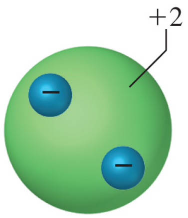
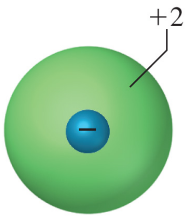
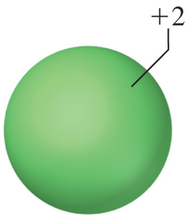
Chapter 2:
The Components of Matter



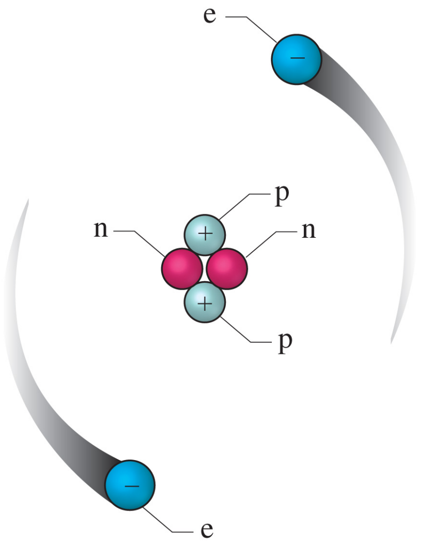
Helium atom
He

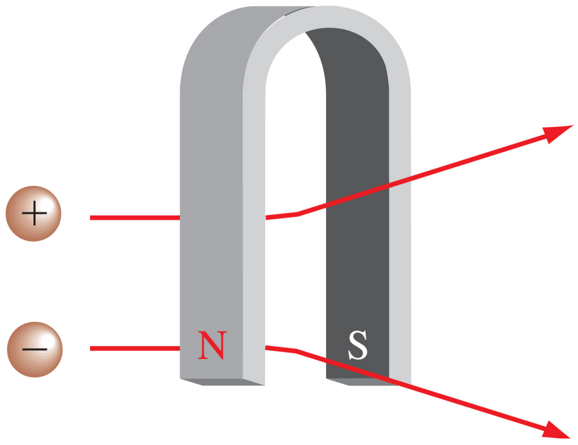


Helium ion
He^{+}



Helium ion
He^{2+}





Elements and Compounds

- Element:
  - cannot be decomposed into a simpler substance through chemical processes; distinguished by the unit of the atom
- Compound:
  - a substance made from the atoms of two or more elements bonded chemically in defined proportions

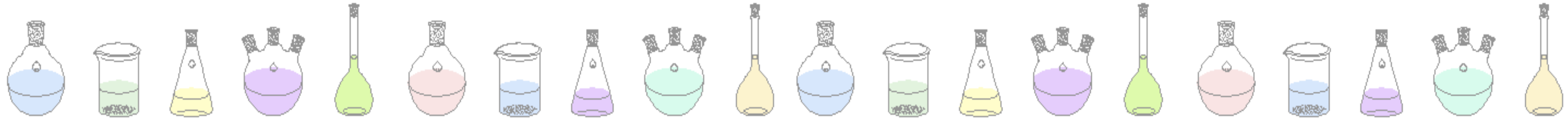
- Compounds can only be decomposed into their respective elements via chemical processes

Pure Substances and Mixtures

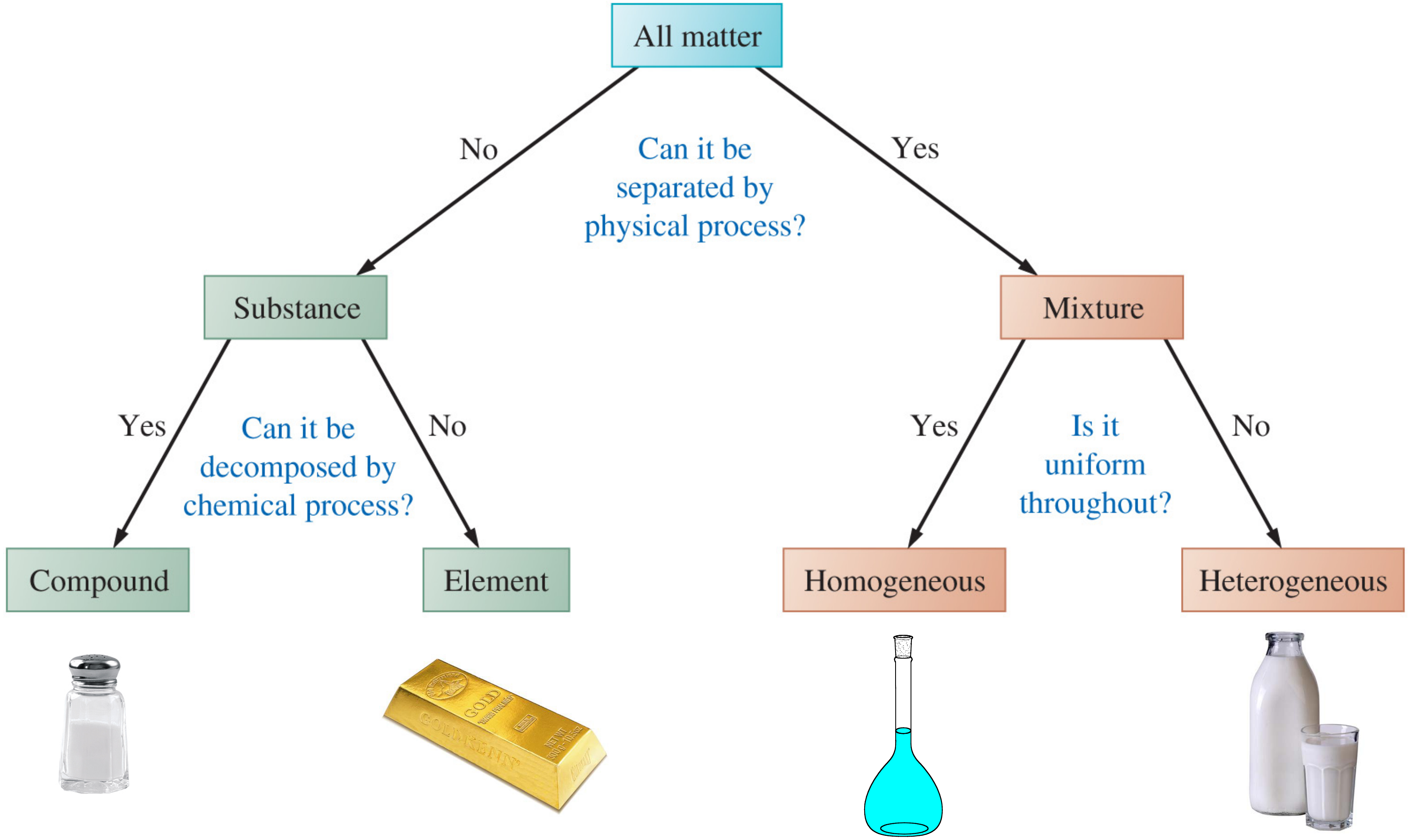
- A pure substance
  - A substance with a fixed and uniform composition and distinct properties (ex: pure water)
- A mixture
  - A combination of two or more pure substances which can vary in composition and properties

  a) homogeneous example: salt water

  b) heterogeneous example: oil and water
- It is possible to separate mixtures through physical processes



Pure Substances and Mixtures



All matter

No

Can it be separated by physical process?

Yes

Substance

Mixture

Yes

Can it be decomposed by chemical process?

No

Yes

Is it uniform throughout?

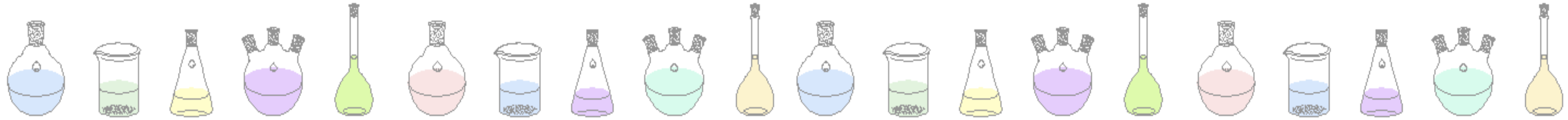
No

Compound

Element

Homogeneous

Heterogeneous



Early Chemical Discoveries

Lavoisier, 1743-1794

Conservation of Matter - A chemical equation must be balanced for mass.

Dalton (1766 – 1844)

Atomic Theory - A chemical equation must have the same number of atoms of the same kind on both sides.

The Law of Conservation of Mass

The total mass during a chemical or physical process must remain constant

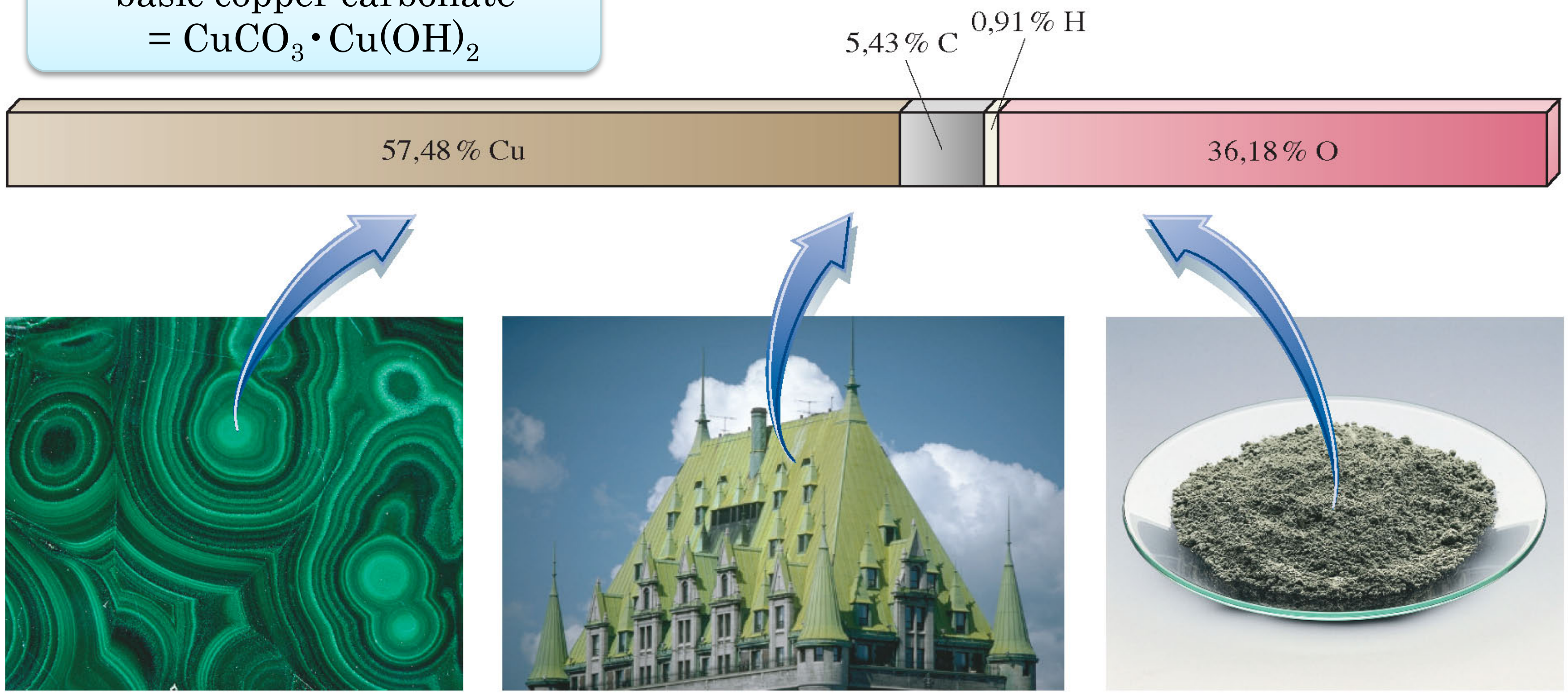
$2\ AgNO_3\ (aq)$
+
$Na_2CrO_4\ (aq)$

$2\ NaNO_3\ (aq)$
+
$Ag_2CrO_4\ (s)$

The Law of Definite Composition

All samples of a given compound have the same composition

‘basic copper carbonate”
= $CuCO_3 \cdot Cu(OH)_2$



The Law of Multiple Proportions

Combinations of elements are in ratios of small whole numbers (to make molecules!)

Carbon monoxide (CO):

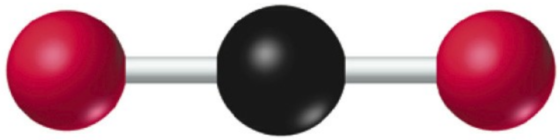
1.33 g of oxygen combines with 1.0 g of carbon or O/C = 1.33

Carbon dioxide (CO_2):

the ratio of O/C is 2.667

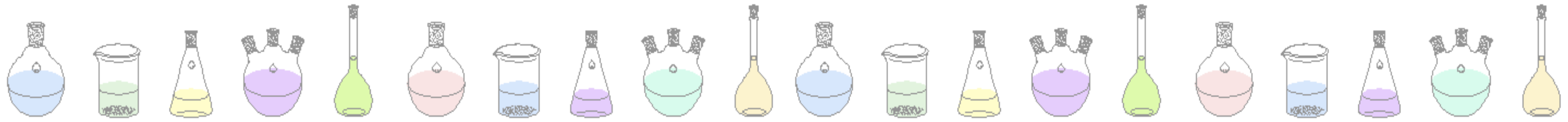
2.667/1.33 = 2.0

…which is the ratio of oxygen in CO_2/oxygen in CO



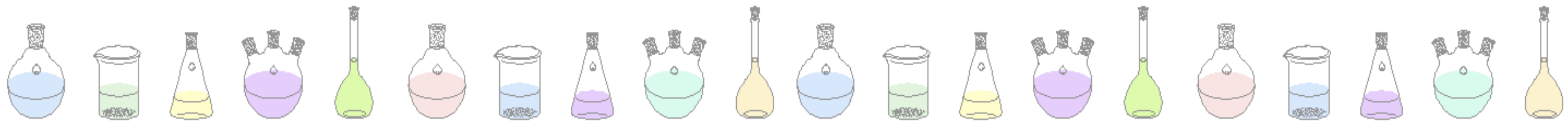
Development of Dalton's Theory

- Dalton's work helped to explain several empirical observations:
  - the law of constant composition
  - the law of multiple proportions
  - the law of conservation of mass



Dalton's Atomic Theory

1. All matter is made of small, indestructible particles called "atoms"
2. All atoms of a given element are identical and atoms differ for different elements
3. Atoms of one element cannot be transformed into atoms of another element
4. Atoms of different elements combine in simple, whole number ratios to form compounds



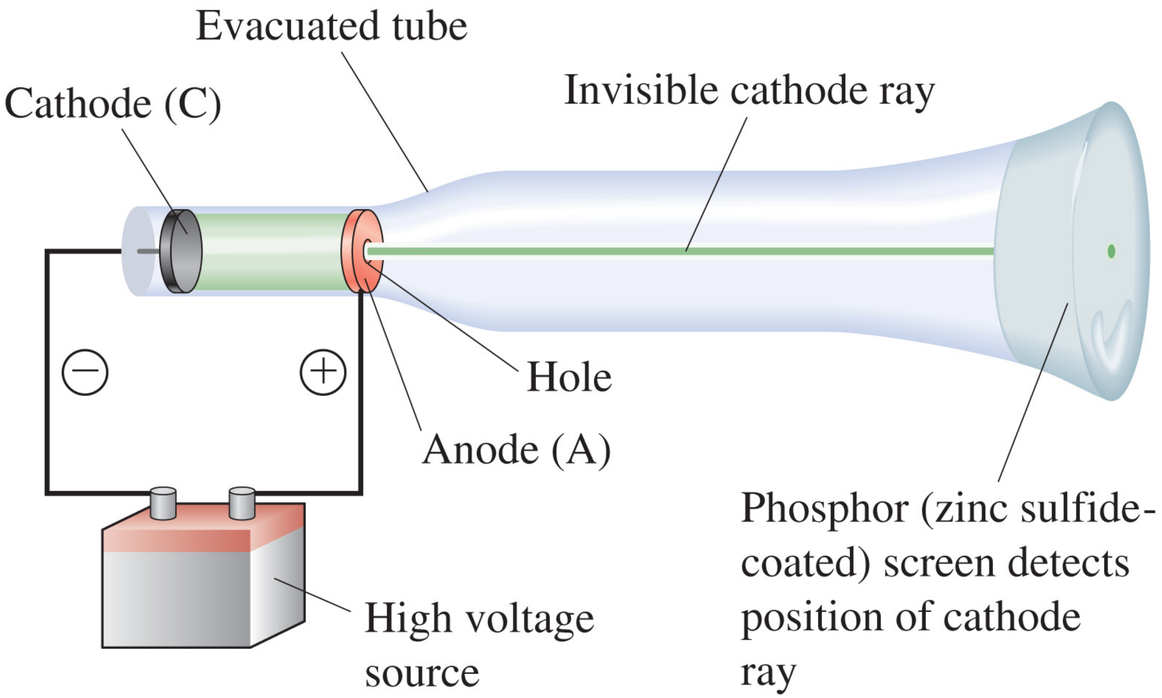
Development of Atomic Theory

- the concept of an indivisible atom proposed by Dalton inspired other scientists
- led to the discoveries of sub-atomic particles
  - J.J. Thompson: Z/m of the electron
  - Millikan: Z of the electron
  - Rutherford: the atomic nucleus

Atomic Structure: the electron

J.J. Thompson: 1898 – 1903

- the existence of a subatomic particle with a negative charge was demonstrated using a cathode-ray tube: the electron

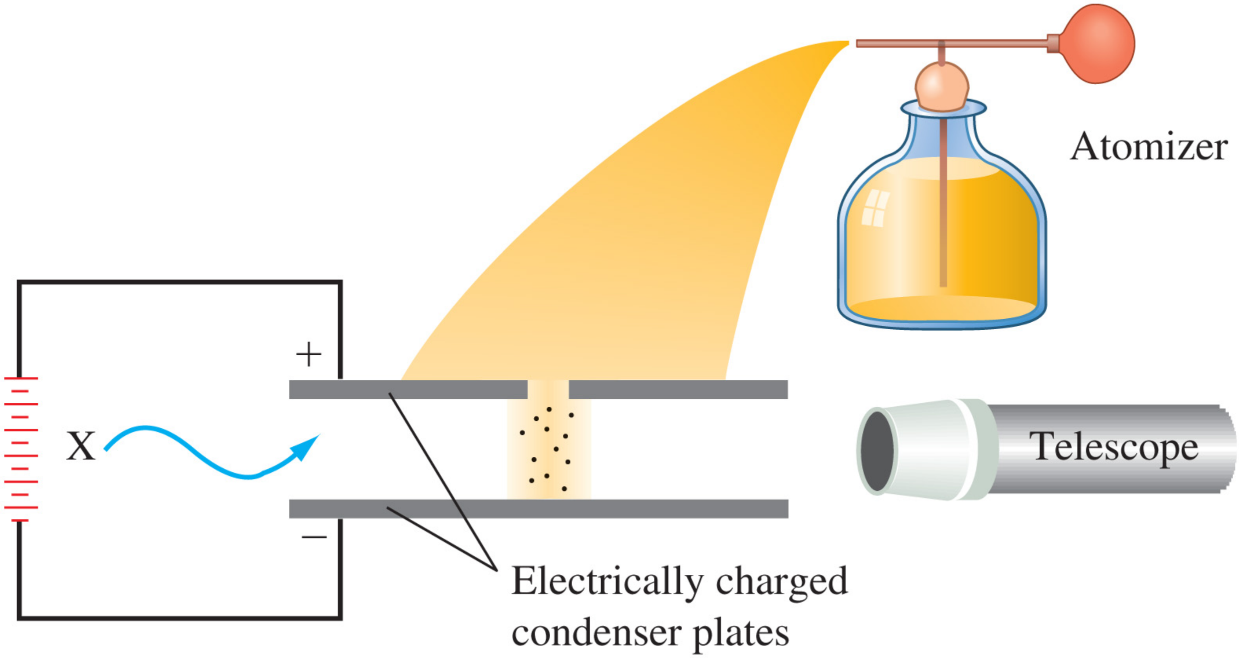


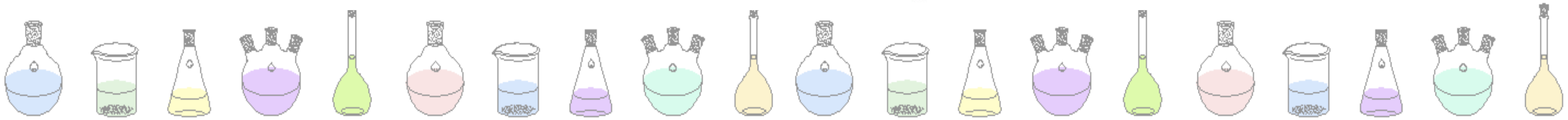
$$\frac{\text{charge}}{\text{mass}} = -1.76 \text{ x } 10^{8} \text{ C/g}$$

Atomic Structure: the electron

R. Millikan, 1909

- conceived an ingenious experiment to determine the exact charge of the electron
- his setup measured the speed of the fall of tiny droplets of oil



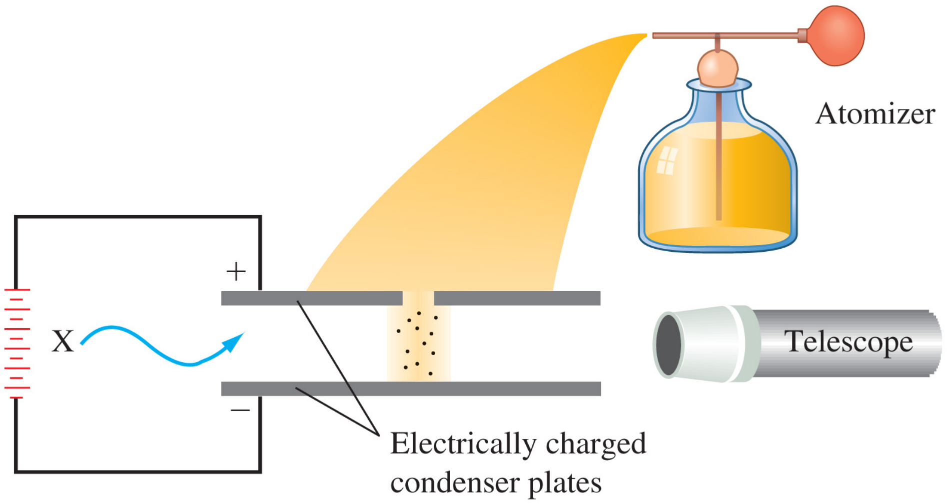


Atomic Structure: the electron

R. Millikan, 1909

- with this experiment, he established the charge on a single electron

$$\text{charge} = -1.60\text{x}10^{-19}\ \text{C}$$

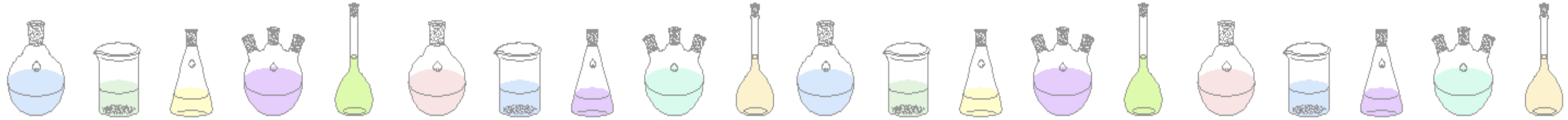


- using the Z/m ratio found by Thompson, the electron's mass was then:

$$\text{mass} = Z \bullet \frac{m}{Z}$$

$$= -1.60 \times 10^{-19}\ \text{C} \bullet \frac{\text{g}}{-1.76 \times 10^{8}\text{C}}$$

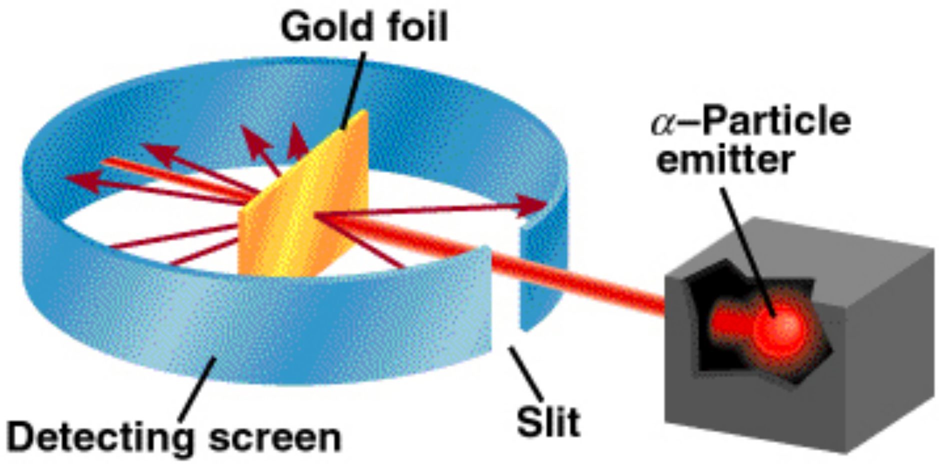
$$= 9.09 \times 10^{-28}\ \text{g}$$

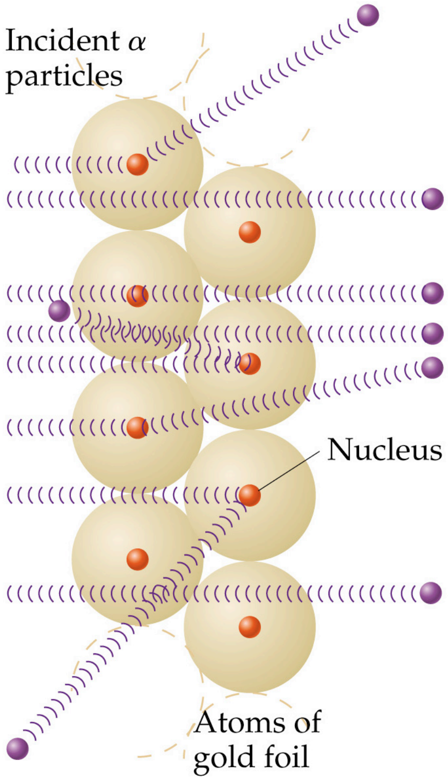


Atomic Structure

E. Rutherford, 1911 (in Montreal)

- in his experiment, a very thin foil of gold was bombarded with α particles

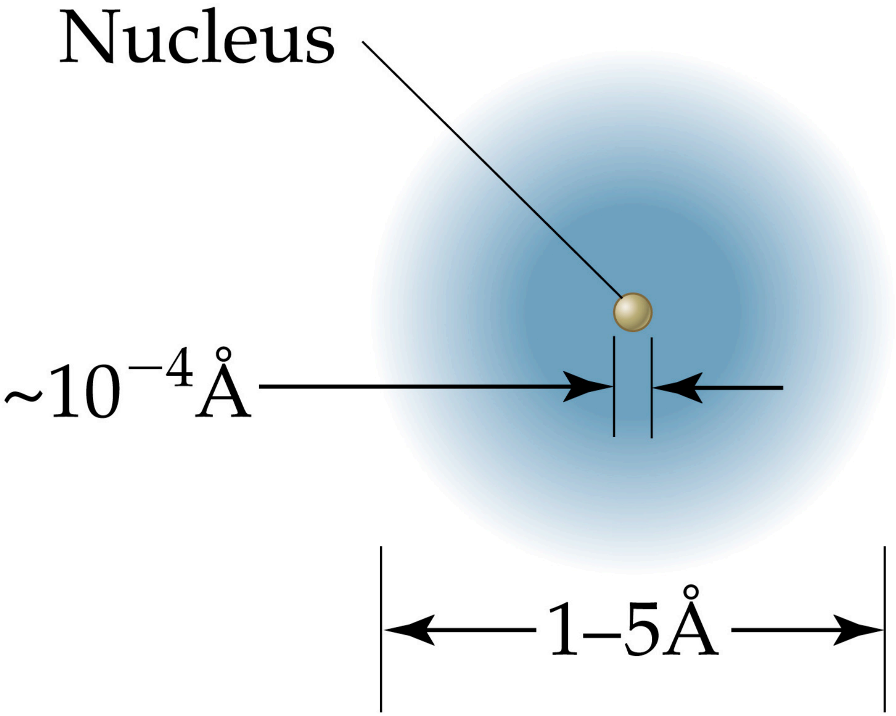




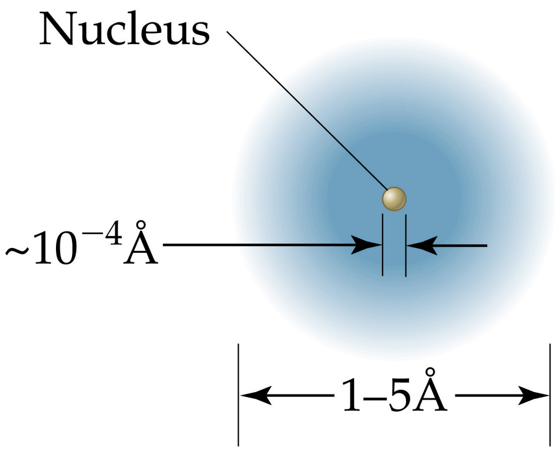
The Modern View of Atomic Structure

The Nucleus:

- Rutherford and Chadwick later discovered that the nucleus contains protons and neutrons
- nuclear radius is 20,000 times smaller than the atomic radius!



Atomic Structure



The mass of the atom is due to the nucleus, the volume is due to the orbits of the electrons

Particle	Symbol	Mass	Charge
Proton	p^+	1.67 x 10^{-24} g	+1.60 x 10^{-19} C
Neutron	n	1.67 x 10^{-24} g	0
Electron	e^-	9.11 x 10^{-28} g	-1.60 x 10^{-19} C

*note: error in Table 2.2

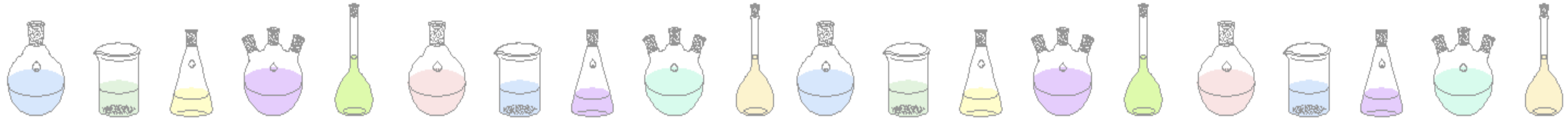
Isotopes, Atomic Numbers, and Mass Numbers

- **atomic number, Z** = number of protons
- **atomic mass, A** = sum of protons and neutrons

- For the element X, we write: $^{A}_{Z}\mathrm{X}$

1. each element has a unique value of Z

2. isotopes of an element will have the same value of Z, but different values of A (due to number of neutrons)



Isotopes

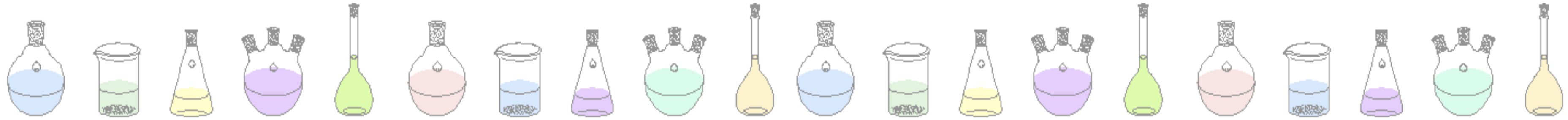
- the different isotopes of an element usually display similar chemistry, since the number of neutrons has little influence on the reaction’s outcome
- we use the atomic mass to identify isotopes:

$^{235}_{92}\text{U}$ is uranium-235

$^{238}_{92}\text{U}$ is uranium-238

- the isotopes of hydrogen have special names:

$^{1}_{1}\text{H}$	$^{2}_{1}\text{H}$	$^{3}_{1}\text{H}$
hydrogen	deuterium	tritium



Atomic Masses

- In the SI system, the standard for atomic mass is pure carbon-12:

$$\text{Mass of one } ^{12}\text{C atom} = 12 \text{ amu}$$

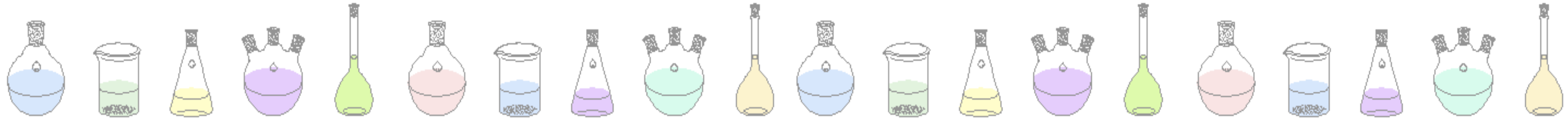
where 12 amu is an exact value (i.e. = 12.0000000000000…)

1 amu (« u ») = 1/12 the mass of a C atom

= 1.66054×10^{-24} g

or 1 g = 6.02214×10^{23} u

- The masses of all other elements are measured relative to the ^{12}C atom



Average Atomic Masses

- **<u>IMPORTANT</u>**: in nature, the elements exist as a mixture of isotopes
- we must account for the relative proportions of each isotope in our calculations
- Atomic mass = weighted average of the respective atomic masses of the natural isotopes of a given element

$$\boxed{\mathrm{AAM} = \mathrm{M}_1\mathrm{f}_1 + \mathrm{M}_2\mathrm{f}_2 + \mathrm{M}_3\mathrm{f}_3 + \ldots}$$

Average Atomic Masses

- ex. carbon has two naturally occurring isotopes:

	^{12}C	^{13}C
Mass	12 u	13.00335 u
Abundance	98.892 %	1.108 %

The average atomic mass of carbon is:

$$(0.98892)(12) + (0.0108)(13.00335) = 12.011 \text{ u}$$

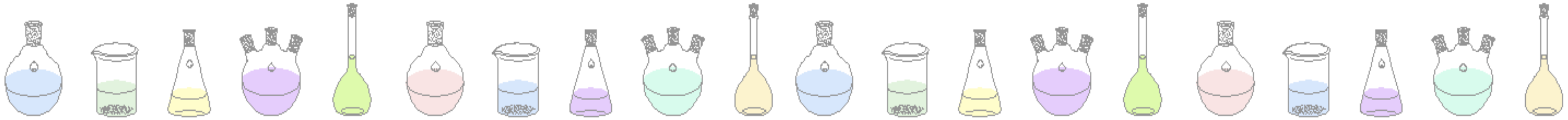
Example: Average Atomic Mass

The atomic masses of the two stable isotopes of boron, boron-10 (19.78%) and boron-11 (80.22%), are 10.0129 u and 11.0093 u, respectively. Calculate the average atomic mass of boron.

$$\text{atomic mass of B} = (M_1 \bullet f_1) + (M_2 \bullet f_2)$$

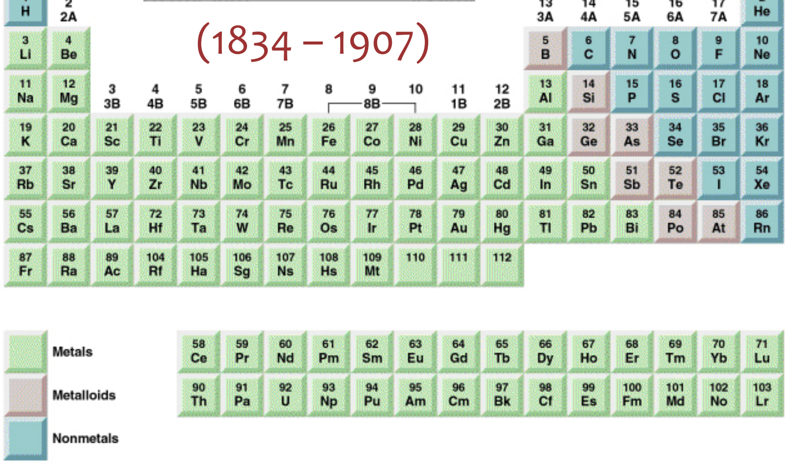
$$= (0.1978)(10.0129\ \text{u}) + (0.8022)(11.0093\ \text{u})$$

$$= 10.81\ \text{u}$$



The Periodic Table

(1834 – 1907)



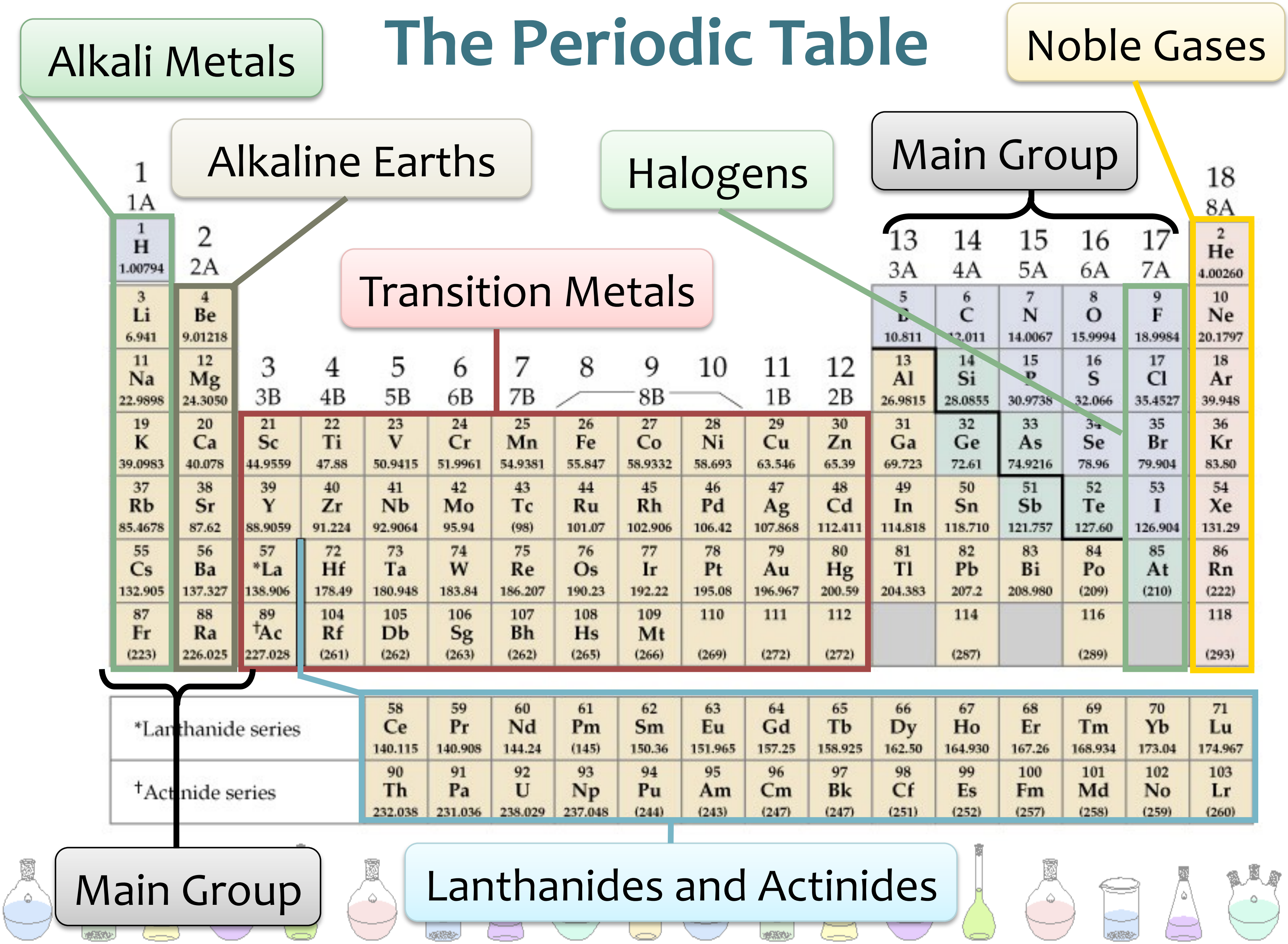
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110	111	112						

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

🟩 Metals
🟫 Metalloids
🟦 Nonmetals

- table in which the elements are grouped together according to their similar physical and chemical properties
- the horizontal rows are called periods
- the vertical columns are called groups (or families)
- three categories of elements:
  - metals (good conductors)
  - non-metals (poor conductors)
  - metalloids or semi-metals

The Periodic Table



Alkali Metals

Alkaline Earths

Transition Metals

Halogens

Main Group

Noble Gases

Main Group

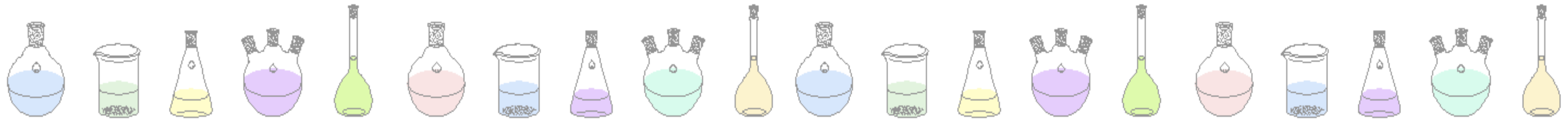
Lanthanides and Actinides

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H** 1.00794																	2 **He** 4.00260
3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.1797
11 **Na** 22.9898	12 **Mg** 24.3050											13 **Al** 26.9815	14 **Si** 28.0855	15 **P** 30.9738	16 **S** 32.066	17 **Cl** 35.4527	18 **Ar** 39.948
19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.9381	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.61	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.906	46 **Pd** 106.42	47 **Ag** 107.868	48 **Cd** 112.411	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.757	52 **Te** 127.60	53 **I** 126.904	54 **Xe** 131.29
55 **Cs** 132.905	56 **Ba** 137.327	57 ***La** 138.906	72 **Hf** 178.49	73 **Ta** 180.948	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.967	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.980	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.025	89 **†Ac** 227.028	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 **Ce** 140.115	59 **Pr** 140.908	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.965	64 **Gd** 157.25	65 **Tb** 158.925	66 **Dy** 162.50	67 **Ho** 164.930	68 **Er** 167.26	69 **Tm** 168.934	70 **Yb** 173.04	71 **Lu** 174.967
†Actinide series	90 **Th** 232.038	91 **Pa** 231.036	92 **U** 238.029	93 **Np** 237.048	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Naming Compounds

- in this course: inorganic compounds
- next semester: organic compounds
- Why do we have to do this???
  - learning how to accurately name compounds now will save you LOTS of time and hassle later!!



Introduction to Bonding

- electrons form chemical bonds between atoms:

$$X - Y$$

electrons are shared

covalent bond

nonmetal + nonmetal

$$X^{+} \; {:}Y^{-}$$

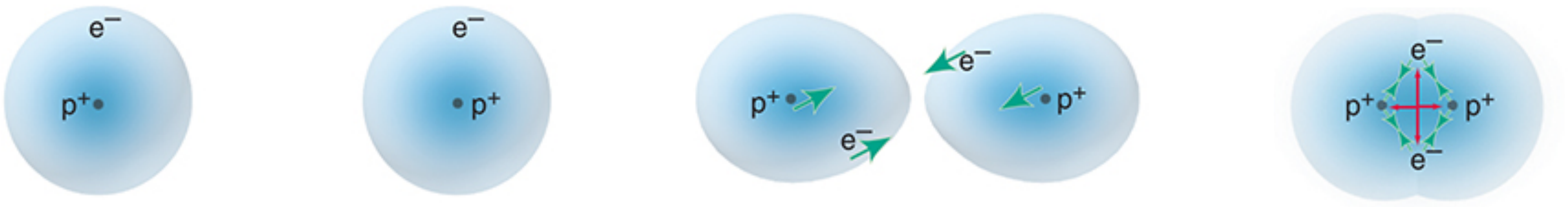
electrons are transferred

ionic bond

metal + nonmetal

Covalent Compounds

- atoms bonded covalently form individual, discrete assemblies called **molecules**



Naming binary covalent compounds

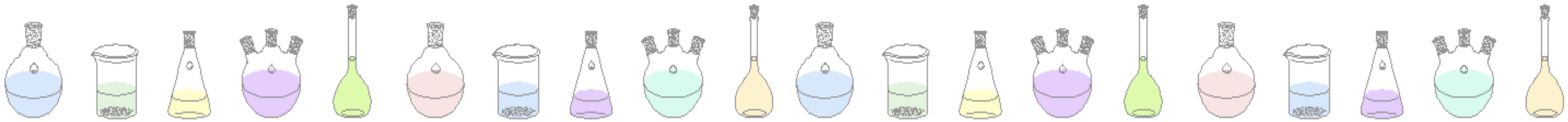
- non-metal + non-metal
- some compounds have common names:

H_2O NH_3

- name lower group number element first

CO NOT OC

- suffix –*ide*
- Greek prefixes are used to indicate the numbers of each element in the compound



Naming binary covalent compounds

Number	Prefix	Example
1	mono-	NO: nitrogen monoxide
2	di-	NO_2: nitrogen dioxide
3	tri-	N_2O_3: dinitrogen trioxide
4	tetra-	N_2O_4: dinitrogen tetroxide
5	penta-	N_2O_5: dinitrogen pentoxide
6	hexa-	SF_6: sulfur hexafluoride

Ionic compounds

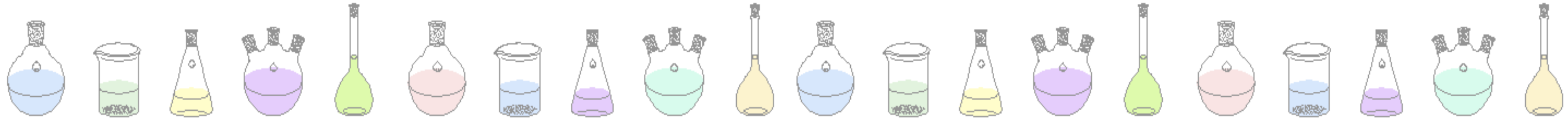
- metal + nonmetal

Metals → lose electrons →

positively charged ions = ***cations***

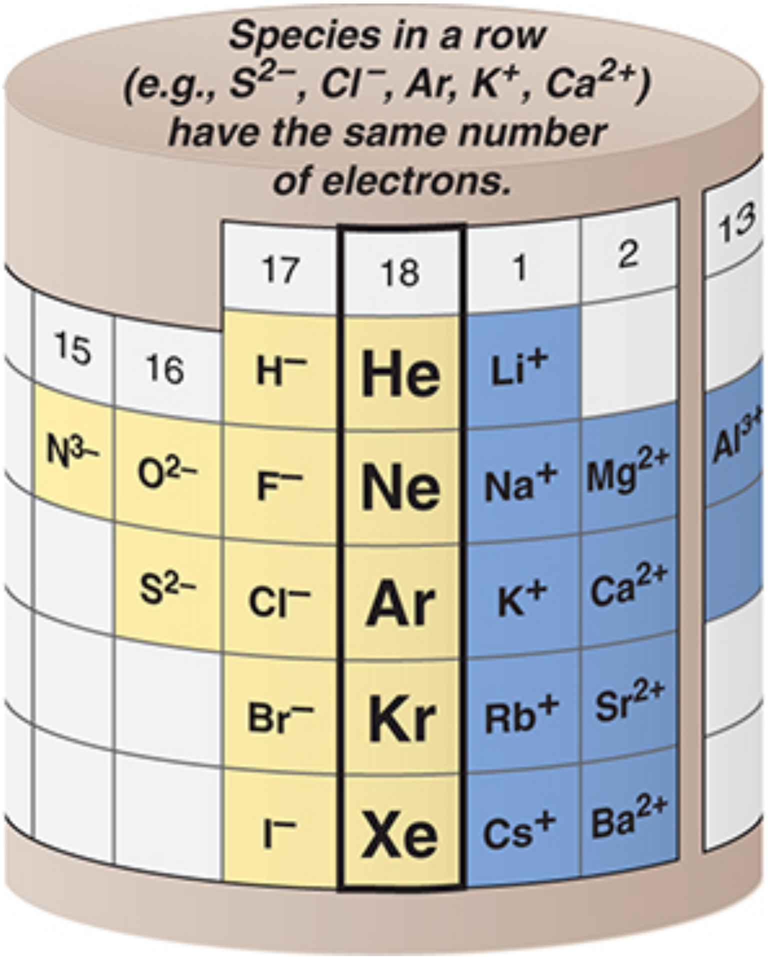
Non-metals → gain electrons →

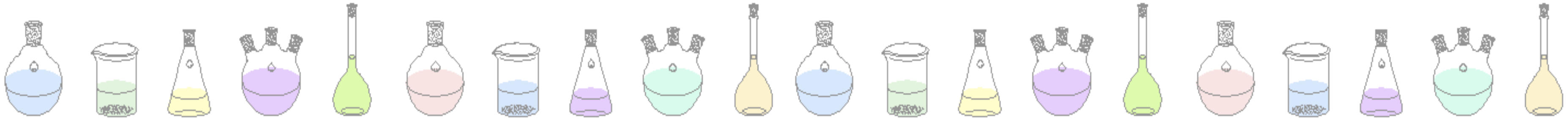
negatively charged ions = ***anions***



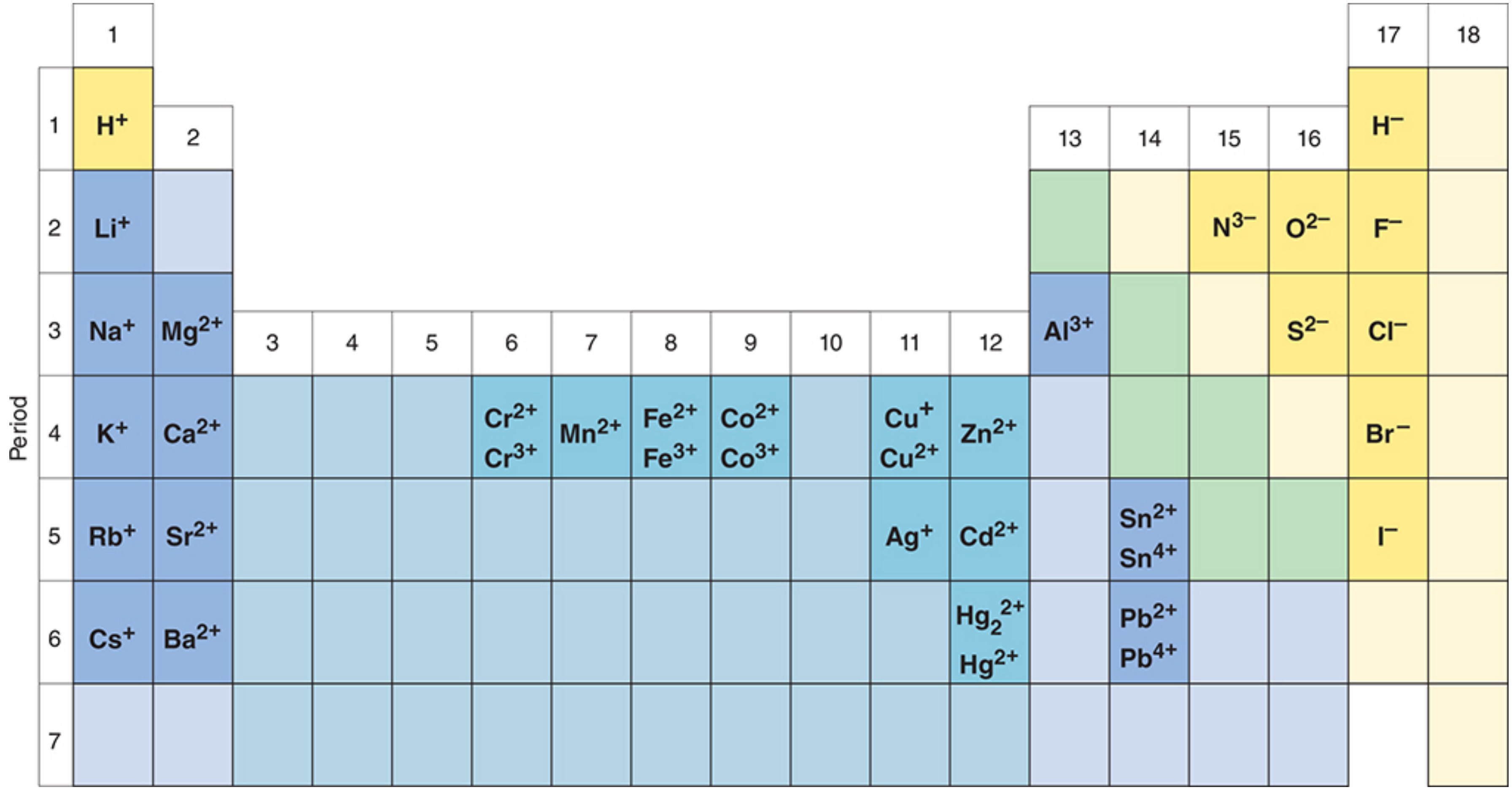
Predicting ions

- atoms tend to gain/lose electrons to form ions with the same number of electrons as that of the *nearest noble gas*

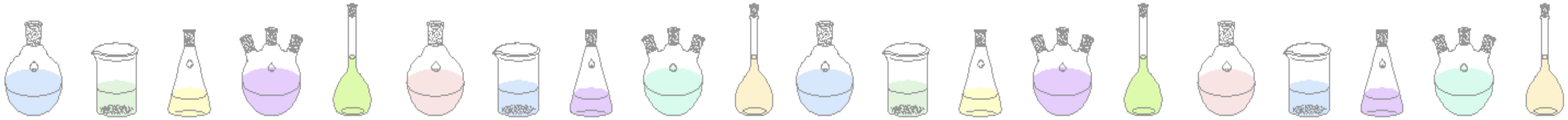




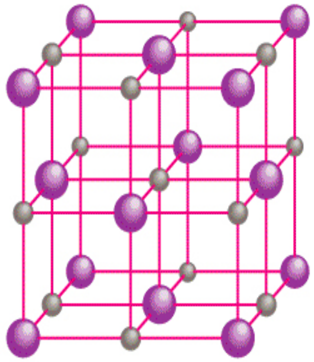
Predicting ions – Fig. 2.17

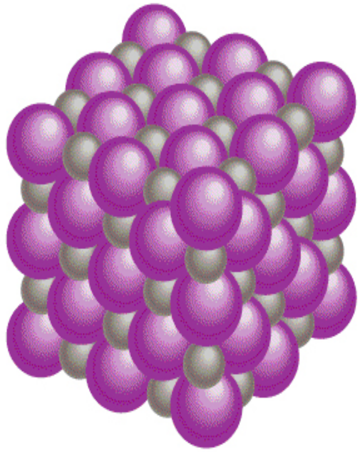


Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H^+																H^-	
2	Li^+														N^{3-}	O^{2-}	F^-	
3	Na^+	Mg^{2+}											Al^{3+}			S^{2-}	Cl^-	
4	K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}		Cu^+ Cu^{2+}	Zn^{2+}					Br^-	
5	Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}			I^-	
6	Cs^+	Ba^{2+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				
7																		



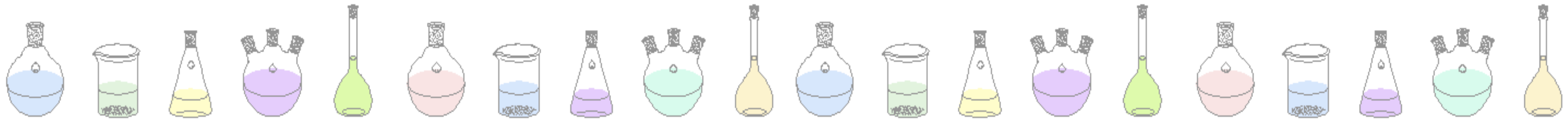
Ionic compounds


(a)


(b)

- there are no distinct molecules
- ions combine to form an ionic lattice
- as shown in the figure, each Na^+ is associated with 6 Cl^-, and each Cl^- is associated with 6 Na^+

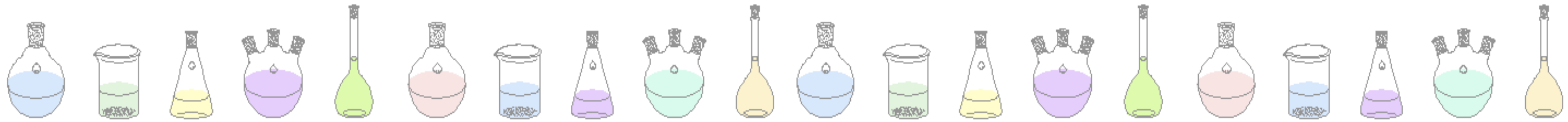
formula unit = NaCl



Ionic compounds

- NaCl = an ionic compound formed of two ions with charges of +1 and -1
- however, other ionic compounds can have different lattice structures with different charges, but:
- the overall crystal must be electrically neutral!

$$Al^{3+} + O^{2-} = Al_2O_3$$



Naming Binary Ionic Compounds

- name the cation first, then the anion
- prefixes are avoided when possible
- if necessary, the oxidation state of the cation is added as a Roman numeral in brackets

KBr	potassium bromide
$CaCl_2$	calcium chloride
Al_2O_3	aluminum oxide
FeI_2	iron (II) iodide

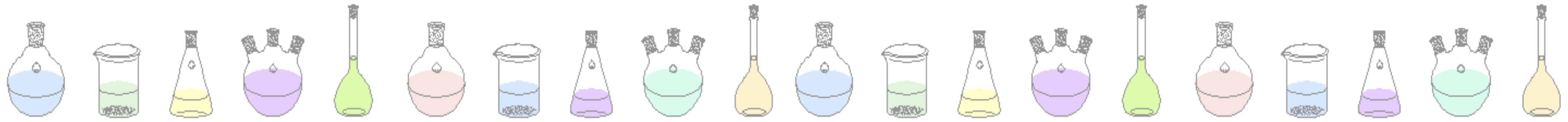


Table 2.4

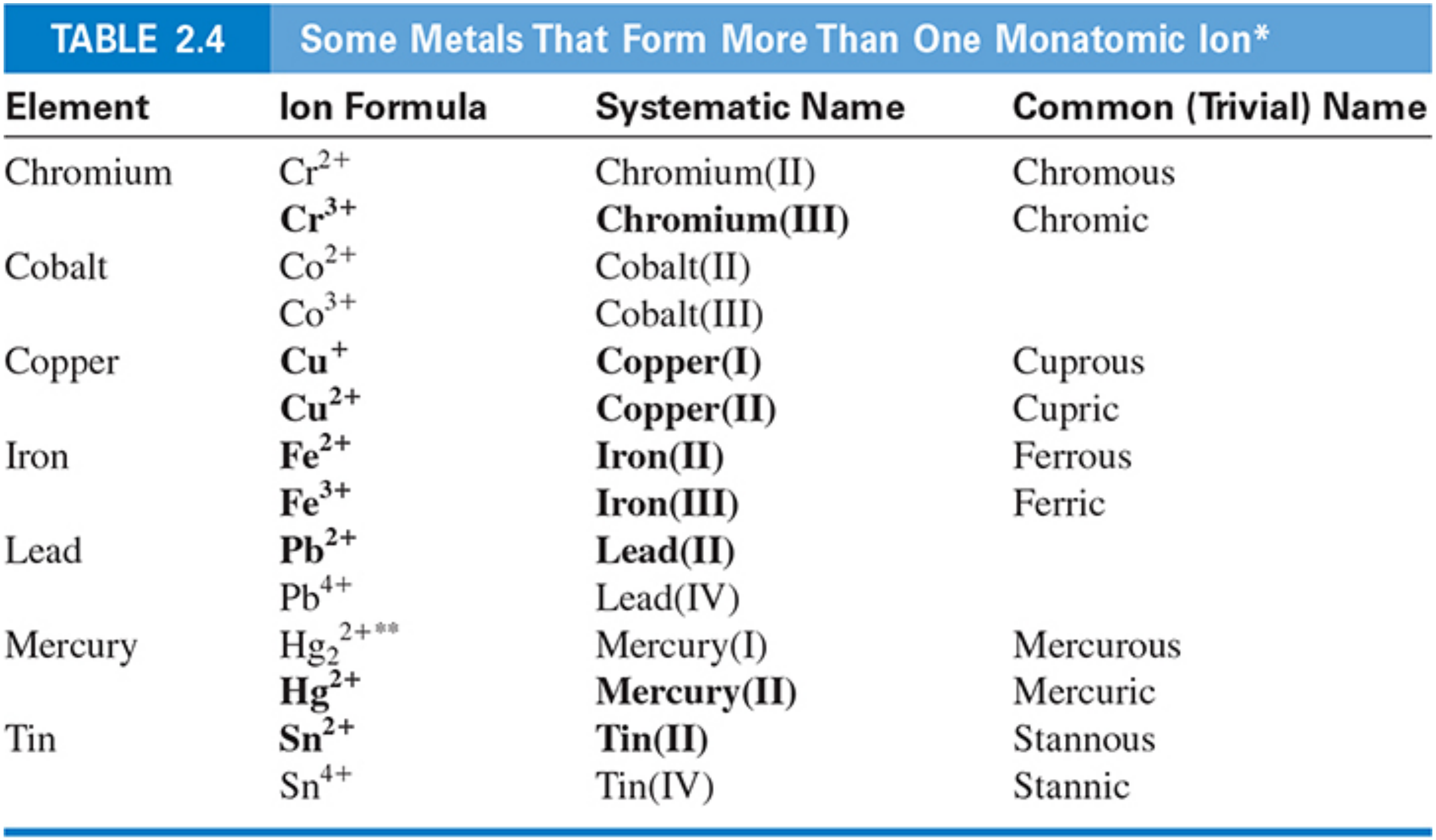
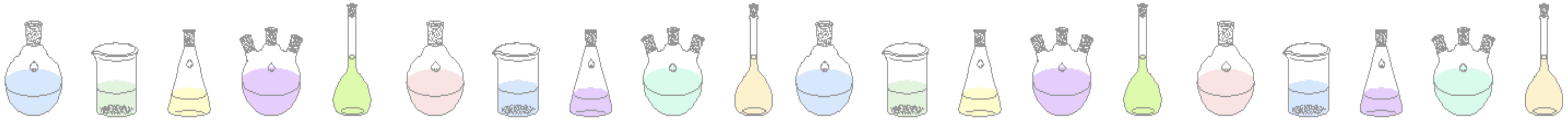


TABLE 2.4 **Some Metals That Form More Than One Monatomic Ion***

Element	Ion Formula	Systematic Name	Common (Trivial) Name
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	**Chromium(III)**	Chromic
Cobalt	Co^{2+}	Cobalt(II)	
	Co^{3+}	Cobalt(III)	
Copper	**Cu^{+}**	**Copper(I)**	Cuprous
	Cu^{2+}	**Copper(II)**	Cupric
Iron	**Fe^{2+}**	**Iron(II)**	Ferrous
	Fe^{3+}	**Iron(III)**	Ferric
Lead	**Pb^{2+}**	**Lead(II)**	
	Pb^{4+}	Lead(IV)	
Mercury	Hg_2^{2+**}	Mercury(I)	Mercurous
	Hg^{2+}	**Mercury(II)**	Mercuric
Tin	**Sn^{2+}**	**Tin(II)**	Stannous
	Sn^{4+}	Tin(IV)	Stannic



Naming Ionic Compounds

- Sometimes, the anion or cation is polyatomic:
  - NH_4^+ : ammonium
  - CO_3^{2-} : carbonate
  - OH^- : hydroxide

- polyatomic ions stay together as a charged unit

Polyatomic Ions

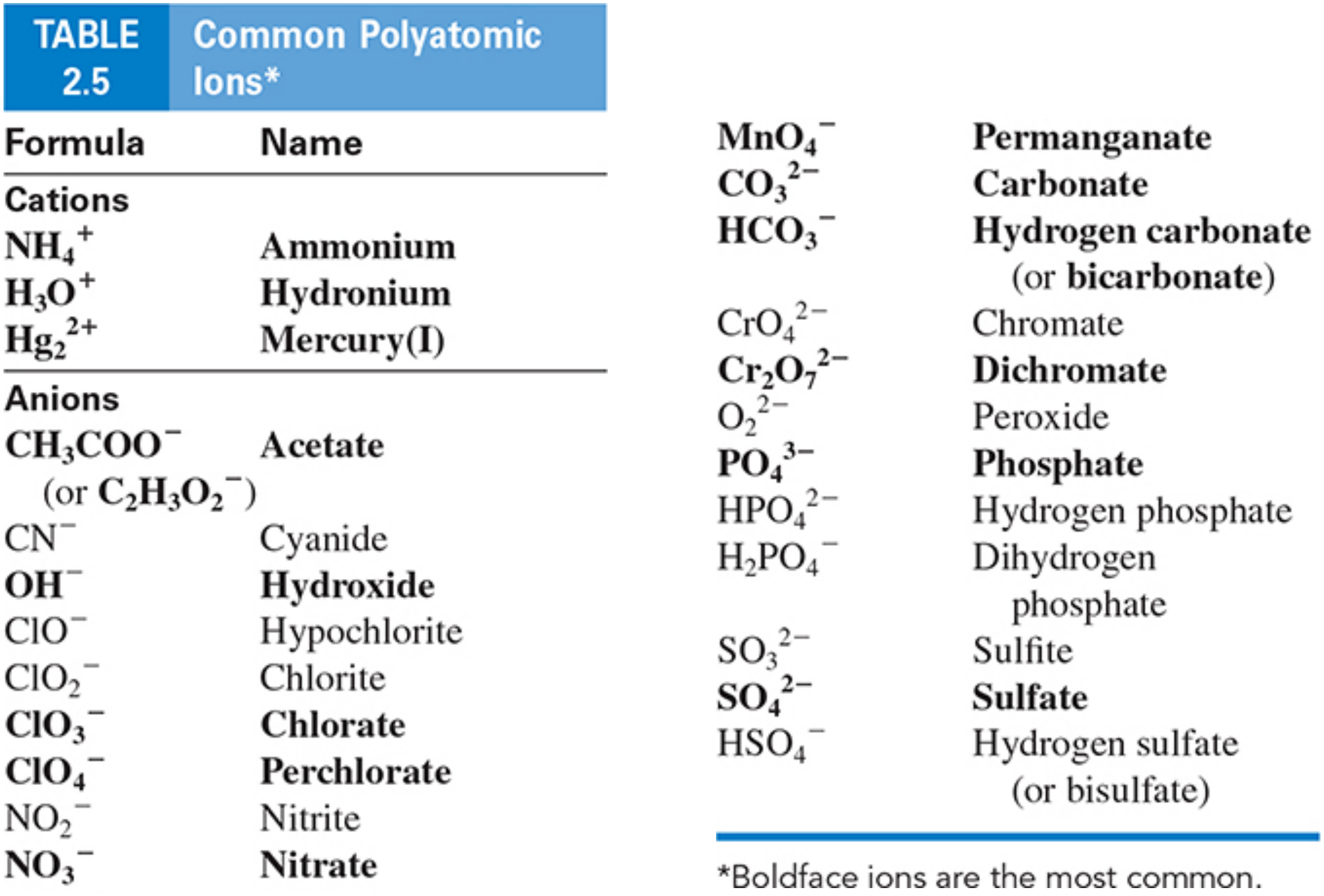


TABLE 2.5 Common Polyatomic Ions*

Formula	Name
Cations	
NH_4^+	**Ammonium**
H_3O^+	**Hydronium**
Hg_2^{2+}	**Mercury(I)**
Anions	
CH_3COO^- (or **$C_2H_3O_2^-$**)	**Acetate**
CN^-	Cyanide
OH^-	**Hydroxide**
ClO^-	Hypochlorite
ClO_2^-	Chlorite
ClO_3^-	**Chlorate**
ClO_4^-	**Perchlorate**
NO_2^-	Nitrite
NO_3^-	**Nitrate**
MnO_4^-	**Permanganate**
CO_3^{2-}	**Carbonate**
HCO_3^-	**Hydrogen carbonate** (or **bicarbonate**)
CrO_4^{2-}	Chromate
$Cr_2O_7^{2-}$	**Dichromate**
O_2^{2-}	Peroxide
PO_4^{3-}	**Phosphate**
HPO_4^{2-}	Hydrogen phosphate
$H_2PO_4^-$	Dihydrogen phosphate
SO_3^{2-}	Sulfite
SO_4^{2-}	**Sulfate**
HSO_4^-	Hydrogen sulfate (or bisulfate)

*Boldface ions are the most common.

Naming inorganic compounds

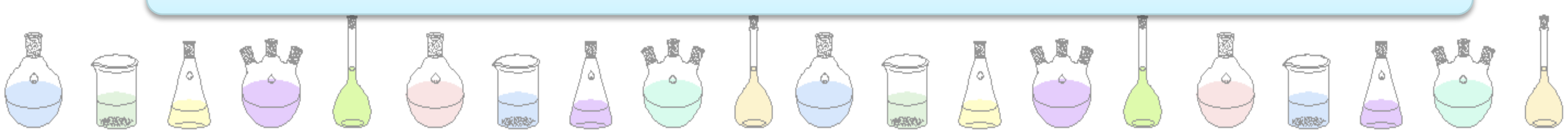
- Some examples:

KNO_3	potassium nitrate
NH_4Cl	ammonium chloride
$KMnO_4$	potassium permanganate
$Ca(SCN)_2$	calcium thiocyanate

- Hydrates:
  - an ionic compound containing a fixed number of molecules of water

$CaCl_2 \cdot 6H_2O$ = calcium chloride hexahydrate

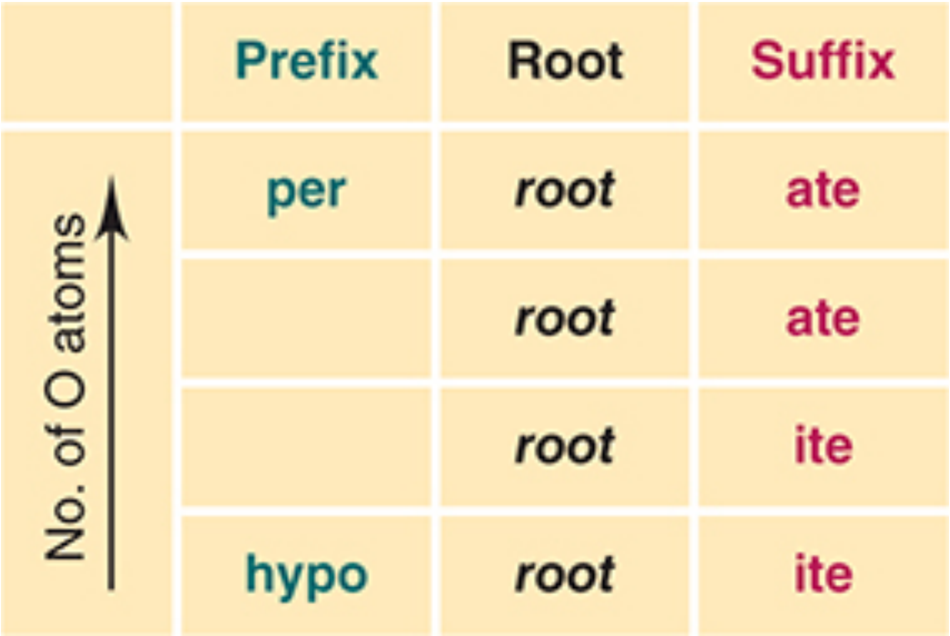
$MgSO_4 \cdot 2H_2O$ = magnesium sulfate dihydrate



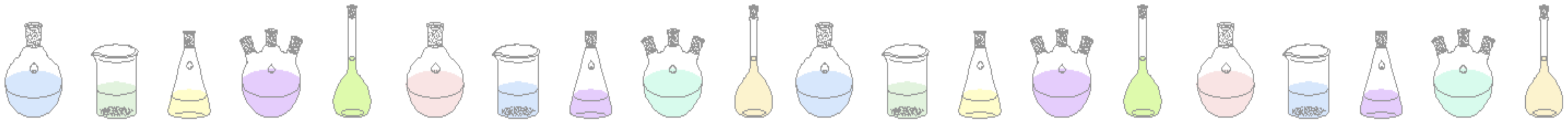
Oxoanions

- polyatomic ions of the general formula XO_n^{m-}

- the suffixes –*ite* and -*ate* and the prefixes *hypo-* and *per-* are used to indicate the oxidation state of the nonmetal atom



	Prefix	Root	Suffix
No. of O atoms ↑	per	*root*	ate
		root	ate
		root	ite
	hypo	*root*	ite



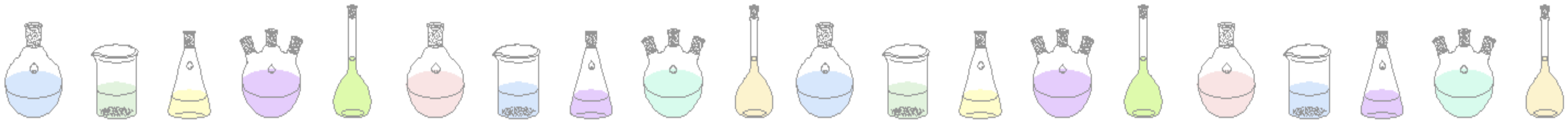
Oxoanions

- examples:

Increasing number of oxygen atoms

SO_4^{2-}	sulfate
SO_3^{2-}	sulfite

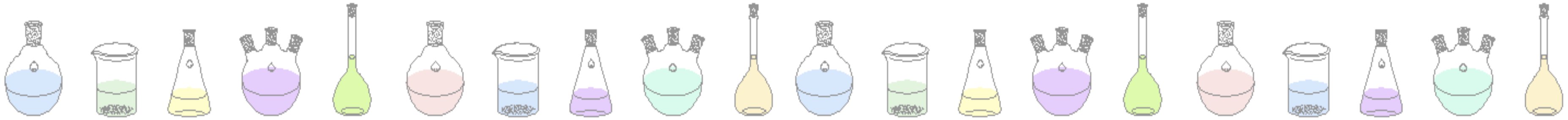
ClO_4^-	perchlorate
ClO_3^-	chlorate
ClO_2^-	chlorite
ClO^-	hypochlorite



Naming acids

- Brønsted-Lowry definition: acids are *proton donors*
- Binary acids (or hydroacids)
  - a hydroacid is a compound with the general formula **H_nX** that ionizes in water
  - they are named used the prefix *hydro-* and the suffix *–ic*:

HF	hydrofluoric acid
HI	hydroiodic acid
H_2S	hydrosulfuric acid



Naming oxoacids

- an oxoacid is a compound with the general formula H_mXO_n that ionizes in water
- oxoacid = oxoanion + H^+ ions
- their names are based on the oxoanion from which they are formed

–ite becomes *–ous*

–ate becomes *–ic*

Naming oxoacids

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

$HClO$	hypochlorous acid
$HClO_2$	chlorous acid
$HClO_3$	chloric acid
$HClO_4$	perchloric acid

Masses from Chemical Formulae

- covalent compounds:

molecular mass = sum of atomic mass of each atom in the molecule (in amu)

- ionic compounds:

formula mass = sum of atomic mass of each atom in the formula unit (in amu)

Masses from Chemical Formulae

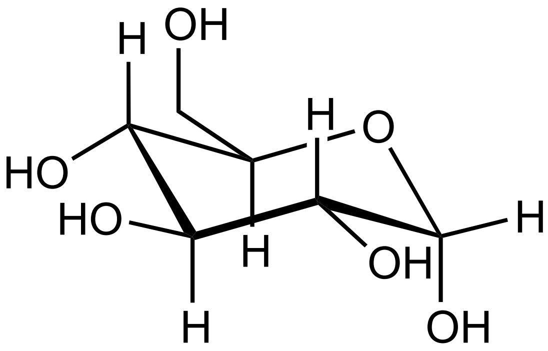
What is the *molecular mass* of H_2O?

$$2 \text{ atoms H} \times \frac{1.008 \text{ amu}}{1 \text{ H atom}} = 2.016 \text{ amu}$$

$$1 \text{ atom O} \times \frac{15.999 \text{ amu}}{1 \text{ O atom}} = 15.999 \text{ amu}$$

1 molecule H_2O = 2.016 amu + 15.999 amu = 18.015 amu

Molecular mass vs. Exact mass



Glucose

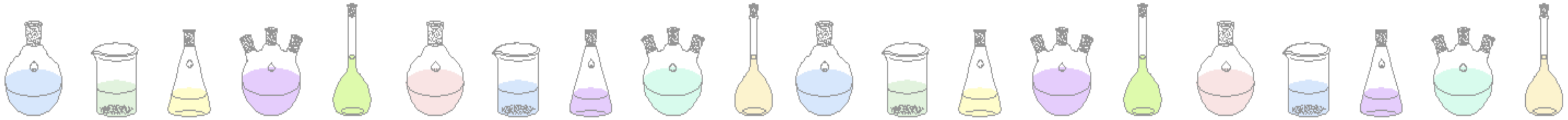
Molecular formula: $C_6H_{12}O_6$

Molecular Mass: Use the naturally occurring mixture of isotopes

$6 \times 12.01 + 12 \times 1.01 + 6 \times 16.00 = \mathbf{180.18}$ **amu**

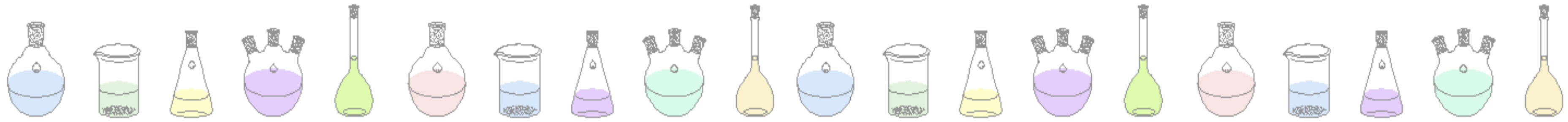
Exact Mass: Use the most abundant isotopes

$6 \times 12.000000 + 12 \times 1.007825 + 6 \times 15.994915 = \mathbf{180.06339}$ **amu**



Chapter 2: Key Concepts

- modern model of the atom
- elements and their isotopes
- atomic mass
- the Periodic Table
- naming inorganic compounds



Chapter 2: Suggested Problems

B2.1, B2.2, 2.3, 2.4, 2.20,
2.24, 2.39, 2.43, 2.49, 2.88,
2.90, 2.92, 2.94, 2.96, 2.100,
2.106, 2.110, 2.124, 2.128,
2.131, 2.135, 2.136