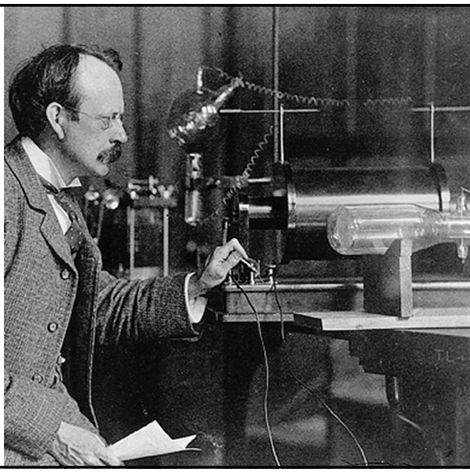


The Components of Matter

- Elements – the basic building blocks of matter
  - Ancient Greeks – four elements: earth, air, fire and water
- The atomic idea
  - Democritus – "… there are atoms and void."
  - Boyle (17th century) – "… simple bodies not made of any other bodies …"
  - Lavoisier (18th century) – natural laws
  - Dalton's atomic theory (19th century)
  - Atomic structure (20th century)

2.1 Elements, Compounds and Mixtures

- **Pure substances** – elements and compounds
  - Have constant composition
- **Elements** – consist of only one kind of atoms;
  - Can't be broken down to simpler substances
  - Have unique properties
  - Some elements consist of **molecules** (independent units made of 2 or more atoms)
- **Compounds** – consist of 2 or more chemically combined elements;
  - Can be broken down to simpler substances
  - Have properties unlike those of their component elements
  - Can't be separated into their components by physical means

- **Mixtures** – two or more elements or compounds that are physically intermingled
  - Have variable composition
  - Retain the properties of their components
  - Can be separated into their components by physical means

Examples:
- Silicon, sodium, chlorine, … – elements
- Water, sodium chloride, … – compounds
- Salt water, air, … – mixtures

2.2 Laws Leading to the Atomic View

- **Law of mass conservation** (Lavoisier) – the total mass of substances does not change during a chemical reaction
  - Matter cannot be created or destroyed

reactant 1 + reactant 2 ⟶ product

total mass = *total mass*

calcium oxide + carbon dioxide ⟶ calcium carbonate

56.08g + 44.01g ⟶ 100.09g

- **Law of constant composition** (Proust) – a pure compound always contains definite proportions (fractions) of its elements by mass

$$Mass\ fraction = \frac{Mass\ of\ element}{Mass\ of\ compound}$$

$$Mass\ \% = Mass\ fraction \times 100\%$$

Example: 9.0 g of water contains **8.0 g** of oxygen and **1.0 g** of hydrogen.

Mass fraction of H = 1.0/9.0 = 0.11 (or 11%)

Mass fraction of O = 8.0/9.0 = 0.89 (or 89%)

$$Mass\ of\ element = Mass\ of\ compound \times Mass\ fraction$$

Example: Analysis shows that **180.2 g** of glucose contains **72.1 g** of carbon, **96.0 g** of oxygen and the remainder is hydrogen. How many ***g*** of hydrogen are in **55.5 g** of glucose?

Mass of H in 180.2 g glucose =

$$= 180.2\text{ g} - 72.1\text{ g} - 96.0\text{ g} = 12.1\text{ g H}$$

$$\boldsymbol{Mass\ of\ H = 55.5\,\cancel{\text{g glucose}} \times \frac{12.1\text{ g H}}{180.2\,\cancel{\text{g glucose}}} =}$$

$$\mathbf{= 3.73\text{ g H}}$$

- **Law of multiple proportions** (Dalton) – if elements A and B form two different compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers

Example: Sulfur has 2 different oxides:

Oxide I $\rightarrow$ 1.0 g oxygen : 1.0 g sulfur

Oxide II $\rightarrow$ 1.5 g oxygen : 1.0 g sulfur

$$\frac{\mathbf{1.0\text{ g O}\ \cancel{\text{per 1 g S in I}}}}{\mathbf{1.5\text{ g O}\ \cancel{\text{per 1 g S in II}}}} = \frac{\mathbf{1.0}}{\mathbf{1.5}} \times \left(\frac{\mathbf{2}}{\mathbf{2}}\right) = \frac{\mathbf{2}}{\mathbf{3}}$$

2.3 Dalton's Atomic Theory

- Postulates of Dalton's atomic theory (1808)
  1. Matter consist of small, indivisible and indestructible atoms.
  2. All atoms of an element are identical in mass and different from the atoms of other elements.
  3. Compounds result from chemical combinations of different elements in specific atomic ratios
  4. Atoms don't change their identities in chemical reactions. They only recombine to form different compounds.

- Explanation of the mass laws
  - Conservation of mass – postulates 1 and 4
  - Constant composition – postulates 2 and 3
  - Multiple proportions – postulates 1, 2 and 3
- Relative atomic masses
  - Hydrogen was assigned a mass of **1** (lightest)
  - In water: **8g O : 1g H**
  - $\Rightarrow$ If the formula of water is HO, O should have relative mass of 8
  - $\Rightarrow$ If the formula of water is H_2O, O should have relative mass of 16 (16:2 = 8:1)

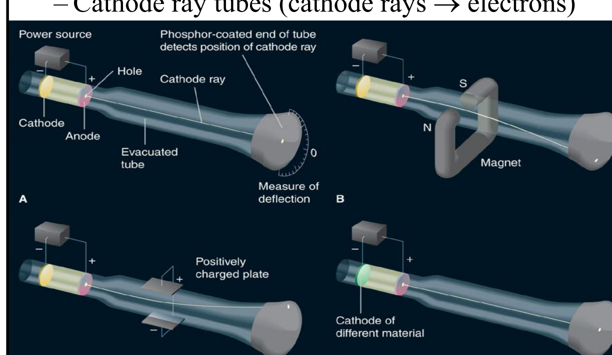
Example:

The two different oxides of sulfur have formulas SO_2 and SO_3.

$$\frac{\mathbf{2\text{ atoms O per 1 atom S in I}}}{\mathbf{3\text{ atoms O per 1 atom S in II}}} = \frac{\mathbf{2 \times 16}}{\mathbf{3 \times 16}} = \frac{\mathbf{2}}{\mathbf{3}}$$

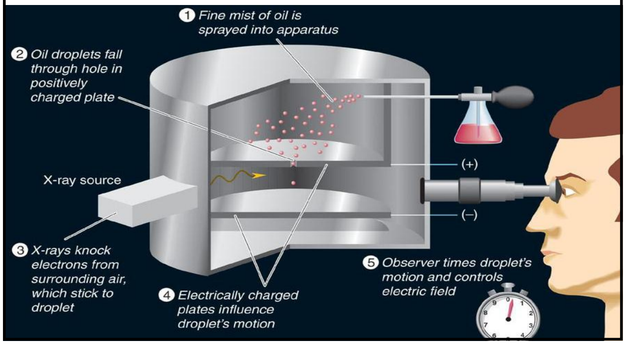
2.4 The Nuclear Model of the Atom

- **Discovery of the electron**
  - Cathode ray tubes (cathode rays $\rightarrow$ electrons)



– Mass-to-charge ratio of the electron (J.J. Thomson, 1897) → -5.686×10^{-12} kg/C

– Charge of the electron (R. Millikan, 1909) → -1.602×10^{-19} C

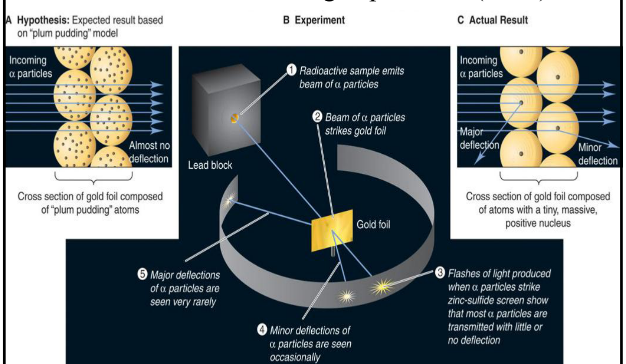


– Mass of the electron →

$(-1.602\times10^{-19}\text{ C})\times(-5.686\times10^{-12}\text{ kg/C})=9.109\times10^{-31}$ kg

- **Discovery of the nucleus**
  - Matter is electrically neutral → the negative electrons must be balanced by positive particles
  - J.J. Thomson's "plum pudding" model (electrons embedded in a diffuse sphere of positive charge)
  - Radioactivity (α, β, γ rays)
    - α-Particles – heavy and positive
    - β-Particles – light and negative
    - γ-Rays – electromagnetic radiation

– Rutherfords α-scattering experiment (1910)



– **Nucleus** – positive, heavy and compact

2.5 The Atomic Theory Today

- Subatomic particles (protons, neutrons and electrons)
  - Nucleus (protons & neutrons) and electron cloud

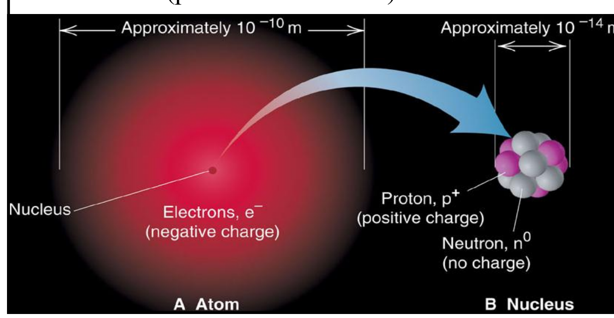


Table 2.2 Properties of the Three Key Subatomic Particles

	Charge		Mass	
Name (Symbol)	Relative	Absolute (C)*	Relative (amu)†	Absolute (g)
Proton (p^+)	1+	$+1.60218\times10^{-19}$	1.00727	1.67262×10^{-24}
Neutron (n^0)	0	0	1.00866	1.67493×10^{-24}
Electron (e^-)	1−	-1.60218×10^{-19}	0.00054858	9.10939×10^{-28}

*The coulomb (C) is the SI unit of charge.
†The atomic mass unit (amu) equals 1.66054×10^{-24} g; discussed later in this section.

- $\mathbf{p^+}$ – positive charge, ~2000 times heavier than the $\mathbf{e^-}$
- $\mathbf{n^0}$ – neutral, almost the same mass as the $\mathbf{p^+}$
- $\mathbf{e^-}$ – negative charge, same absolute charge as the $\mathbf{p^+}$
- Atoms are neutral ⇒ $\#e^- = \#p^+$

- **Atomic number** (**Z**) – number of protons in the atomic nucleus
  - All atoms of a given element have the same **Z**

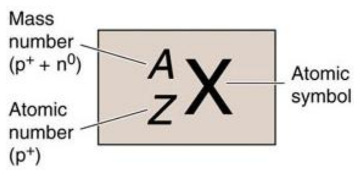
$$\mathbf{Z = \#p^+ = \#e^-}$$

- **Mass number** (**A**) – total number of protons and neutrons

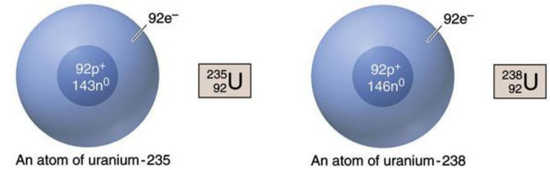
$$\mathbf{A = \#p^+ + \#n^0}$$

- Atomic symbols
  - H (hydrogen), C (carbon), O (oxygen), Ar (argon), Cl (chlorine)
  - Fe (iron, ferrum), Ag (silver, argentum), Sn (tin, stannum)

- **Isotopes and atomic masses**
  - The **#p⁺** in the nucleus of a given element is always the same, but the **#n⁰** can vary (**Z** is the same; **A** can vary)
- **Isotopes**
  - Atoms with the same **Z**, but different **A**
  - Belong to the same element, but have different atomic mass
- Isotopic symbols



- **Example:** Two of the isotopes of uranium (Z=92) have symbols ^{235}U and ^{238}U.



Problem:

- How many **p⁺**, **n⁰**, and **e⁻** are present in an atom of Plutonium-239?

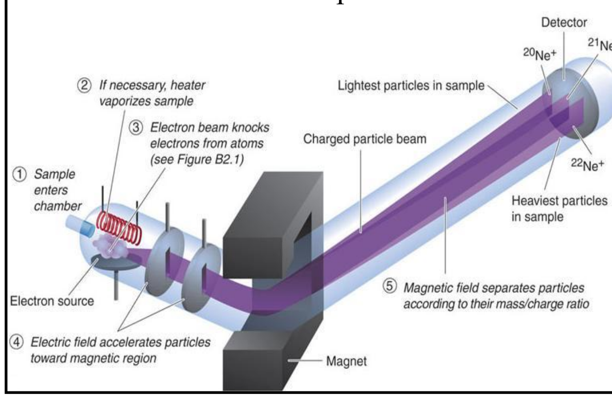
A = 239 Z = 94 (from list in textbook)

#n = A – Z = 239 – 94 = 145

#e = #p = Z = 94

- **Atomic mass unit** (**amu or D**) – **1/12** of the mass of a carbon-12 atom
  - Isotopic mass of $^{12}C \rightarrow$ 12 amu (exactly)
  - Isotopic mass of $^{1}H \rightarrow$ 1.008 amu
  - Isotopic mass of $^{29}Si \rightarrow$ 28.976 amu
- Elements occur in nature as mixtures of isotopes with certain abundances
- **Atomic mass** of an element – average of the masses of its naturally occurring isotopes (atomic masses are listed in the periodic table)

- **Mass spectrometer** – can measure the mass and abundance of isotopes



Problem:

Calculate the atomic mass of Cu, given that it naturally occurs as 69.17% ^{63}Cu (62.94 amu) and 30.83% ^{65}Cu (64.93 amu).

Use a weighted average:

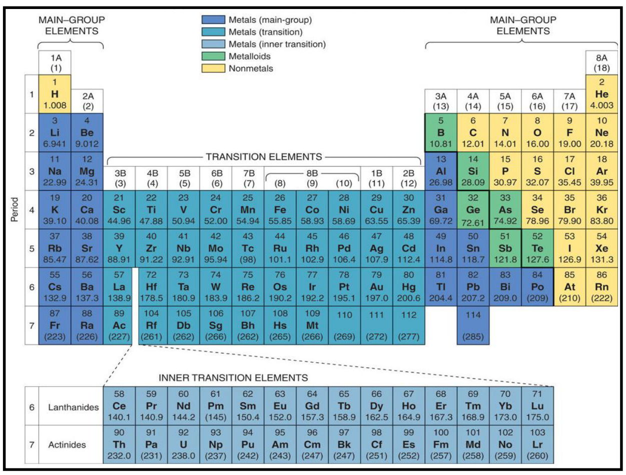
$$\textbf{Atomic mass of Cu} =$$
$$= \mathbf{0.6917 \times 62.94}\ \boldsymbol{amu} + \mathbf{0.3083 \times 64.93}\ \boldsymbol{amu}$$
$$= \mathbf{63.55}\ \boldsymbol{amu}$$

- **Reassessment of Dalton's atomic theory:**
  1. Matter consist of atoms that are *divisible and composed of protons, neutrons and electrons*.
  2. All atoms of an element have the same *number of protons in their nucleus* which is different from the atoms of other elements.
  3. Compounds result from chemical combinations of different elements in specific atomic ratios
  4. Atoms don't change their identities in chemical reactions. *Nuclear reactions can convert atoms of one element to another*.

2.6 Elements and the Periodic Table

- Periodicity in the properties of the elements
  - Mendeleev's table, 1871 – arrangement by atomic mass
  - Modern version of the table – arrangement by atomic number
- **Groups** – vertical columns in the table
  - ***A*** groups (1, 2, 13-18) – representative elements
  - ***B*** groups (3-12) – transition elements
  - Inner transition elements – lanthanides & actinides
- **Periods** – horizontal rows in the table

- Elements in a group have similar properties
- Elements in a period have different properties
- **Metals**
  - good electrical and heat conductivity, malleable, ductile
- **Nonmetals**
  - poor electrical and heat conductivity, neither malleable nor ductile, often gases or liquids
- **Metalloids**
  - semiconductors, intermediate properties



- Properties change gradually down in a group
  - Group 1A (1) **- alkali metals** (Li, Na, K, Rb,...)
    - soft, easy melting metals; react violently with water
    - reactivity increases down in the group
  - Group 2A (2) **- alkaline earth metals** (Be, Mg, ...)
    - similar but less reactive than Group 1
    - reactivity increases down in the group
  - Group 7A (17) **- halogens** (F, Cl, Br, I,...)
    - very reactive - reactivity increases up in the group
    - gradual change in physical properties - F, Cl (yellow gases), Br (red-brown liquid), I (purple-black solid)
  - Group 8A (18) **- noble gases** (He, Ne, Ar,...)
    - very low reactivity - inert gases
    - colorless, odorless gases

2.7 Compounds

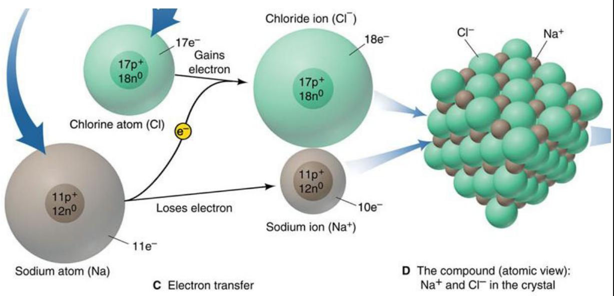
- Combination of two or more elements in some definite proportion
- **Chemical bonds** – the forces that hold the atoms of elements together in compounds
  - **Ionic bonding** – results from transfer of electrons from one atom to another
  - **Covalent bonding** – results from sharing of electrons between atoms
- **Ions** – el. charged atoms or groups of atoms
- **Molecules** – el. neutral groups of atoms covalently bonded together

- **Ionic compounds** – consist of positive and negative ions held together by electrostatic attraction (NaCl, CaO, ...)
  - Positive ions (**cations**) – often produced when metals lose electrons (Na^+, Ca^{2+}, ...)
  - Negative ions (**anions**) – often produced when nonmetals gain electrons (Cl^-, O^{2-}, ...)
- Binary ionic compounds – composed of just 2 elements (typically a metal and a nonmetal)
- Monatomic ions – formed through gain or loss of e^- by single atoms

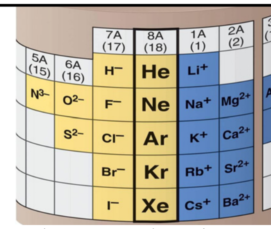
- Formation of binary ionic compounds

Example: NaCl

– The electrons lost by Na are gained by Cl



- Charges of monoatomic ions can be predicted from the periodic table



– Typically metals loose e^- and nonmetals gain e^- until they reach the same number of e^- as in the nearest noble gas (high stability)

– Groups **1A–3A** form cations with charges equal to the **group#** (only the lighter members of 3A)

– Groups **5A–7A** - anions with charges equal to the **group# - 8** (only the lighter members of 5&6A)

- The strength of ionic bonds depends on the charges and sizes of the ions
  – Potential energy of interaction between two ions with charges q_1 and q_2 separated by a distance r_{12}

$$E_p = \frac{q_1 \times q_2}{r_{12}}$$

⇒**Ions with higher charges and smaller sizes attract each other stronger**

- Ionic compounds are neutral → the # of positive charges must equal the # of negative charges (charge balance)

Problems:

1. What are the charges of the monatomic ions formed by Al and Br?

Al → Group 3A → 3+ → Al^{3+}

(loss of $3e^-$ →Ne)

Br → Group 7A → 7 – 8 = -1 → Br^-

(gain of $1e^-$ → Kr)

2. What is the ratio of **Al^{3+}** to **Br^-** ions in the binary ionic compound of these elements?

Al^{3+} : Br^- → **1 : 3** ← 1(+3) + 3(-1) = 0

- **Covalent compounds** – typically consist of molecules in which atoms are bonded together through sharing of electrons → **molecular compounds** (H_2O, NH_3, …)
  – Formed usually between nonmetals
  – Some elements occur in nature in a molecular form (H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2, P_4, S_8, …)
- **Polyatomic ions** – consist of two or more covalently bonded atoms with a net overall charge (NH_4^+, SO_4^{2-}, …) → participate in ionic bonding

2.9 Mixtures

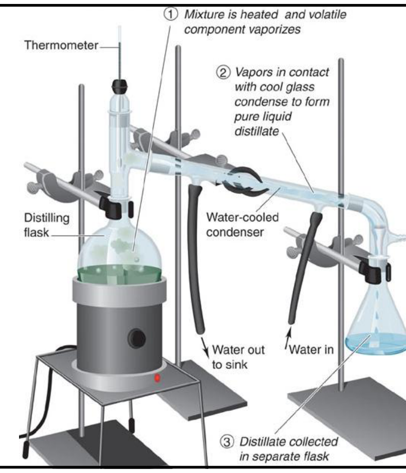
- Contain more than one pure substances
- **Heterogeneous mixtures** - composition changes from one part to another (soil, blood, milk, dust, fog, ...)
- **Homogeneous mixtures** - composition is uniform throughout (sea water, air, gasoline, vinegar, brass, ...)
- **Solutions** - homogeneous mixtures
  – solvent - present in the larger amount
  – solute - the dissolved substance
- Aqueous solutions - the solvent is water

- Differences between mixtures and compounds

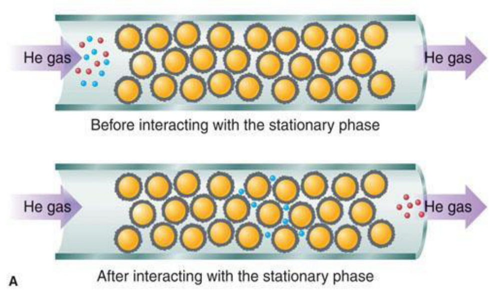
Mixture	Compound
Components can be separated by using physical techniques.	Components cannot be separated by using physical techniques.
Composition is variable.	Compositon is fixed.
Properties are related to those of its components.	Properties are unlike those of its components.

- Separation of mixtures (relies on differences in the physical properties of the components)
  - **Extraction -** differences in the solubility
  - **Filtration -** differences in particle size

- **Distillation -** differences in the volatility (boiling point)



- **Chromatography -** differences in the ability to adsorb on surfaces or absorb into liquids
  - Stationary and mobile phases
  - GC
  - LC



2.8 Formulas, Names and Masses of Compounds

- **Chemical formula** – shows the elemental composition of a compound
- **Molecular formula (MF)** – gives the number of atoms of each type in a molecule
  - butane $\rightarrow C_4H_{10} \rightarrow$ 4C &10H atoms in 1 molecule
- **Empirical formula (EF)** – shows the relative number of atoms of each type in terms of the smallest whole numbers
  - butane, MF $\rightarrow C_4H_{10}$, EF $\rightarrow C_2H_5$

- **Structural formula** – gives the type and number of atoms in a molecule and how they are bonded (water $\rightarrow$ H–O–H)
- **Nomenclature** – system of naming compounds (common and systematic names)

Names of Ions

- Monatomic cations
  - ***name of element* + *ion*** ($Ca^{2+} \rightarrow$ calcium ion)
  - *Roman numeral* for the charge of the ion, if more than one charges are possible ($Fe^{2+} \rightarrow$ iron(II) ion, $Fe^{3+} \rightarrow$ iron(III) ion)
  - No roman numerals for group 1&2 cations, Al^{3+}, Zn^{2+}, Cd^{2+} and Ag^{+}

- Monatomic anions
  - ***root of element name* + *-ide* + *ion*** (Cl^- $\rightarrow$ chlor*ide* ion, O^{2-} $\rightarrow$ ox*ide* ion)
- Polyatomic ions - Table 2.5 (memorize)
  - oxoanions
    - ***root of element name* + *-ate* + *ion*** ($SO_4^{2-} \rightarrow$ sulf*ate* ion)
    - oxoanions with different number of O atoms - ***hypo- (per-)* + *root of element name* + *-ite* (*-ate*) + *ion*** ($ClO^- \rightarrow$ *hypo*clor*ite*, $ClO_2^- \rightarrow$ chlor*ite*, $ClO_3^- \rightarrow$ chlor*ate*, $ClO_4^- \rightarrow$ *per*chlor*ate* ion)
  - H-containing anions – add ***"hydrogen"*** to the name ($H_2PO_4^-$ dihydrogen phosphate ion)

Names & Formulas of Ionic Compounds

– ***Name of cation + name of anion***

- Names from formulas
  – $MgSO_4 \rightarrow Mg^{2+}, SO_4^{2-} \rightarrow$ magnesium sulfate
- Charge balance
  – $Co_2S_3 \rightarrow Co_2^{X+}S_3^{2-} \rightarrow 2(+X) + 3(-2) = 0 \rightarrow X=+3$
  $\rightarrow$ cobalt(III) sulfide
- Formulas from names (**cross rule**)
  – Aluminum sulfate $\rightarrow Al^{3+}, SO_4^{2-} \rightarrow Al_2(SO_4)_3$
- Formulas of ionic compounds are empirical

- **Hydrates** – ionic compounds containing a definite proportion of water ($CoCl_2{\cdot}6H_2O \rightarrow$ cobalt(II) chloride hexahydrate)
  – Greek prefixes - ***mono-, di-***,... (memorize)

Table 2.6 Numerical Prefixes for Hydrates and Binary Covalent Compounds

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

Examples:

- Give the systematic names of Na_2HPO_4 and $CuSO_4{\cdot}5H_2O$.
  $Na^+ , HPO_4^{2-} \rightarrow$ sodium hydrogen phosphate
  $Cu^{2+} , SO_4^{2-} \rightarrow$ copper(II) sulfate pentahydrate
- Give the formulas of manganese(II) fluoride and barium oxide.
  Manganese(II) fluoride ($Mn^{2+}\ F^-$) $\rightarrow MnF_2$
  Barium oxide ($Ba^{2+}\ O^{2-}$) $\rightarrow Ba_2O_2 \rightarrow BaO$

Names & Formulas of Covalent Compounds

- **Binary molecular compounds**
  – ***name of 1st element + root for 2nd element + ide***
  – Greek prefixes for the # of atoms of each type; *mono-* can be omitted except for O
  – The element with lower group # is written 1st (exception $\rightarrow$ O is always last)

Examples:

$N_2O \rightarrow$ *di*nitrogen *monox*ide

$SF_4 \rightarrow$ sulfur *tetra*fluor*ide*

$H_2S \rightarrow$ *di*hydrogen sulf*ide*

- **Acids** – release H^+ in water
  – Aqueous solutions of **binary** compounds containing H ($HCl_{(aq)}$, $H_2S_{(aq)}$, $HCN_{(aq)}$, ...)
    - ***hydro- + root of element + -ic + acid*** (*hydro*chlor*ic* acid, *hydro*sulfur*ic* acid, *hydro*cyan*ic* acid)
    - parent acids of anions ending on *-ide*
  – **Oxoacids** – parent acids of oxoanions; the # of acidic H atoms equals the charge of the anion (HNO_3, H_2SO_3, H_3PO_4, …)
    - ***root of element + -ic + acid*** - for anions ending on *-ate* ($HNO_3 \rightarrow$ nitric acid)
    - ***root of element + -ous + acid*** - for anions ending on *-ite* ($H_2SO_3 \rightarrow$ sulfurous acid)

Examples:

- Name the acids: $HBr_{(aq)}$, $HClO$, H_2CO_3, $HBrO_4$
  *hydro*brom*ic* acid $\leftarrow$ (Br^- , brom*ide*)
  hypochlor*ous* acid $\leftarrow$ (ClO^- , hypochlor*ite*)
  carbon*ic* acid $\leftarrow$ (CO_3^{2-} , carbon*ate*)
  perbrom*ic* acid $\leftarrow$ (BrO_4^- , perbrom*ate*)

- How to distinguish between ionic and molecular compounds
  - Molecular – typically consist of nonmetals (H_2O, NH_3, CO_2, C_2H_6O, ...)
  - Ionic – combination of metals and nonmetals ($NaCl$, $MgSO_4$, $AlPO_4$, KOH, ... exception NH_4^+ containing)

Example:

- Classify the compounds CH_5N, NH_4NO_3 and HCl as ionic or covalent and identify the ions if present.

Problem:

- Name the compounds: PF_5, CrF_3, N_2O_3, Fe_2O_3
  - molecular → phosphorus pentafluoride
  - ionic, F^- , Cr^{3+} → chromium(III) fluoride
  - molecular → dinitrogen trioxide
  - ionic, O^{2-}, Fe^{3+} → iron(III) oxide

Problem:

- Write the formulas of: zinc phosphate, vanadium(V) oxide, xenon tetrafluoride, cobalt(II) chloride hexahydrate
  - Zn^{2+} , PO_4^{3-} (cross rule) → $Zn_3(PO_4)_2$
  - V^{5+} , O^{2-} (cross rule) → V_2O_5
  - molecular → XeF_4
  - Co^{2+} , Cl^- (cross rule) → $CoCl_2{\cdot}6H2O$

Molecular and Formula Mass

- **Molecular mass –** sum of the atomic masses of the atoms in a molecule
  - $CO_2 \rightarrow 1\times12.01 + 2\times16.00 = 44.01$ amu
- **Formula mass –** sum of the atomic masses of the atoms in one formula unit of an ionic compound
  - $Ca(ClO_4)_2 \rightarrow 1\times40.08 + 2\times35.45 + 8\times16.00 =$
    $= 238.98$ amu