 times heavier than water, the height of mercury in a eudiometer will be 1/13.6 the height of water.

- For the Ideal Gas Law, $R = 0.0821\frac{L \cdot atm}{mol \cdot k}$
- density: $\rho = \frac{PM}{RT}$
- At high pressures or low temperatures, molecules get closer so attractive forces come into play - hence volume is lower than ideal gas law predict
- Gay-Lussac's law (another one):
  - the ratio between the volumes of the reactant gases and the products can be expressed in simple whole numbers.

Liquids, solids and phase changes

- Above the critical temperature, the liquid phase of a substance cannot exist.
- At critical temperature, critical temperature is required to liquefy the gas.
- Amorphous solids have a random structure
- Polycrystalline solids have a large number of crystals arranged randomly.
- For solids, solubility increases with temperature (opposite in gases)
- The solubility of gas in liquid is directly proportional to the gas pressure (Henry's law).
- Water mixtures can be solutions (particles smaller than 1nm), suspensions (greater than 1000nm) or colloids (in between).

Table 9. Solubilities of Compounds

Soluble	Except
Na^+, NH_4^+, K^+ compounds	
Acetates	
Bicarbonates	
Chlorates	
Chlorides	Ag, Hg, Pb ($PbCl_2$, sol. in hot water)
Nitrates	
Sulfates	Ba, Ca (slight), Pb
INSOLUBLE	
Carbonates, phosphates	Na, NH_4, K compounds
Sulfides, hydroxides	Na, NH_4, K, Ba, Ca

- Saturated solutions are in equilibrium.
- Molality is the number of moles per 1000 grams of solvent.
- The freezing point of a substance decreases 1.86K per molality of solute per mole of solute particles, the boiling point increases by 0.51K.

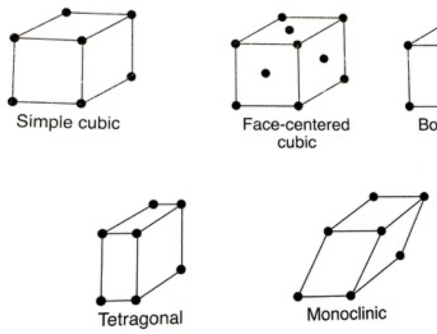


Figure 33. Kinds of Unit Cells

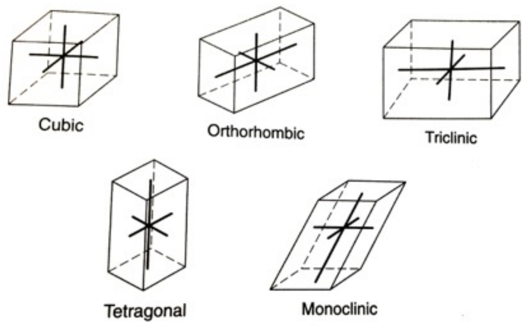


Figure 34. Crystal Structure Classified by Internal Axis

- Hydrates that lose water of hydration are efflorescent
- Hydrates that absorb water are deliquescent or hydroscopic e.g calcium chloride.
- Osmotic pressure is proportional to molarity (M). $\prod$=MRT.

Chemical reactions and equilibria

- For double replacement reactions to go to completion, one of the following must be formed
  - insoluble precipitate

- nonionising substance
- gaseous product given off

- Since energy is conserved (First law of thermodynamics), we can add enthalpies e.g in Hess's law.
- The overall change in entropy is positive. $\Delta G = \Delta H - T\Delta S$

- The Law of Mass Action states that for a reaction between A and B, aA+bB →…, the rate expression is:

$r \propto [A]^a[B]^b$ or $r = k[A]^a[B]^b$ where k is the specific rate constant.

Equilibria

- A saturated solution of a substance has an equilibrium e.g $AgCl \rightleftharpoons Ag^+ + Cl^-$
- The solubility constant:

$K_{sp} = [Ag^+][Cl^-]$

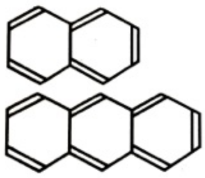
$$\Delta G^\circ = -nFE_{cell} = RT\ln K_{eq}$$

- If predicted concentrations are too large, precipitates form
- The common ion effect states that a large increase in ions on one side of an equilibrium may cause some substance to move out of solution
- The equilibrium constant K depends on the Gibbs free energy: $k = e^{-\Delta G^\circ/RT}$

Acids and bases

- Conjugate base is an acid minus H. Strong acid = weak conjugate base.
- pH +pOH = 14
- Buffers are weak acids and their conjugate base.
- Metallic oxides can react with nonmetallic oxides.
- The **normality** of a solution regards the number of number of mole equivalents of H+.
- The gram equivalent weight is how much of the acid solution is needed to release the equivalent H+

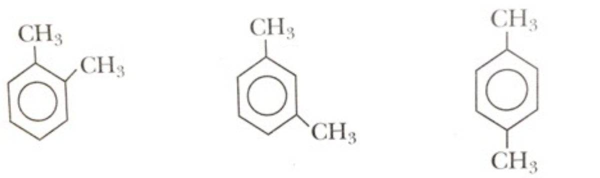
Organic chemistry



$C_{10}H_8$ naphthalene

$C_{14}H_{10}$ anthracene

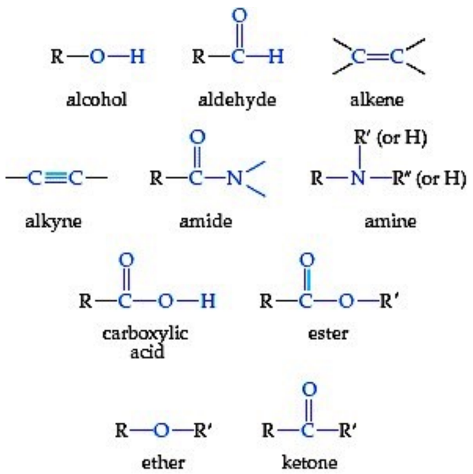
- For cycloalkanes, assign position 1 to the alkyl that comes first in alphabetical order then give the rest the lowest numbers possible.
- The aromatic groups are unsaturated ring structures, the general formula is C_6H_{2n-6}. The simplest is benzene (C6H6).
  - The phenyls are C6H5, groups can then be added, e.g methyl benzene = toulene.
  - Different benzene substitutions have different isomer forms, ortho, meta and para.



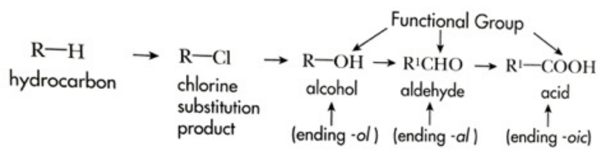
ortho- structure 1,2-dimethylbenzene or *o*-xylene

meta- structure 1,3-dimethylbenzene or *m*-xylene

para- structure 1,4-dimethylbenzene or *p*-xylene

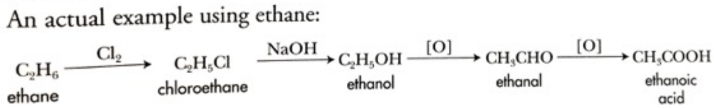


Summary of Oxygen Derivatives



Note: R^1 indicates a hydrocarbon chain different from R by having one less carbon in the chain.

An actual example using ethane:



Practical chemistry

- Some precise tools are:
  - Gravimetric balances with readings to thousandths of a gram
  - pH meters
  - Spectrophotometers (percentage of light transmitted through a substance - determines molality)

Tests for substances

- Ammonia - expose to HCl fumes - NH4CL (white solid) forms
- HCl - exhale over gas, vapour fumes
- Hydrogen sulfide turns led acetate paper brown black
- Oxygen - add nitric oxide gas, turns red-brown
- Acetate - add sulphuric acid and warm - vinegar odour.
- Sulphate - add BaCl2 then HCL - white precipitate forms
- Ammonium - add strong base and heat - ammonia released
- Fe 2/3, potassium ferricyanide/ferrocyanide then dark blue precipitate
- Flame tests:
  - sodium - yelow
  - Potassium - violet
  - Lithium - crimson
  - calcium - orange-red
  - barium - green
  - strontium - bright red
- Hydrogen sulfide tests:
  - lead - brown black
  - copper -black
  - silver - black
  - mercury - black
  - nickel - black
  - iron - black
  - cadmium - yellow
  - arsenic - yellow
  - antimony - orange
  - zinc - white
  - bismuth -brown.