General Chemistry

لجان الدفعات - كليّة الصيدلة – الجامعة الأردنية

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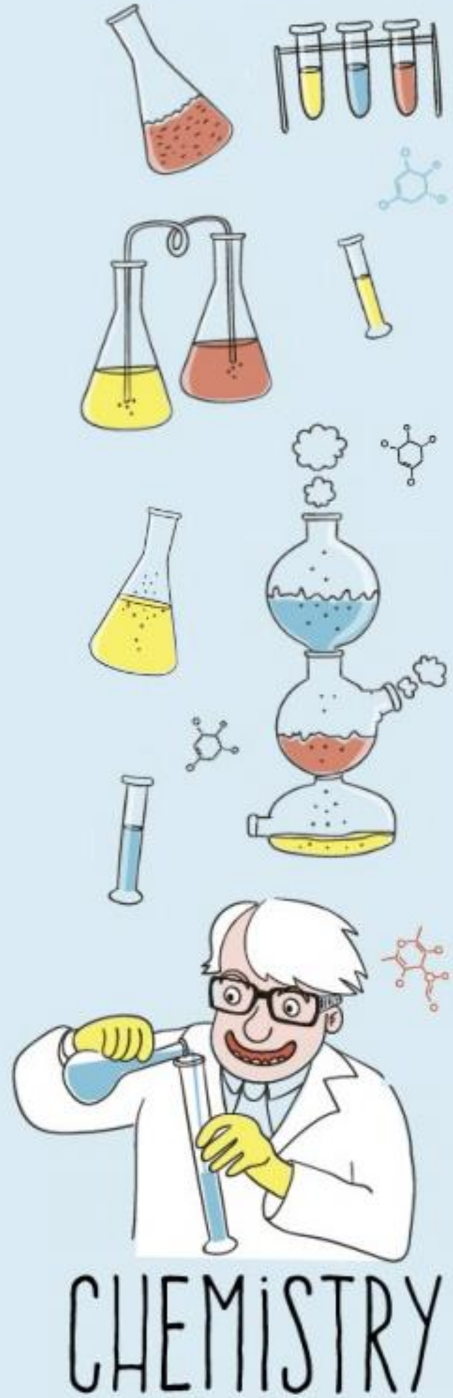
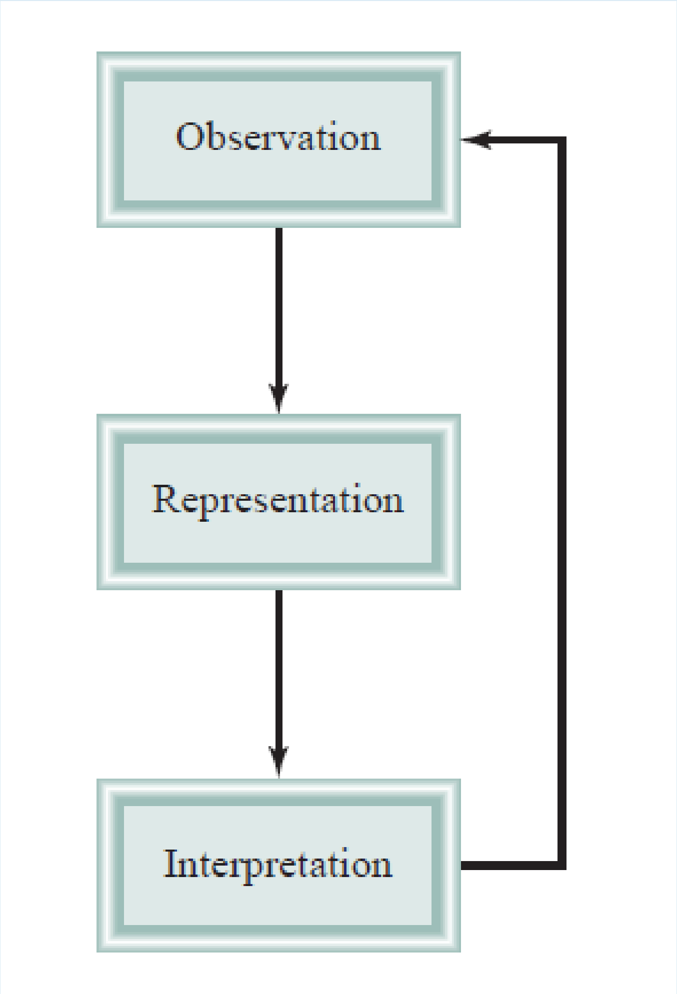


Chapter 1: Introduction

Scientific Method:



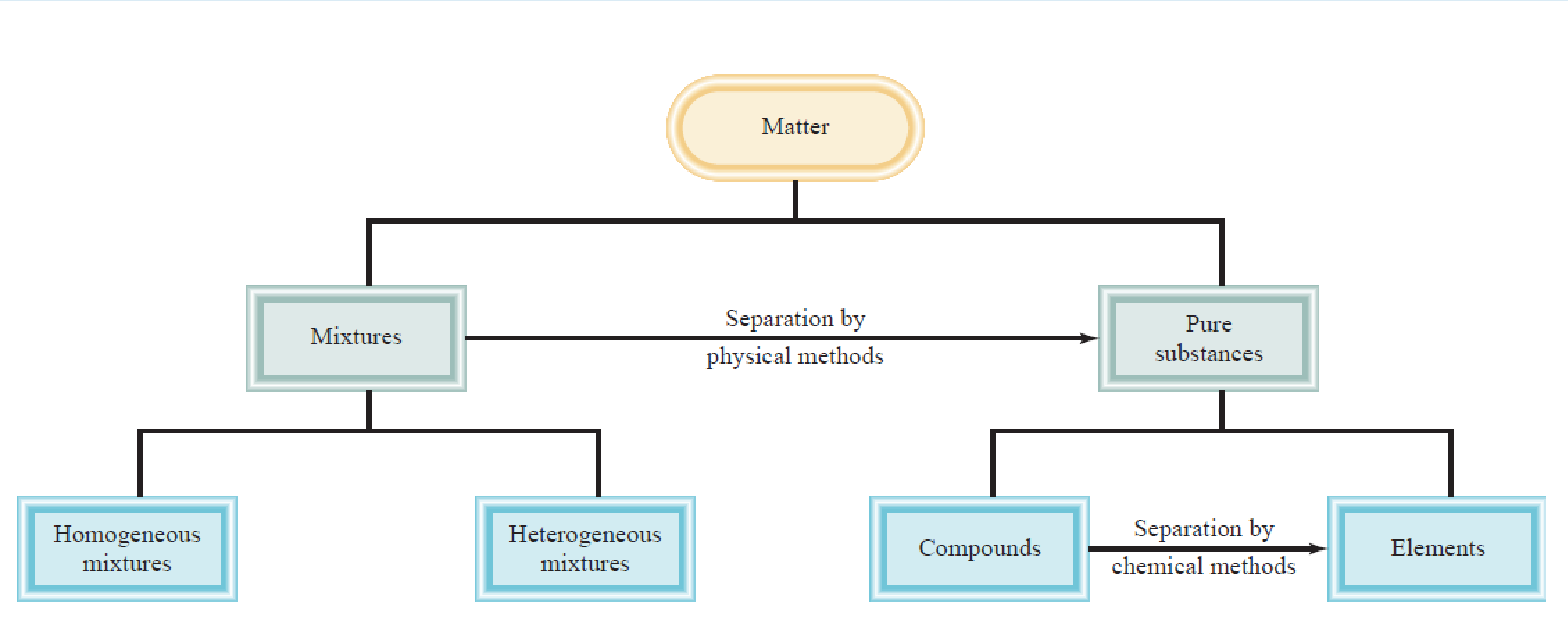
Scientific Method: Systematic approach of research.

Observation: of a problem. It includes collecting quantitative and qualitative data.

Representation: Formation of a law verbally (words) or mathematically (symbols and equations).

Interpretation: Generating hypothesis and consequent theories.

Classification of Matter



Properties: (Physical vs. chemical) , (intensive temperature vs. extensive mass) and (macroscopic vs. microscopic).

TABLE 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bromine	Br	Iodine	I	Potassium	K
Calcium	Ca	Iron	Fe	Silicon	Si
Carbon	C	Lead	Pb	Silver	Ag
Chlorine	Cl	Magnesium	Mg	Sodium	Na
Chromium	Cr	Mercury	Hg	Sulfur	S
Cobalt	Co	Nickel	Ni	Tin	Sn
Copper	Cu	Nitrogen	N	Zinc	Zn

Measurements:

SI : International System of Units.

TABLE 1.2 SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

TABLE 1.3 **Prefixes Used with SI Units**

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

Q1.

Example 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3. Calculate the density of gold.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

$$d = \frac{m}{V}$$

$$= \frac{301\ \text{g}}{15.6\ \text{cm}^3}$$

$$= \boxed{19.3\ \text{g/cm}^3}$$

Practice Exercise A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

Temperature:

$$°F = (°C \times \frac{9}{5}) + 32$$

$$°C = (°F - 32) \times \frac{5}{9}$$

$$°C = (K - 273.15)$$

$$K = (°C + 273.15)$$

Q2.

Example 1.2

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at −452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at −38.9°C. Convert its melting point to kelvins.

Solution These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This conversion is carried out by writing

$$\frac{9°\text{F}}{5°\text{C}} \times (224°\text{C}) + 32°\text{F} = 435°\text{F}$$

(b) Here we have

$$(-452°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = -269°\text{C}$$

(c) The melting point of mercury in kelvins is given by

$$(-38.9°\text{C} + 273.15°\text{C}) \times \frac{1\text{ K}}{1°\text{C}} = 234.3\text{ K}$$

Practice Exercise Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K, the boiling point of liquid nitrogen, to degrees Celsius.

Handling Numbers:

Scientific Notation: $N \times 10^n$

Addition and subtraction: unify the exponent (n) of 10, add or subtract N(s).

Multiplication: add the exponents of 10, multiply N(s).

Division: subtract the exponents of 10, divide N(s).

$$
\begin{aligned}
(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3 \\
(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4) \\
&= 4.70 \times 10^4 \\
(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) \\
&= 1.81 \times 10^{-2}
\end{aligned}
$$

$$
\begin{aligned}
(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2}) \\
&= 40 \times 10^6 \\
&= 4.0 \times 10^7 \\
(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3}) \\
&= 28 \times 10^{-2} \\
&= 2.8 \times 10^{-1} \\
\frac{6.9 \times 10^7}{3.0 \times 10^{-5}} &= \frac{6.9}{3.0} \times 10^{7-(-5)} \\
&= 2.3 \times 10^{12} \\
\frac{8.5 \times 10^4}{5.0 \times 10^9} &= \frac{8.5}{5.0} \times 10^{4-9} \\
&= 1.7 \times 10^{-5}
\end{aligned}
$$

Significant Figures:

Rules:

1. Any digit that is not zero is significant.
2. Zeros between nonzero digits are significant.
3. Zeros to the left of nonzero digits are insignificant (with or without decimal points).
4. Zeros to the right of nonzero digits:
   a. Number with decimal points= significant.
   b. Number without decimal points =

~~Questionable!~~ scientific notation is not counted!

5. Exact numbers = infinite numbers of significant figures!

Example 1.3

Determine the number of significant figures in the following measurements: (a) 478 cm, (b) 6.01 g, (c) 0.825 m, (d) 0.043 kg, (e) 1.310×10^{22} atoms, (f) 7000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four, because the number is greater than one, all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (7.000×10^3), three (7.00×10^3), two (7.0×10^3), or one (7×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Practice Exercise Determine the number of significant figures in each of the following measurements: (a) 24 mL, (b) 3001 g, (c) 0.0320 m^3, (d) 6.4×10^4 molecules, (e) 560 kg.

Rounding:

Addition and Subtraction: the answer cannot have more digits to the right of the decimal point than either of the original numbers (i.e. equal to the one with the smaller number of digits).

Multiplication and Division: the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures.

Round the final answer!

Q3.

Example 1.4

Carry out the following arithmetic operations to the correct number of significant figures: (a) 11,254.1 g + 0.1983 g, (b) 66.59 L − 3.113 L, (c) 8.16 m × 5.1355, (d) 0.0154 kg ÷ 88.3 mL, (e) 2.64×10^3 cm + 3.27×10^2 cm.

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r} 11{,}254.1 \text{ g} \\ +\quad 0.1983 \text{ g} \\ \hline 11{,}254.2983 \text{ g} \end{array} \longleftarrow \text{round off to } 11{,}254.3 \text{ g}$$

(b)
$$\begin{array}{r} 66.59 \text{ L} \\ -\ 3.113 \text{ L} \\ \hline 63.477 \text{ L} \end{array} \longleftarrow \text{round off to } 63.48 \text{ L}$$

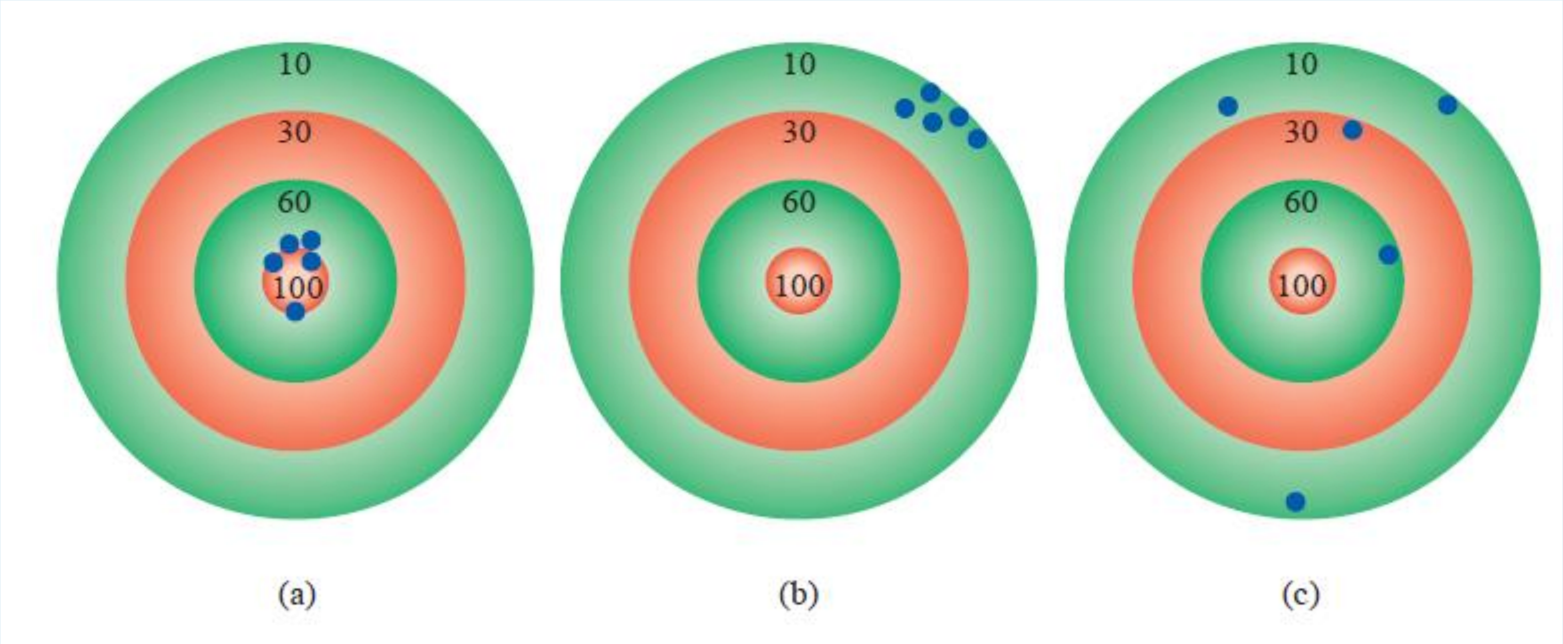
(c) $8.16 \text{ m} \times 5.1355 = 41.90568 \text{ m} \longleftarrow$ round off to 41.9 m

(d) $\dfrac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.000174405436 \text{ kg/mL} \longleftarrow$ round off to 0.000174 kg/mL or 1.74×10^{-4} kg/mL

(e) First we change 3.27×10^2 cm to 0.327×10^3 cm and then carry out the addition $(2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is 2.97×10^3 cm.

Practice Exercise Carry out the following arithmetic operations and round off the answers to the appropriate number of significant figures: (a) 26.5862 L + 0.17 L, (b) 9.1 g − 4.682 g, (c) 7.1×10^4 dm × 2.2654×10^2 dm, (d) 6.54 g ÷ 86.5542 mL, (e) $(7.55 \times 10^4 \text{ m}) - (8.62 \times 10^3 \text{ m})$.

Accuracy and Precision:



	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

Dimensional Analysis:
" Factor Label Method "

X Required (Desired) = Given * Conversion Factor

X Required (Desired) = Given * $\frac{Desired}{Given}$

X Required (Desired) = ~~Given~~ * $\frac{Desired}{\cancel{Given}}$

X Required (Desired) = Desired !

Ex. Convert 10 m to cm ? 1000

Ex. Convert 5 kg/m^3 to g/cm^3 ? $5 * 10^{-3}$

Example 1.5

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams ($1 \text{ mg} = 1 \times 10^{-3} \text{ g}$). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors:

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = 3.78 \times 10^4 \text{ mg}$$

Check As an estimate, we note that 1 lb is roughly 500 g and that 1 g = 1000 mg. Therefore, 1 lb is roughly 5×10^5 mg. Rounding off 0.0833 lb to 0.1 lb, we get 5×10^4 mg, which is close to the preceding quantity.

Practice Exercise A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

Q4.

Example 1.6

An average adult has 5.2 L of blood. What is the volume of blood in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 5.2 \text{ L}$$

How many conversion factors are needed for this problem? Recall that $1 \text{ L} = 1000 \text{ cm}^3$ and $1 \text{ cm} = 1 \times 10^{-2} \text{ m}$.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}$$

Because the second conversion factor deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} = \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3$$

This means that $1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$. Now we can write

$$? \text{ m}^3 = 5.2 \cancel{\text{L}} \times \frac{1000 \text{ cm}^3}{1 \cancel{\text{L}}} \times \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3 = 5.2 \times 10^{-3} \text{ m}^3$$

Check From the preceding conversion factors you can show that $1 \text{ L} = 1 \times 10^{-3} \text{ m}^3$. Therefore, 5 L of blood would be equal to $5 \times 10^{-3} \text{ m}^3$, which is close to the answer.

Practice Exercise The volume of a room is $1.08 \times 10^8 \text{ dm}^3$. What is the volume in m^3?

Example 1.7

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research. The density of the liquid at its boiling point (−196°C or 77 K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

Strategy The Problem can be stated as

$$? \text{ kg/m}^3 = 0.808 \text{ g/cm}^3$$

Two separate conversions are required for this problem: $g \longrightarrow kg$ and $cm^3 \longrightarrow m^3$. Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m.

Solution In Example 1.6 we saw that 1 cm^3 = 1×10^{-6} m^3. The conversion factors are

$$\frac{1 \text{ kg}}{1000 \text{ g}} \quad \text{and} \quad \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3}$$

Finally,

$$? \text{ kg/m}^3 = \frac{0.808 \cancel{\text{g}}}{1 \cancel{\text{cm}^3}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} \times \frac{1 \cancel{\text{cm}^3}}{1 \times 10^{-6} \text{ m}^3} = \boxed{808 \ kg/m^3}$$

Check Because 1 m^3 = 1×10^6 cm^3, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

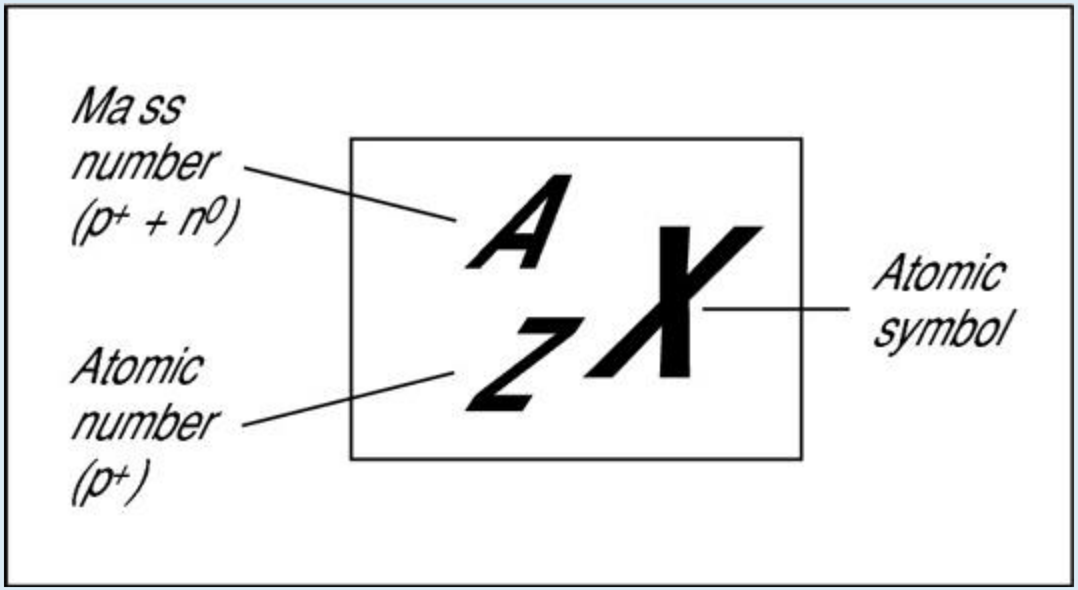
Practice Exercise The density of the lightest metal, lithium (Li), is 5.34×10^2 kg/m^3. Convert the density to g/cm^3.

Answers to Practice Exercises

1.1 96.5 g. **1.2** (a) 621.5°F, (b) 78.3°C, (c) -196°C. **1.3** (a) Two, (b) four, (c) three, (d) two, (e) three or two. **1.4** (a) 26.76 L, (b) 4.4 g, (c) 1.6×10^7 dm^2, (d) 0.0756 g/mL, (e) 6.69×10^4 m. **1.5** 2.36 lb. **1.6** 1.08×10^5 m^3. **1.7** 0.534 g/cm^3.

Chapter 2: Atoms, Molecules and Ions

Atoms:



Mass Number (atomic mass or weight) – Atomic Number = Neutrons^0

Protons = electrons. Except:

Cations$^+$ = positively charged ions.

Anions$^-$ = negatively charged ions.

Isotopes.

Example 2.1

Give the number of protons, neutrons, and electrons in each of the following species: (a) $^{17}_{8}O$, (b) $^{199}_{80}Hg$, and (c) $^{200}_{80}Hg$.

Strategy Recall that the superscript denotes mass number and the subscript denotes atomic number. Mass number is always greater than atomic number. (The only exception is $^{1}_{1}H$, where the mass number is equal to the atomic number.)

Solution

(a) The atomic number is 8, so there are 8 protons. The mass number is 17, so the number of neutrons is $17 - 8 = 9$. The number of electrons is the same as the number of protons, that is, 8.

(b) The atomic number is 80, so there are 80 protons. The mass number is 199, so the number of neutrons is $199 - 80 = 119$. The number of electrons is 80.

(c) Here the number of protons is the same as in (b), or 80. The number of neutrons is $200 - 80 = 120$. The number of electrons is also the same as in (b), 80. The species in (b) and (c) are chemically similar isotopes of mercury.

Practice Exercise How many protons, neutrons, and electrons are in the following isotope of copper: $^{63}_{29}Cu$?

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	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 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)

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

Groups (families-verticals) and Rows (Periods-horizontal).

You have to know groups A1-A8:

A1: Alkali metals (M+, loss of an electron) (H,Li,Na,k)

A2: Alkaline earth metals (M++, loss of two electrons) as (Be, Mg, ca).

A3: X+++ (B,Al)

A4: X ? (C,si, Se)

A5: X--- (N, P)

A6: X- - (O, S,Se)

A7: Halogens (X-, gain of an electron) (F, Cl, Br ,I)

A8: Nobel gases (He, Ne, Ar, Kr, Xe, Rn)

B Groups: Transition Metals (different forms of electron loss M*). (Cu, Fe, Hg)

Metalloids: in-between:

Characteristics: Metals

Electron behavior: Nonmetals.

Types of Bonds:

Ionic Bonds: metal +nonmetal (or metalloids).
Covalent Bonds: nonmetal +nonmetal (molecules)

Naming of Compounds:

According to compound's type:

1. Ionic compound.
2. Molecular compound.
3. Acid/base.

1. Ionic Compounds:

$$X^{+}Y^{-}$$

X: cation, positively charged.

Y: anion, negatively charged.

→Name the cation (metal) then name the anion (nonmetal/metalloid) ended by –ide suffix if it's an element or by its name if it's not a single element (ionic group).

Ex. NaCl, $(NH_4)_2SO_4$ and $Al(Br)_3$.

Potassium Sulfate, Cupric oxide and Calcium Phosphate.

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)

mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^{+})	permanganate (MnO_4^-)
rubidium (Rb^{+})	peroxide (O_2^{2-})
silver (Ag^{+})	phosphate (PO_4^{3-})
sodium (Na^{+})	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

TABLE 2.2 The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C Carbide (C^{4-})*	N Nitride (N^{3-})	O Oxide (O^{2-})	F Fluoride (F^-)
Si Silicide (Si^{4-})	P Phosphide (P^{3-})	S Sulfide (S^{2-})	Cl Chloride (Cl^-)
		Se Selenide (Se^{2-})	Br Bromide (Br^-)
		Te Telluride (Te^{2-})	I Iodide (I^-)

The last table is not to be memorized, just get familiar with anions names.

Example 2.4

Name the following compounds: (a) $Cu(NO_3)_2$, (b) KH_2PO_4, and (c) NH_4ClO_3.

Strategy Our reference for the names of cations and anions is Table 2.3. Keep in mind that if a metal can form cations of different charges (see Figure 2.10), we need to use the Stock system.

Solution

(a) The nitrate ion (NO_3^-) bears one negative charge, so the copper ion must have two positive charges. Because copper forms both Cu^+ and Cu^{2+} ions, we need to use the Stock system and call the compound copper(II) nitrate.

(b) The cation is K^+ and the anion is $H_2PO_4^-$ (dihydrogen phosphate). Because potassium only forms one type of ion (K^+), there is no need to use potassium(I) in the name. The compound is potassium dihydrogen phosphate.

(c) The cation is NH_4^+ (ammonium ion) and the anion is ClO_3^-. The compound is ammonium chlorate.

Practice Exercise Name the following compounds: (a) PbO and (b) Li_2SO_3.

Example 2.5

Write chemical formulas for the following compounds: (a) mercury(I) nitrite, (b) cesium sulfide, and (c) calcium phosphate.

Strategy We refer to Table 2.3 for the formulas of cations and anions. Recall that the Roman numerals in the Stock system provide useful information about the charges of the cation.

Solution

(a) The Roman numeral shows that the mercury ion bears a +1 charge. According to Table 2.3, however, the mercury(I) ion is diatomic (that is, Hg_2^{2+}) and the nitrite ion is NO_2^-. Therefore, the formula is $Hg_2(NO_2)_2$.

(b) Each sulfide ion bears two negative charges, and each cesium ion bears one positive charge (cesium is in Group 1A, as is sodium). Therefore, the formula is Cs_2S.

(c) Each calcium ion (Ca^{2+}) bears two positive charges, and each phosphate ion (PO_4^{3-}) bears three negative charges. To make the sum of the charges equal zero, we must adjust the numbers of cations and anions:

$$3(+2) + 2(-3) = 0$$

Thus, the formula is $Ca_3(PO_4)_2$.

Practice Exercise Write formulas for the following ionic compounds: (a) rubidium sulfate and (b) barium hydride.

2. Molecular Compounds (covalent bonds):

$$X_n\text{-}Y_z\text{-}$$

It contains ionic groups and most of acids.
We use prefixes here to indicate number of each Element in the compound.

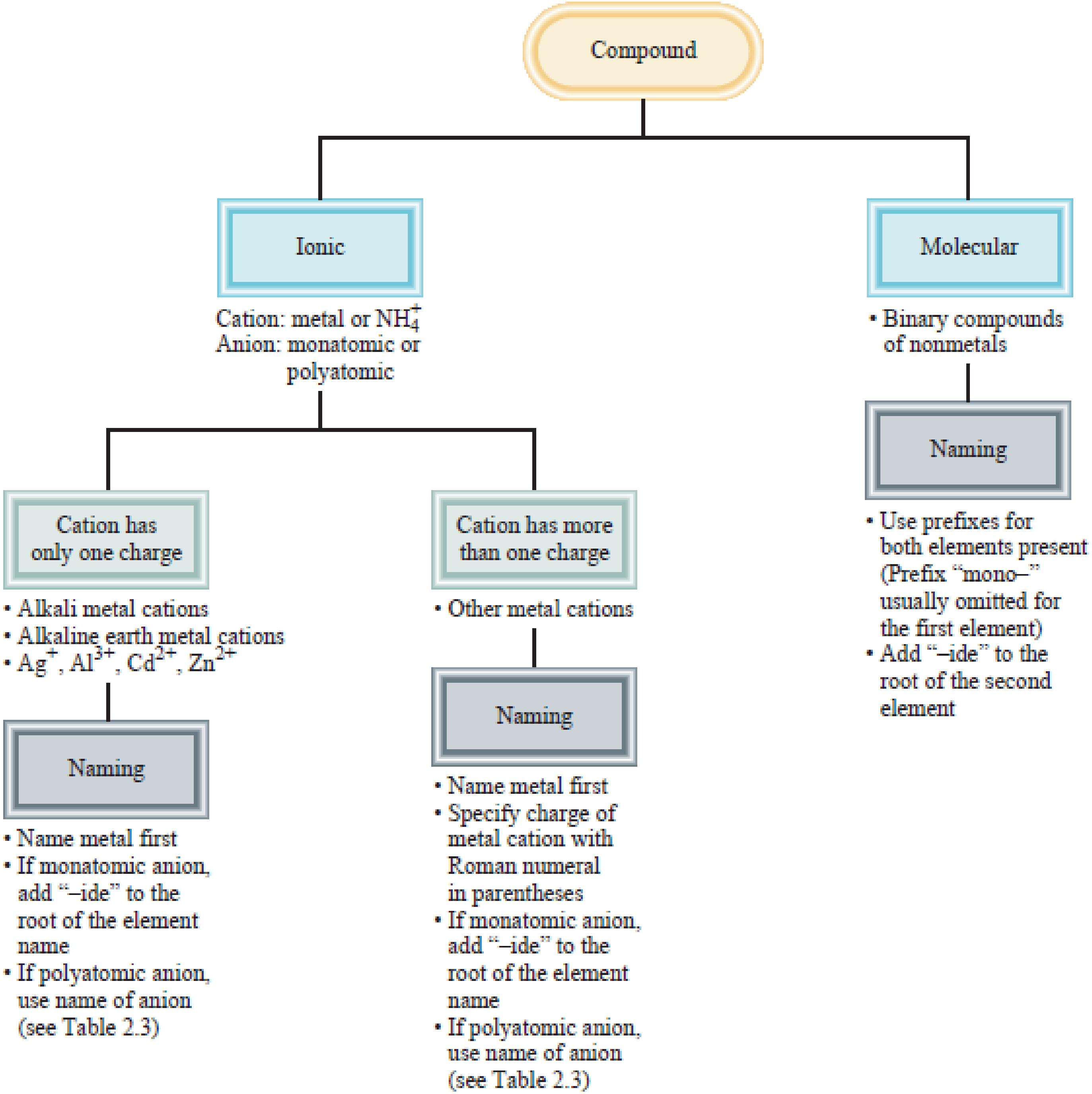
→Prefix of (n) name of element X , prefix (z) name of element Y ended by –ide (as anion).

Ex. N_2O_4 and SO_3

TABLE 2.4

Greek Prefixes Used in Naming Molecular Compounds

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


Compound
Ionic
Cation: metal or NH_4^+
Anion: monatomic or polyatomic
Molecular
• Binary compounds of nonmetals
Naming
• Use prefixes for both elements present (Prefix "mono–" usually omitted for the first element)
• Add "–ide" to the root of the second element
Cation has only one charge
• Alkali metal cations
• Alkaline earth metal cations
• Ag^+, Al^{3+}, Cd^{2+}, Zn^{2+}
Naming
• Name metal first
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)
Cation has more than one charge
• Other metal cations
Naming
• Name metal first
• Specify charge of metal cation with Roman numeral in parentheses
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)

Example 2.6

Name the following molecular compounds: (a) $SiCl_4$ and (b) P_4O_{10}.

Solution We refer to Table 2.4 for prefixes.

(a) Because there are four chlorine atoms present, the compound is silicon tetrachloride.

(b) There are four phosphorus atoms and ten oxygen atoms present, so the compound is tetraphosphorus decoxide. Note that the "a" is omitted in "deca."

Practice Exercise Name the following molecular compounds: (a) NF_3 and (b) Cl_2O_7.

Example 2.7

Write chemical formulas for the following molecular compounds: (a) carbon disulfide and (b) disilicon hexabromide.

Solution We refer to Table 2.4 for prefixes.

(a) Because there are two sulfur atoms and one carbon atom present, the formula is CS_2.

(b) There are two silicon atoms and six bromine atoms present, so the formula is Si_2Br_6.

Practice Exercise Write chemical formulas for the following molecular compounds: (a) sulfur tetrafluoride and (b) dinitrogen pentoxide.

3. Naming Acids:

Usually the first element is Hydrogen.

Name it as a molecular compound:

- If ended with –ide : add (hydro) prefix and replace –ide with –ic acid
- If ended with –ate: replace –ate with –ic acid.
- If ended with –ite: replace –its with –ous acid.

Per, -ic (reference) , -ous and hypo. → iodate (IO3)

Ex. HCl, $HClO_3$ and HClO

Example 2.8

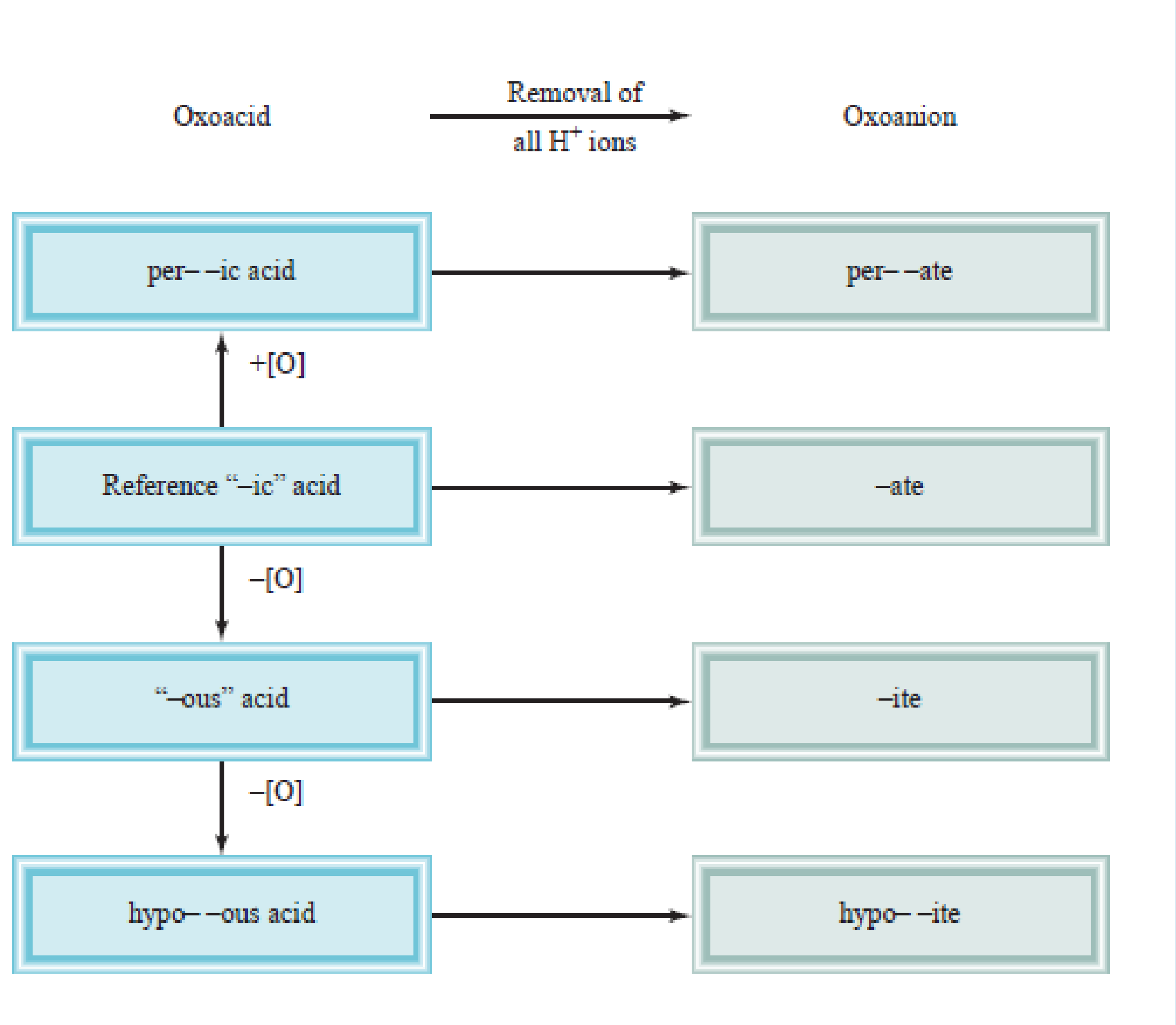
Name the following oxoacid and oxoanion: (a) H_3PO_3 and (b) IO_4^-.

Solution We refer to Figure 2.14 and Table 2.6.

(a) We start with our reference acid, phosphoric acid (H_3PO_4). Because H_3PO_3 has one fewer O atom, it is called phosphorous acid.

(b) The parent acid is HIO_4. Because the acid has one more O atom than our reference iodic acid (HIO_3), it is called periodic acid. Therefore, the anion derived from HIO_4 is called periodate.

Practice Exercise Name the following oxoacid and oxoanion: (a) HBrO and (b) HSO_4^-.


Oxoacid
Removal of
all H$^+$ ions
Oxoanion
per– –ic acid
per– –ate
+[O]
Reference "–ic" acid
–ate
–[O]
"–ous" acid
–ite
–[O]
hypo– –ous acid
hypo– –ite

4. Naming Bases and Hydrates:

Bases: usually ends with hydroxide anion.

Hydrates: contain number of water molecules.

Naming:

Bases → First element name then hydroxide

" if the first element is TE name its charge"

Hydrates→ Name compound then prefix of number

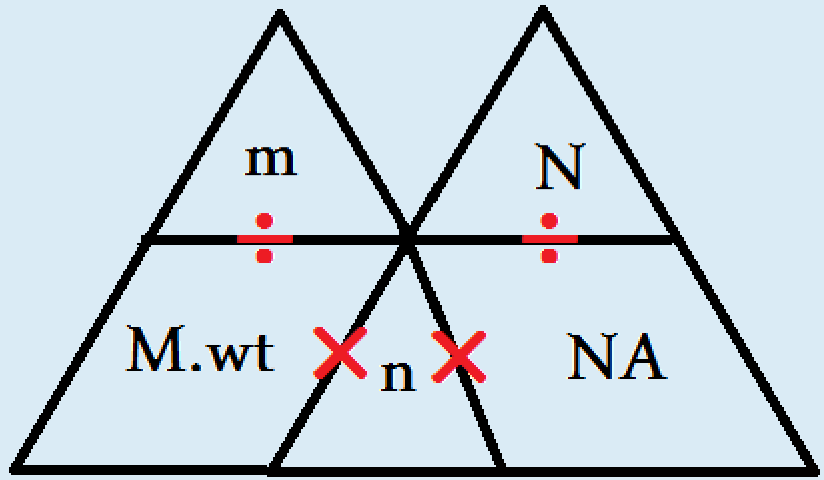
. of water molecules followed by hydrate.

Ex. $NaOH$, $Fe(OH)_2$ and $BaCl.2H_2O$

Answers to Practice Exercises

2.1 29 protons, 34 neutrons, and 29 electrons. **2.2** $CHCl_3$. **2.3** $C_4H_5N_2O$. **2.4** (a) Lead(II) oxide, (b) lithium sulfite. **2.5** (a) Rb_2SO_4, (b) BaH_2. **2.6** (a) Nitrogen trifluoride, (b) dichlorine heptoxide. **2.7** (a) SF_4, (b) N_2O_5. **2.8** (a) Hypobromous acid, (b) hydrogen sulfate ion.

Chapter 3: Stoichiometry



- Always use moles for proportionality (نسبة وتناسب).
- In the above figure:

m= weight or mass (gram) , N= Atoms number

n= moles number (moles) , NA= Avogadro's number

M.wt= Molecular/ Atomic weight (gram/ mole)

- check the book for examples! (pages 58- 94).

Chapter 4: Reactions in Aq Solutions

References:

General Chemistry : The Essential concepts by Chang, 5th edition.

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Thank you for listening and sorry for any unintended disruptions!