Nomenclature of inorganic compounds
Ionic equations

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Structure of matter:

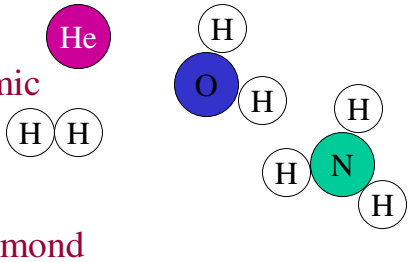
- Atom
- Molecule
- Ion
- Compound
- Mixture (dispersion system)

Atom

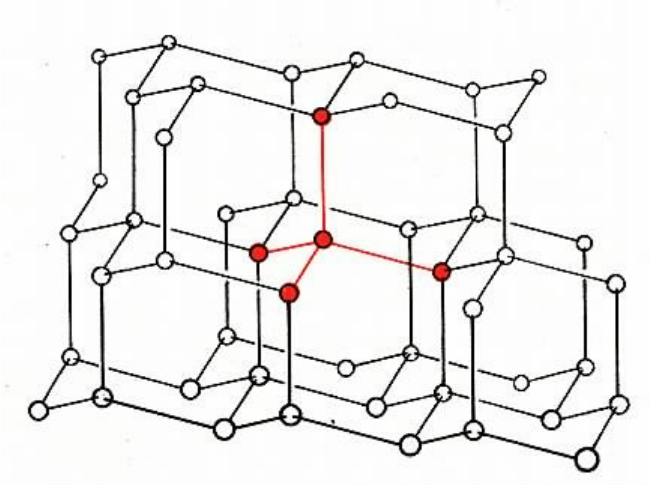
- **Smallest particle of a pure element having its chemical properties**
- **Positively charged nucleus (neutrons, protons)**
- **Electron shell:**
  - **Electron is wave/particle**
  - **Behavior of electron described by quantum mechanics (... wave function, quantum numbers)**
  - **Orbital: space area within the atom shell where occurrence of an electron or pair of electrons is more probable**

Molecule

- smallest particle of a substance having its chemical properties
- Atoms connected via covalent bonds
- Examples:
  - noble gases: monoatomic
  - other gases: diatomic
  - H_2O, NH_3 etc.
  - molecular crystals: diamond
  - ...many thousands of atoms in proteins and nucleic acids



Molecular crystal of diamond

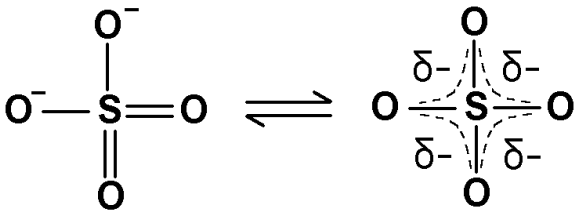


Ion

- atom or molecule with non zero charge (number of electrons does not match number of protons)
- tendency to form ions depends on **electronegativity** of element
- cations (+) or anions (-)
- monoatomic: Na^+, Cl^-, H^+, Fe^{2+}
- molecular: NO_3^- , SO_4^{2-}
- complex: $[Fe(CN)_6]^{4-}$

Molecular ions of oxo-acids:

e.g. sulfate, SO_4^{2-} :



resonance stabilization of sulfate ion

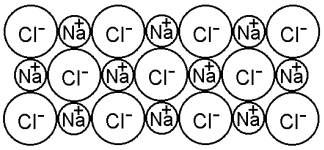
..similar is nitrate NO_3^-, phosphate PO_4^{3-}, carbonate CO_3^{2-}, etc.

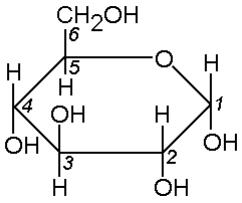
Compound

- Chemically pure substance consisting of the same molecules that containing two or more different atoms
- Atoms are held together in a defined spatial arrangement by chemical bonds
- Independent molecules (e.g. CO_2) or crystalline structures

Chemical formulas

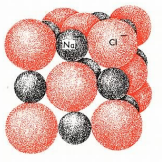
- Stoichiometric (empirical)
  - e.g.: sodium chloride NaCl
  - e.g.: glucose CH_2O
- Molecular
  - e.g.: sodium chloride NaCl
  - e.g.: glucose $C_6H_{12}O_6$
- Structural



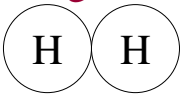


Chemical bond

- Cohesive force holding together atoms in molecules and crystals
- Ionic bond: electrostatic forces among ions having opposite charges



- Covalent bond: sharing a couple of electrons by the two bonding atoms



Polarity of chemical bond

Determined by difference in electronegativity of the two connected atoms:

< 0.4 nonpolar covalent bond

e.g. H-H, carbon-hydrogen

0.4 - 1.7 polar covalent bond

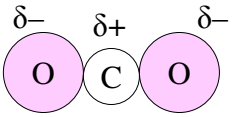
e.g. H-O-H, NH_3, carbon-oxygen, carbon-nitrogen

>1.7 ionic bond

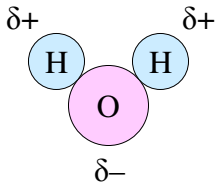
e.g NaCl...

Gradual transition !

Molecular shape affects electric dipole moments:

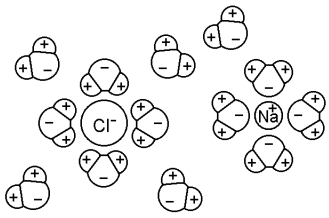


CO_2: linear, nonpolar



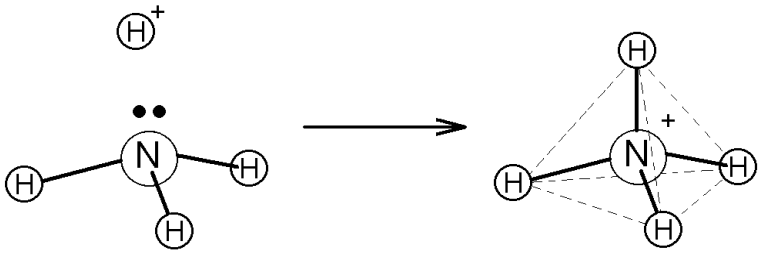
H_2O: angled, polar

... Water as polar solvent:



Coordination covalent bond

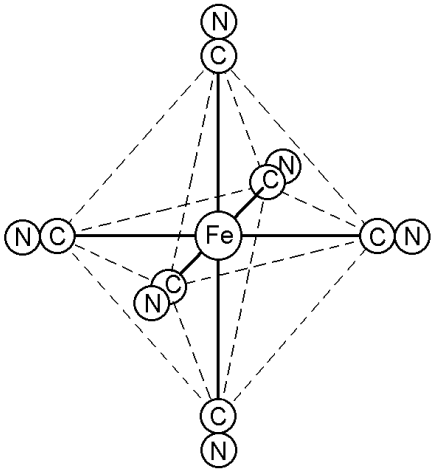
- (also dative, donor-acceptor bond)
- Both bonding electrons provided by one of the atoms (donor), whereas the other atoms provides an empty orbital (akceptor)



Coordination (complex) ions

- central atom of transition metal providing empty orbitals
- ligands providing free electron pairs
- Number of ligands (coordination number) is usually 4 or 6

e.g. $[Fe(CN)_6]^{4-}$



Valence

- Number of covalent bonds formed by an atom
- octet rule: tendency to achieve electron configuration of the nearest noble gas
  - e.g. H-F: H achieves configuration of He, F configuration of Ne
- Therefore O usually bivalent, N trivalent, C tetravalent etc.

Oxidation number (formal valency)

- Oxidation number of element in compound equals its charge after giving all bonding electron pairs to the more electronegative atom
- Can be zero, positive or negative integer
- Basis for nomenclature of inorganic compounds
- Redox reactions: oxidation number increases in oxidation, decreases in reduction

Rules for determination of oxidation numbers

- Free electroneutral atom, or atom in molecule of pure element: oxidation number = 0
- Oxidation number of a monoatomic ion equals its charge
- In heteroatomic compounds the bonding electrons are given to the more electronegative atom, practically:
  - H has nearly always oxidation number I (only in metallic hydrides -I)
  - O almost always -II (only in peroxides -I)
  - F always -I
  - Alkaline metals (Na, K..) always I
  - Alkaline earth elements (Ca, Mg..) always II

Rules for determination of oxidation numbers

Examples:

CO_2 : C^{IV}, O^{-II}

H_2SO_4: H^{I}, S^{VI}, O^{-II}

Sum of oxidation numbers of all atoms in electroneutral molecule is 0, in polyatomic ion equals the ion charge

e.g: CO_3^{2-}: C^{IV}, O^{-II} $\quad 1\times IV + 3 \times(-II) = -2$

Czech nomenclature of oxides:

Oxidation number	Suffix	General formula
I	-ný	X_2O
II	-natý	XO
III	-itý	X_2O_3
IV	-ičitý	XO_2
V	-ečný/-ičný	X_2O_5
VI	-ový	XO_3
VII	-istý	X_2O_7
VIII	-ičelý	XO_4

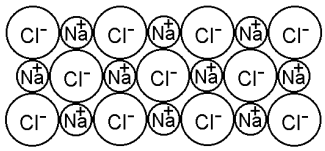
English (international) nomenclature

- Example 1: Two oxidation numbers II and III possible for Fe:
  - $FeCl_2$ Iron(II) chloride, ferrous chloride
  - $FeCl_3$ Iron(III) chloride, ferric chloride
- Example 2: oxoacids of chlorine, four oxidation numbers I, III, V and VII possible for Cl:
  - HClO hypochlorous acid
  - $HClO_2$ chlorous acid
  - $HClO_3$ chloric acid
  - $HClO_4$ perchloric acid

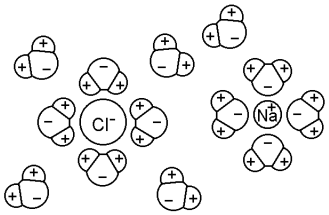
IONIC EQUATIONS

Ionic salts: no true molecule

- Crystal lattice of NaCl:



- Dissolution of NaCl in water: electrolytic dissociation producing hydrated independent ions Na^+, Cl^-



Reaction I

Stoichiometric equation:

$AgNO_3 + KCl \longrightarrow AgCl + KNO_3$

Ionic equation:

$Ag^+ + NO_3^- + K^+ + Cl^- \longrightarrow AgCl + K^+ + NO_3^-$

Net ionic equation:

$Ag^+ + \cancel{NO_3^-} + \cancel{K^+} + Cl^- \longrightarrow AgCl + \cancel{K^+} + \cancel{NO_3^-}$

$\mathbf{Ag^+ + Cl^- \longrightarrow AgCl}$ white ppt

Also possible:

$AgNO_3(aq) + KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$

$\mathbf{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$

(aq)	... aqueous
(s)	... solid
(l)	... liquid
(g)	... gaseous

Reaction II

$$CuSO_4 + 2NaOH \longrightarrow Cu(OH)_2 + Na_2SO_4$$

ionic:

$$Cu^{2+} + SO_4^{2-} + 2Na^+ + 2\,OH^- \longrightarrow Cu(OH)_2 + SO_4^{2-} + 2\,Na^+$$

net ionic:

$$Cu^{2+} + \cancel{SO_4^{2-}} + \cancel{2Na^+} + 2\,OH^- \longrightarrow Cu(OH)_2 + \cancel{SO_4^{2-}} + \cancel{2\,Na^+}$$

$$\boxed{\mathbf{Cu^{2+} + 2\,OH^- \longrightarrow Cu(OH)_2}}$$

pale blue ppt

Reaction III

$$2\,NaOH + (NH_4)_2SO_4 \longrightarrow 2\,NH_3 + Na_2SO_4 + 2H_2O$$

ionic:

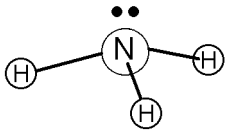
$$2Na^+ + 2OH^- + 2NH_4^+ + SO_4^{2-} \longrightarrow 2NH_3 + 2Na^+ + SO_4^{2-} + 2H_2O$$

net ionic:

$$\cancel{2Na^+} + \cancel{2OH^-} + \cancel{2NH_4^+} + \cancel{SO_4^{2-}} \longrightarrow 2NH_3 + \cancel{2Na^+} + \cancel{SO_4^{2-}} + \cancel{2}H_2O$$

$$\boxed{\mathbf{OH^- + NH_4^+ \longrightarrow NH_3 + H_2O}}$$

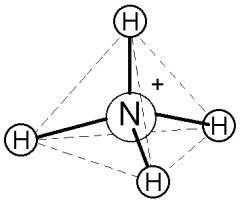
Ammonia gas: NH_3, $NH_3(g)$



Aqueous ammonia: $NH_3(aq)$, $NH_3.H_2O$, NH_4OH

$$NH_3.H_2O \longleftrightarrow NH_4OH$$

$$NH_4OH \longleftrightarrow NH_4^+ + OH^-$$



Reaction IV

$$Cu(OH)_2 + 2(NH_4)_2SO_4 + 2NaOH \longrightarrow [Cu(NH_3)_4]SO_4 + Na_2SO_4 + 4H_2O$$

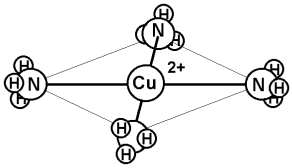
Ionic:

$$Cu(OH)_2 + 4NH_4^+ + 2SO_4^{2-} + 2Na^+ + 2OH^- \longrightarrow [Cu(NH_3)_4]^{2+} + 2SO_4^{2-} + 2Na^+ + 4H_2O$$

Net ionic:

$$Cu(OH)_2 + 4NH_4^+ + 2OH^- \longrightarrow [Cu(NH_3)_4]^{2+} + 4H_2O$$

Dark blue complex



Writing Ionic equations: Summary

1. write correct and balanced stoichiometric equation first
2. rewrite to ionic: write separately any species that exist separately and indicate its charge if present, but write together what exists joined (usually a precipitate of insoluble salt, or a soluble coordination complex)
3. Cancel out all species not involved in the reaction
4. Check that the equation is still balanced

What combinations of cations and anions are insoluble?

- All nitrates (NO_3^-) and acetates ($CH3COO^-$) are soluble
- All salts of Na, K, Li, and NH_4^+ are soluble
- All chlorides, bromides and iodides are soluble except salts of Pb^{2+}, Ag^+, and Hg_2^{2+}
- Most sulfate salts are soluble except $BaSO_4$, $PbSO_4$, $HgSO_4$, and $CaSO_4$.
- Most hydroxides are insoluble. Soluble are only NaOH and KOH. $Ba(OH)_2$, and $Ca(OH)_2$ are marginally soluble.
- Most sulfides (S^{2-}), carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble.

Names of coordination compounds

- Names of neutral ligands:
  - H_2O aqua
  - NH_3 ammin
  - NO nitrosyl
  - CO carbonyl
- Names of anionic ligands always end in –o:
  - F^- fluoro
  - Cl^- chloro
  - Br^- bromo
  - I^- iodo
  - OH^- hydroxo
  - CN^- cyano
  - etc..

Names of coordination compounds

1. Complex particle is cation:

e.g. $[Cu(NH_3)_4]SO_4$

$[Cu(NH_3)_4]^{2+} + SO_4^{2-}$

Tetraamminecopper(II) sulfate

2. Complex particle is anion:

e.g. $K_3[CoF_6]$

$3\ K^+ + [CoF_6]^{3-}$

Potassium hexafluorocobaltate(III)

Names of coordination compounds

3. Both cation and anion are complexes:

e.g. $[Pt(NH_3)_4][PtCl_4]$

$[Pt(NH_3)_4]^{2+} + [PtCl_4]^{2-}$

Tetraammineplatinum(II) tetrachloroplatinate(II)

4. Neutral complexes:

e.g. $[CrCl_3(H_2O)_3]$

Triaquatrichlorochromium(III) complex