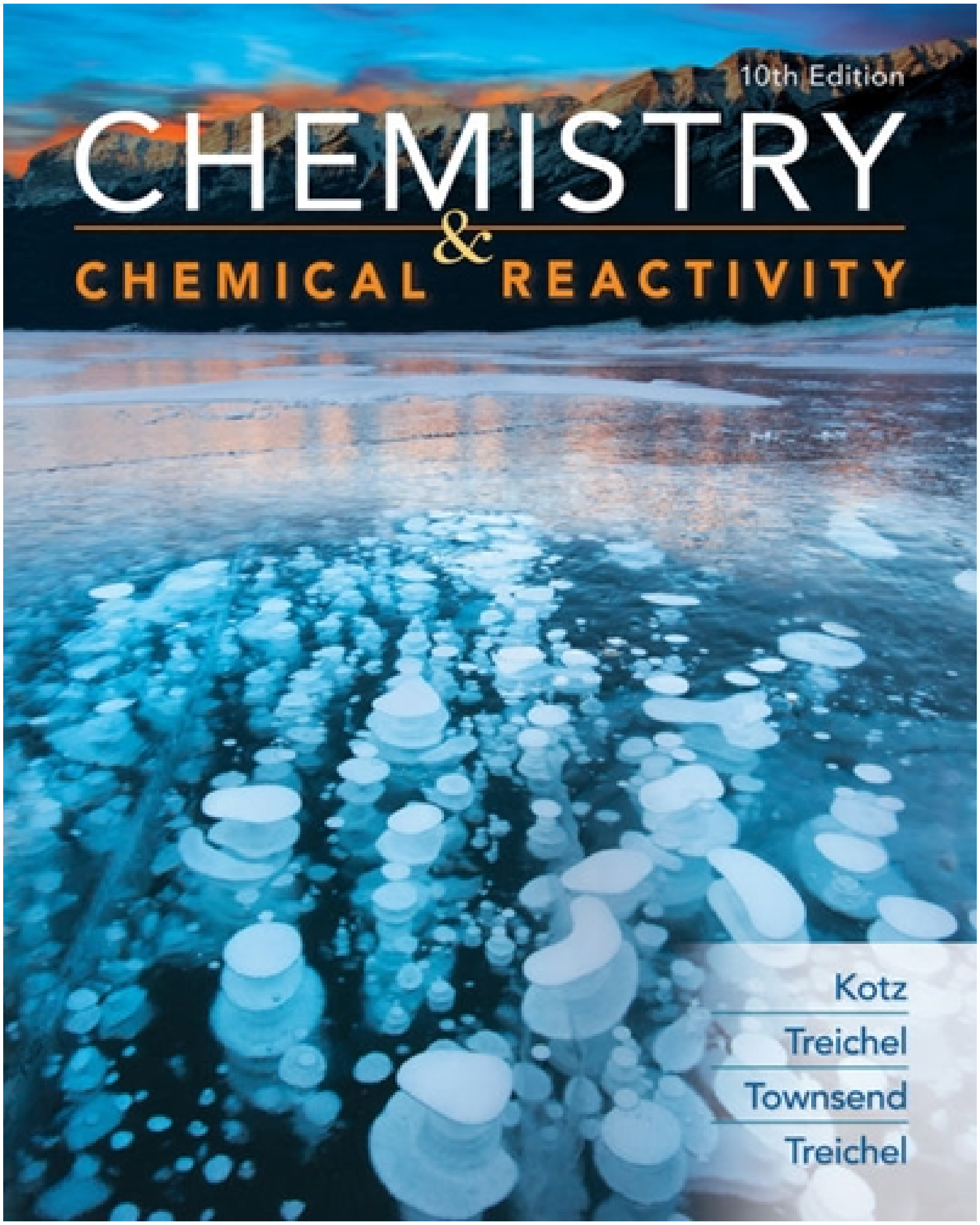
Solutions for Chemistry and Chemical Reactivity 10th Edition by Kotz



Solutions

Chapter 2
Atoms, Molecules and Ions

Applying Chemical Principles:
Using Isotopes

2.1.1. Neutrons in O-18; Neutrons in the two Pb isotopes:

Neutrons in O-18 : 18-8 = 10 neutrons

Neutrons in Pb-206: 206-82 = 124 neutrons

Neutrons in Pb-204: 204-82 = 122 neutrons

2.1.2. Using the abundance and mass of the three stable isotopes of O, calculate the atomic weight:

Atomic weight is a weighted average of the various isotopes. Add the product of (percent abundance)(isotopic mass) for all of the isotopes to obtain the atomic weight:

Mass	Abundance	Product
15.9949	0.99763	15.95699209
16.9991	0.000375	0.006374663
17.9991	0.001995	0.035908205
	SUM:	15.99927495

In this calculation, I have left more sf than is normally used. The sum (reported to 3sf) is 16.0, the atomic weight usually assigned to oxygen.

Arsenic, Medicine and the Formula of Compound 606

2.2.1. What is the empirical formula of the compound enargite?

The percentages total to 100%, so think of the percent of each element in 100. g of the compound: Mass As: 19.024 g; Mass Cu: 48.407 g; Mass S: 32.569 g.

Now convert the masses to moles of each element:

$$\frac{19.024\text{ g As}}{1} \cdot \frac{1\text{ mol As}}{74.9216\text{ g As}} = 0.25392\text{ mol As}$$

$$\frac{48.407\text{ g Cu}}{1} \cdot \frac{1\text{ mol Cu}}{63.546\text{ g Cu}} = 0.76176\text{ mol Cu}$$

$$\frac{32.569\text{ g S}}{1} \cdot \frac{1\text{ mol S}}{32.066\text{ g S}} = 1.01569\text{ mol S}$$

Calculate the mole: mole ratio:

$$\frac{1.01569 \text{ mol S}}{0.25392 \text{ mol As}} = 4.0000 \qquad \frac{0.76176 \text{ mol Cu}}{0.25392 \text{ mol As}} = 3.0000$$

The ratio of the three elements is: $As_1Cu_3S_4$ or usually written: $AsCu_3S_4$.

2.2.2. Molecular formulas of the compounds composing Salvarsan:

First determine empirical formulas for the substances:

$$\frac{39.37 \text{ g C}}{1} \cdot \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.2778 \text{ mol C} \qquad \frac{3.304 \text{ g H}}{1} \cdot \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 3.2781 \text{ mol H}$$

$$\frac{8.741 \text{ g O}}{1} \cdot \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.54633 \text{ mol O}$$

$$\frac{7.652 \text{ g N}}{1} \cdot \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 0.54631 \text{ mol N}$$

$$\frac{40.932 \text{ g As}}{1} \cdot \frac{1 \text{ mol As}}{74.9216 \text{ g As}} = 0.54633 \text{ mol As}$$

Determine mole:mole ratios:

Using the smallest # of mol (either for O, N, or As), determine ratios of C and H:

$$\frac{3.2778 \text{ mol C}}{0.54631 \text{ mol N}} = 5.99989 \text{ mol C/mol N or 6 mol C/mol N}$$

The ratio for H/N is obviously also 6 mol H/mol N.

These ratios imply an empirical formula that is C_6H_6NOAs.

Calculating the mass associated with this empirical formula gives: 183.04 g/empirical formula.

We are told that one of the compounds has a molar mass of 549 g/mole.

$$\frac{549 \text{ g / mol compound}}{183.04 \text{ g / empirical formula}} = 2.999 \text{ empirical formulas/molecular compound}$$

So that compound has a molecular formula of $C_{18}H_{18}N_3O_3As_3$.

The other compound has a molar mass of 915 g/mol. Using a similar procedure to that above, there are 5 empirical formulas/molecular compound, or a molecular formula of $C_{30}H_{30}N_5O_5As_5$.

Argon-An Amazing Discovery

2.3.1. Volume of globe:

$$\frac{0.20389 \text{ g N}}{1} \cdot \frac{1 \text{ L}}{1.25718 \text{ g N}} = 0.16218 \text{ L or } 162.18 \text{ mL or } 162.18 \text{ cm}^3$$

2.3.2. Determine the density of argon:

Using the prescribed procedure of summing the products of

(fractional volume)(density):

(0.2096)(1.42952 g/L) + (0.7811)(1.25092 g/L) + (0.00930)X = 1.000(1.29327 g/L)

and X = 1.78 g/L

2.3.3. From the periodic table, the atomic weight of Argon = 39.948 u. Now we deduce the abundance of Ar-40. 100 % – 0.337 % – 0.063 % = 99.600 %

Now that we know the abundance of Ar-40, we can determine the weighted average of the three isotopes: (0.00337)(35.967545 u) + (0.00063)(37.96732 u) + (0.99600)X = 39.948 u

and X = 39.963 u

2.3.4. Number of Ar atoms in a room 4.0 m × 5.0 m × 2.4 m:

Volume of room: $4.0 \text{ m} \cdot 5.0 \text{ m} \cdot 2.4 \text{ m} \cdot \frac{1 \text{ L}}{1.00 \times 10^{-3} \text{ m}^3} = 4.8 \times 10^4 \text{ L}$

Use this volume and the density of argon:

$$\frac{4.8 \times 10^4 \text{L}}{1} \cdot \frac{1.78 \text{ g Ar}}{1 \text{ L}} \cdot \frac{1 \text{ mol Ar}}{39.948 \text{ g Ar}} \cdot \frac{6.022 \times 10^{23} \text{atoms Ar}}{1 \text{ mol Ar}} =$$

1.3×10^{27} atoms Ar

PRACTICING SKILLS

Atoms: Their Composition and Structure

2.1.

Fundamental Particles	Protons	Electrons	Neutrons
Electrical Charges	+1	-1	0
Present in nucleus	Yes	No	Yes
Least Massive	1.007 u	**0.00055 u**	1.009 u

2.2. Mass number is the sum of the number of protons and number of neutrons for an atom. Atomic mass is the mass of an atom. Mass number is a simple TOTAL, while atomic mass conveys the sum of the weight of the individual particles (frequently expressed in units of u).

2.3. Begin by expressing the diameter of the nucleus and the electron cloud in the same units.

2mm (diameter of nucleus) = 2×10^{-3} m (since 1 m = 10^{3} mm, 1 mm = 10^{-3} m)

The ratio of diameters: $\dfrac{\text{Electron cloud}}{\text{nucleus}} = \dfrac{200\text{ m}}{2 \times 10^{-3}\text{ m}}$. So we set the actual diameters in the same ratio: $\dfrac{200\text{ m}}{2 \times 10^{-3}\text{m}} = \dfrac{1\times 10^{-8}\text{cm}}{\text{x}}$ and solving for x:

$200\text{ x} = (2 \times 10^{-3})(1 \times 10^{-8}) = 2 \times 10^{-11}$ and $\text{x} = 1 \times 10^{-13}$ cm.

Note that we left the actual diameter of the electron cloud in units of centimeters, so the **ratio** would be the same as if we had changed the units to **meters**.

2.4. Each gold atom has a diameter of 2×145 pm = 290. pm

$$\frac{36\text{ cm}}{1}\cdot\frac{1\text{ m}}{100\text{ cm}}\cdot\frac{1\times10^{12}\text{ pm}}{1\text{ m}}\cdot\frac{1\text{ Au atom}}{290.\text{ pm}} = 1.2\times10^{9}\text{ atoms Au}$$

2.5. Isotopic symbol for:

(a) Mg (at. no. 12) with 15 neutrons: 27 $^{27}_{12}Mg$

(b) Ti (at. no. 22) with 26 neutrons: 48 $^{48}_{22}Ti$

(c) Zn (at. no. 30) with 32 neutrons: 62 $^{62}_{30}Zn$

The mass number represents the SUM of the protons + neutrons in the nucleus of an atom. The atomic number represents the # of protons, so (atomic no. + # neutrons) =mass number

2.6. Isotopic symbol for:

(a) Ni (at. no. 28) with 31 neutrons: 59 $^{59}_{28}\text{Ni}$

(b) Pu (at. no. 94) with 150 neutrons: 244 $^{244}_{94}\text{Pu}$

(c) W (at. no. 74) with 110 neutrons: 184 $^{184}_{74}\text{W}$

2.7.

substance	protons	neutrons	electrons
(a) magnesium-24	12	12	12
(b) tin-119	50	69	50
(c) thorium-232	90	142	90

substance	protons	neutrons	electrons
(d) carbon-13	6	7	6
(e) copper-63	29	34	29
(f) bismuth-205	83	122	83

Note that the number of protons and electrons are **equal** for any **neutral atom**. The number of protons is **always** equal to the atomic number. The mass number equals the sum of the numbers of protons and neutrons.

2.8. (a) Number of protons = number of electrons = 43; number of neutrons = 56

(b) Number of protons = number of electrons = 95; number of neutrons = 146

Key Experiments Developing Atomic Structure

2.9. The accepted mass of a proton is 1.672622×10^{-24} g, while that for the electron is 9.109383×10^{-28} g. The ratio of these two masses is: $\frac{\text{mass of electron}}{\text{mass of proton}} = \frac{9.109383\times 10^{-28}\text{g}}{1.672622\times 10^{-24}\text{g}}$.

If one calculates the reciprocal of this 0.00054 value, one obtains about 1800:1, so Thomson's estimate was "off" by a factor of 2.

2.10. Negatively charged electrons in the cathode-ray tube collide with He atoms, splitting the atom into an electron and a He^{+} cation.

2.11. One can observe from the accompanying graphic, that the β particles are attracted to the "+" plate, while the α particles are attracted to the "-" one. The gamma rays are attracted to neither.

From this information, we know that α particles are **positively charged**, and that β particles are **negatively charged**. With the α particle being essentially the nucleus of a He atom, it is the heavier of the two particles.

2.12. Atoms are not solid, hard, or impenetrable. We know, thanks to the Kinetic-Molecular Theory, that the atoms are in rapid motion at all temperatures above absolute zero. They have mass (an important aspect of Dalton's hypothesis). The Rutherford "gold foil" experiment proved that atoms were *not* impenetrable, solid or hard, as alpha particles penetrated the atom and revealed a better picture of the atom.

Isotopes

2.13. The mass of a ^{16}O atom is 15.995 u. The mass relative to the mass of an atom of ^{12}C, which has a mass of 12.000 u will be 15.995/12.000 or 1.3329 (5 significant figures or 5 sf).

2.14. The mass of a ^{16}O atom is 15.995 u. So the mass in grams is:

$15.995 \text{ u} \cdot 1.661 \times 10^{-24} \text{ g/u} = 2.657 \times 10^{-23} \text{ g}$

2.15. Isotopes of cobalt (atomic number 27) with 30, 31, and 33 neutrons would have symbols of

$^{57}_{27}Co$, $^{58}_{27}Co$, and $^{60}_{27}Co$ respectively.

2.16. Atomic number of Ag is 47; both isotopes have 47 protons and 47 electrons.

^{107}Ag 107 – 47 = 60 neutrons and ^{109}Ag 109 – 47 = 62 neutrons

2.17. Hydrogen has three isotopes:

Name	# Protons	# Neutrons	# Electrons
Protium	1	0	1
Deuterium	1	1	1
Tritium	1	2	1

The ONLY difference between the isotopes of an element is in the **number** of neutrons.

2.18. Recalling that isotopes of an element must all have the same number of protons, the following are isotopes of element X: $^{19}_{9}X$, $^{20}_{9}X$, and $^{21}_{9}X$

Isotope Abundance and Atomic Weight

2.19. Thallium has two stable isotopes ^{203}Tl and ^{205}Tl. The more abundant isotope is:___?___

The atomic weight of thallium is 204.4 u. The fact that this weight is closer to 205 than 203 indicates that the **205 isotope is the more abundant** isotope. Recall that the atomic weight is the "weighted average" of all the isotopes of each element. Hence the more abundant isotope will have a "greater contribution" to the atomic weight than the less abundant one.

2.20. Strontium has an atomic weight of 87.62 so ^{88}Sr is the most abundant.

2.21. The atomic mass of lithium is: (0.0750)(6.015121) + (0.9250)(7.016003) = 6.94 u

Recall that the atomic mass is a weighted average of all isotopes of an element, and is obtained by **adding** the *product* of (relative abundance x mass) for all isotopes.

2.22. (^{24}Mg mass)(% abundance) + (^{25}Mg mass)(% abundance) + (^{26}Mg mass)(% abundance) = atomic weight of Mg

(23.985 u)(0.7899) + (24.986 u)(0.1000) + (25.983 u)(0.1101) = 24.31 u

2.23. The average atomic weight of gallium is 69.723 (from the periodic table). If we let **x** represent the abundance of the lighter isotope, and (**1-x**) the abundance of the heavier isotope, the expression to calculate the atomic weight of gallium may be written:

(**x**)(68.9257) + (**1 - x**)(70.9249) = 69.723

[Note that the sum of all the isotopic abundances must add to 100% -- or 1 (in decimal notation).]

Simplifying the equation gives:

$$68.9257 \text{ u } \mathbf{x} + 70.9249 \text{ u} - 70.9249 \text{ u } \mathbf{x} = 69.723 \text{ u}$$
$$-1.9992 \text{ u } \mathbf{x} = (69.723 \text{ u} - 70.9249)$$
$$-1.9992 \text{ u } \mathbf{x} = -1.202 \text{ u}$$
$$\mathbf{x} = 0.6012$$

So the relative abundance of isotope 69 is 60.12 % and that of isotope 71 is 39.88 %.

2.24. Let x represent the abundance of ^{151}Eu and (1 – x) represent the abundance of ^{153}Eu.

(x)(150.9197 u) + (1 – x)(152.9212 u) = 151.965 u, and solving for x yields

x = 0.4777; ^{151}Eu abundance is 47.77%, ^{153}Eu abundance is 52.23%

The Periodic Table

2.25. Comparison of Titanium and Thallium:

Name	Symbol	Atomic #	Atomic Weight	Group #	Period #	Metal, Metalloid, or nonmetal
Titanium	Ti	22	47.867	4B (4)	4	Metal
Thallium	Tl	81	204.3833	3A (13)	6	Metal

2.26. Elements with symbols beginning with S:

	silicon	tin	antimony	sulfur	selenium
Symbol	Si	Sn	Sb	S	Se
At. number	14	50	51	16	34
Period	3	5	5	3	4
Group	4A	4A	5A	6A	6A
	metalloid	metal	metalloid	nonmetal	nonmetal

2.27. Periods with 8 elements: **2**; Periods 2 (at.no. 3-10) and 3 (at.no. 11-18)

Periods with 18 elements: **2**; Periods 4 (at.no 19-36) and 5 (at.no. 37-54)

Periods with 32 elements: **2**; Periods 6 (at.no. 55-86) and 7 (at.no. 87-118)

2.28. There are 32 elements in the seventh period, the majority of them are called the Actinides, and many of them are man-made elements.

2.29. Elements fitting the following descriptions:

	Description	Elements
(a)	Nonmetals	C, Cl
(b)	Main group elements	C, Ca, Cl, Cs
(c)	Lanthanides	Ce
(d)	Transition elements	Cr, Co, Cd, Cu, Ce, Cf, Cm
(e)	Actinides	Cf, Cm
(f)	Gases	Cl

2.30. Name and chemical symbol for:

(a) C, carbon; N, nitrogen; O, oxygen; F, fluorine; Ne, neon

(b) Rb, rubidium

(c) Cl, chlorine

(d) H, hydrogen; He, helium; N, nitrogen; O, oxygen; F, fluorine; Ne, neon; Cl, chlorine; Ar, argon; Kr, krypton; Xe, xenon; Rn, radon

2.31. Classify the elements as metals, metalloids, or nonmetals:

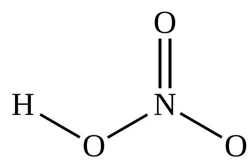
	Metals	Metalloids	Nonmetals
N			X
Na	X		
Ni	X		
Ne			X
Np	X		

2.32. Match symbol with description:

(a) A radioactive element Bk
(b) A liquid at room temperature Br
(c) A metalloid B
(d) An alkaline earth metal Ba
(e) A group 5A element Bi

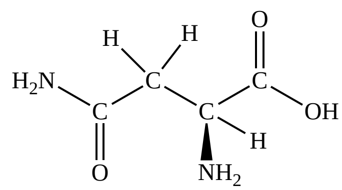
Molecular Formulas and Models

2.33. The molecular formula for nitric acid is HNO_3.



The structural formula is shown at right (without lone-pair electrons). The oxygens are in a trigonal planar arrangement, and the H-O-N atoms reside in bent arrangement. However all the atoms can reside in the same plane, so the molecule is planar.

2.34. Molecular formula for asparagine (an amino acid) is $C_4H_8N_2O_3$. The structural formula is:



Ions and Ion Charges

2.35. Most commonly observed monatomic ion for:

(a) Magnesium: 2+ —like all the alkaline earth metals
(b) Zinc: 2+
(c) Nickel: 2+
(d) Gallium: 3+ —an analog of Aluminum

2.36. Most commonly observed monatomic ion for:

(a) Selenium: 2-
(b) Fluorine: 1-
(c) Iron: 2+ and 3+ (transition metals frequently form multiple cations)
(d) Nitrogen: 3-

2.37. The symbol and charge for the following ions:

(a)	barium	Ba^{2+}		(e)	sulfide	S^{2-}
(b)	titanium(IV)	Ti^{4+}		(f)	perchlorate	ClO_4^-
(c)	phosphate	PO_4^{3-}		(g)	cobalt(II)	Co^{2+}
(d)	hydrogen carbonate	HCO_3^-		(h)	sulfate	SO_4^{2-}

2.38. Symbol and charge for the ions:

(a)	permanganate	MnO_4^-		(d)	ammonium	NH_4^+
(b)	nitrite	NO_2^-		(e)	phosphate	PO_4^{3-}
(c)	dihydrogen phosphate	$H_2PO_4^-$		(f)	sulfite	SO_3^{2-}

2.39. When potassium becomes a monatomic ion, potassium—like all alkali metals—**loses 1 electron.** The noble gas atom with the same number of electrons as the potassium ion is **argon**.

2.40. Both atoms gain two electrons. O^{2-} has the same number of electrons as Ne and S^{2-} has the same number of electrons as Ar.

Ionic Compounds

2.41. Barium is in Group 2A, and is expected to form a 2+ ion while bromine is in group 7A and expected to form a 1- ion. Since the compound would have to have an **equal amount** of negative and positive charges, the formula would be $BaBr_2$.

2.42. Cobalt is a transition metal and is can form a 3+ ion while fluorine is in group 7A and is expected to form a 1- ion. Since the compound would have to have an **equal amount** of negative and positive charges, the formula would be CoF_3.

2.43. Formula, Charge, and Number of ions in:

		cation	# of	anion	# of
(a)	K_2S	K^+	2	S^{2-}	1
(b)	$CoSO_4$	Co^{2+}	1	SO_4^{2-}	1
(c)	$KMnO_4$	K^+	1	MnO_4^-	1
(d)	$(NH_4)_3PO_4$	NH_4^+	3	PO_4^{3-}	1
(e)	$Ca(ClO)_2$	Ca^{2+}	1	ClO^-	2
(f)	$NaCH_3CO_2$	Na^+	1	$CH_3CO_2^-$	1

2.44. Formula, Charge, and Number of ions in:

		cation	# of	anion	# of
(a)	$Mg(CH_3CO_2)_2$	Mg^{2+}	1	$CH_3CO_2^-$	2
(b)	$Al(OH)_3$	Al^{3+}	1	OH^-	3
(c)	$CuCO_3$	Cu^{2+}	1	CO_3^{2-}	1
(d)	$Ti(SO_4)_2$	Ti^{4+}	1	SO_4^{2-}	2

		cation	# of	anion	# of
(e)	KH_2PO_4	K^+	1	$H_2PO_4^-$	1
(f)	$CaHPO_4$	Ca^{2+}	1	HPO_4^{2-}	1

2.45. Regarding cobalt oxides: Cobalt(II) oxide, CoO cobalt ion : Co^{2+}
Cobalt(III) oxide, Co_2O_3 Co^{3+}

2.46. Platinum (II and IV) ions with chloride and sulfide ions:
(a) Pt^{2+}: $PtCl_2$ Pt^{4+}: $PtCl_4$
(b) Pt^{2+}: PtS Pt^{4+}: PtS_2

2.47. Provide correct formulas for compounds:
(a) $AlCl_3$ The tripositive aluminum ion requires three chloride ions.
(b) KF Potassium is a monopositive cation. Fluoride is a mononegative anion.
(c) Ga_2O_3 is correct; Ga is a 3+ ion and O forms a 2- ion
(d) MgS is correct; Mg forms a 2+ ion and S forms a 2- ion

2.48. Provide correct formulas for compounds if given formula is incorrect:
(a) Ca_2O Incorrect; Correct is CaO.
(b) $SrBr_2$ Correct as is.
(c) Fe_2O_3 Correct for iron(III) oxide; could be FeO for iron(II) oxide
(d) Li_2O Correct as is.

Naming Ionic Compounds

2.49. Names for the ionic compounds
(a) K_2S potassium sulfide
(b) $CoSO_4$ cobalt(II) sulfate
(c) $(NH_4)_3PO_4$ ammonium phosphate
(d) $Ca(ClO)_2$ calcium hypochlorite

2.50. Name the following:
(a) $Ca(CH_3CO_2)_2$, calcium acetate
(b) $Ni_3(PO_4)_2$, nickel(II) phosphate
(c) $Al(OH)_3$, aluminum hydroxide
(d) KH_2PO_4, potassium dihydrogen phosphate

2.51. Formulas for the ionic compounds

(a) ammonium carbonate $(NH_4)_2CO_3$

(b) calcium iodide CaI_2

(c) copper(II) bromide $CuBr_2$

(d) aluminum phosphate $AlPO_4$

(e) silver(I) acetate $AgCH_3CO_2$

2.52. Formula for the following:

(a) calcium hydrogen carbonate $Ca(HCO_3)_2$

(b) potassium permanganate $KMnO_4$

(c) magnesium perchlorate $Mg(ClO_4)_2$

(d) potassium hydrogen phosphate K_2HPO_4

(e) sodium sulfite Na_2SO_3

2.53. Names and formulas for ionic compounds:

	anion	anion
cation	CO_3^{2-}	I^-
Na^+	Na_2CO_3 sodium carbonate	NaI sodium iodide
Ba^{2+}	$BaCO_3$ barium carbonate	BaI_2 barium iodide

2.54. Write formulas and name:

	anion	anion
cation	NO_3^-	PO_4^{3-}
Mg^{2+}	$Mg(NO_3)_2$ magnesium nitrate	$Mg_3(PO_4)_2$ magnesium phosphate
Fe^{3+}	$Fe(NO_3)_3$ iron(III) nitrate	$FePO_4$ iron(III) phosphate

Coulomb's Law

2.55. The fluoride ion has a smaller radius than the iodide ion. Hence the distance between the sodium and fluoride ions will be less than the comparable distance between sodium and iodide. Coulomb’s Law indicates that the attractive force becomes greater as the distance between the charges grows smaller—hence NaF will have stronger forces of attraction.

2.56. The attractive forces are stronger in CaO because the ion charges are greater (+2/–2 in CaO and +1/–1 in NaCl).

Naming Binary, Nonmetal Compounds

2.57. Names of binary nonionic compounds

(a) NF_3 nitrogen trifluoride

(b) HI hydrogen iodide

(c) BI_3 boron triiodide

(d) PF_5 phosphorus pentafluoride

2.58. Names of binary nonionic compounds

(a) N_2O_5 dinitrogen pentaoxide

(b) P_4S_3 tetraphosphorus trisulfide

(c) OF_2 oxygen difluoride

(d) XeF_4 xenon tetrafluoride

2.59. Formulas for:

(a) sulfur dichloride SCl_2

(b) dinitrogen pentaoxide N_2O_5

(c) silicon tetrachloride $SiCl_4$

(d) diboron trioxide B_2O_3

2.60. Formulas for:

(a) bromine trifluoride BrF_3

(b) xenon difluoride XeF_2

(c) hydrazine N_2H_4

(d) diphosphorus tetrafluoride P_2F_4

(e) butane C_4H_{10}

Atoms and the Mole

2.61. The mass, in grams of:

(a) $$\frac{2.5 \text{ mol Al}}{1} \cdot \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 67 \text{ g Al (2 sf)}$$

(b) $\frac{1.25 \text{ x } 10^{-3} \text{ mol Fe}}{1} \cdot \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 0.0698 \text{ g Fe (3 sf)}$

(c) $\frac{0.015 \text{ mol Ca}}{1} \cdot \frac{40.1 \text{ g Ca}}{1 \text{ mol Ca}} = 0.60 \text{ g Ca (2 sf)}$

(d) $\frac{653 \text{ mol Ne}}{1} \cdot \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 1.32 \text{ x } 10^{4} \text{ g Ne (3 sf)}$

Note that, whenever possible, one should use a molar mass of the substance that contains **one more** significant figure than the data, to reduce round-off error.

2.62. The mass (in g) of the following:

(a) $4.24 \text{ mol Au} \cdot \frac{197.0 \text{ g Au}}{1 \text{ mol Au}} = 835 \text{ g Au}$

(b) $15.6 \text{ mol He} \cdot \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 62.4 \text{ g He}$

(c) $0.063 \text{ mol Pt} \cdot \frac{195 \text{ g Pt}}{1 \text{ mol Pt}} = 12 \text{ g Pt}$

(d) $3.63 \times 10^{-4} \text{ mol Pu} \cdot \frac{244.7 \text{ g Pu}}{1 \text{ mol Pu}} = 0.0888 \text{ g Pu}$

2.63. The amount (moles) of substance represented by:

(a) $\frac{127.08 \text{ g Cu}}{1} \cdot \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} = 1.9998 \text{ mol Cu (5 sf)}$

(b) $\frac{0.012 \text{ g Li}}{1} \cdot \frac{1 \text{ mol Li}}{6.94 \text{ g Li}} = 1.7 \text{ x } 10^{-3} \text{ mol Li (2 sf)}$

(c) $\frac{5.0 \text{ mg Am}}{1} \cdot \frac{1 \text{ g Am}}{10^{3} \text{ mg Am}} \cdot \frac{1 \text{ mol Am}}{243 \text{ g Am}} = 2.1 \text{ x } 10^{-5} \text{ mol Am (2 sf)}$

(d) $\frac{6.75 \text{ g Al}}{1} \cdot \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.250 \text{ mol Al (3 sf)}$

2.64. The amount (moles) of substance represented by:

(a) $16.0 \text{ g Na} \cdot \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.696 \text{ mol Na}$

(b) $0.876 \text{ g Sn} \cdot \frac{1 \text{ mol Sn}}{118.7 \text{ g Sn}} = 7.38 \times 10^{-3} \text{ mol Sn}$

(c) $0.0034 \text{ g Pt} \cdot \frac{1 \text{ mol Pt}}{195 \text{ g Pt}} = 1.7 \times 10^{-5} \text{ mol Pt}$

(d) $0.983 \text{ g Xe} \cdot \frac{1 \text{ mol Xe}}{131.1 \text{ g Xe}} = 7.49 \times 10^{-3} \text{ mol Xe}$

2.65. 1-gram samples of He, Fe, Li, Si, C:

Which sample contains the **largest number** of atoms? …the **smallest number** of atoms?
If we calculate the number of atoms of any one of these elements, say He, the process is:

$$\frac{1.0 \text{ g He}}{1} \cdot \frac{1 \text{ mol He}}{4.0026 \text{ g He}} \cdot \frac{6.022 \times 10^{23} \text{ atoms He}}{1 \text{ mol He}} = 1.5 \times 10^{23} \text{ atoms He}$$

All the calculations proceed analogously, with the ONLY numerical difference attributable to the molar mass of the element. Therefore the element with the **smallest** molar mass (He) will have the **largest number** of atoms, while the element with the **largest** molar mass (Fe) will have the **smallest number** of atoms. This is a great question to answer by **thinking** rather than by calculating.

2.66. List samples in increasing amounts:

$$0.10 \text{ g K} \cdot \frac{1 \text{ mol K}}{39.0983 \text{ g K}} = 0.0026 \text{ mol K}$$

$$0.10 \text{ g Mo} \cdot \frac{1 \text{ mol Mo}}{95.96 \text{ g Mo}} = 0.0010 \text{ mol Mo}$$

$$0.10 \text{ g Cr} \cdot \frac{1 \text{ mol Cr}}{51.9961 \text{ g Cr}} = 0.0019 \text{ mol Cr}$$

$$0.10 \text{ g Al} \cdot \frac{1 \text{ mol Al}}{26.9815 \text{ g Al}} = 0.0037 \text{ mol Al}$$

0.0010 mol Mo < 0.0019 mol Cr < 0.0026 mol K < 0.0037 mol Al

2.67. Analysis of a 10.0-g sample of apatite contained 3.99g Ca, 1.85g P, 41.4g O, and 0.020g H.

Calculating moles of each element gives:

$$3.99 \text{ g Ca} \cdot \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = 0.0996 \text{ mol Ca} \text{ and } 1.85 \text{ g P} \cdot \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 0.0597 \text{ mol P}$$

$$4.14 \text{ g O} \cdot \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.259 \text{ mol O} \text{ and } 0.020 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.0 \text{ g H}} = 0.020 \text{ mol H}$$

From smallest to largest # moles: 0.020 mol H < 0.0597 mol P < 0.0996 mol Ca < 0.259 mol O

2.68. Semiconducting material contains 52 g Ga, 9.5 g Al, 112 g As. Element with largest number of atoms:

$$\frac{52 \text{ g Ga}}{1} \cdot \frac{1 \text{ mol Ga}}{69.7 \text{ g Ga}} \cdot \frac{6.022 \times 10^{23} \text{ atoms Ga}}{1 \text{ mol Ga}} = 4.5 \text{ x } 10^{23} \text{ atoms Ga}$$

$$\frac{9.5 \text{ g Al}}{1} \cdot \frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \cdot \frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}} = 2.1 \text{ x } 10^{23} \text{ atoms Al}$$

$$\frac{112\text{ g As}}{1}\cdot\frac{1\text{ mol As}}{74.92\text{ g As}}\cdot\frac{6.022\times10^{23}\text{ atoms As}}{1\text{ mol As}} = 9.00\times10^{23}\text{ atoms As}$$

Arsenic has the largest number of atoms in the mixture.

Molecules, Compounds, and the Mole

2.69. Molar mass of the following: (with atomic weights expressed to 4 significant figures)

(a) Fe_2O_3 (2)(55.85) + (3)(16.00) = 159.7 g/mol

(b) BCl_3 (1)(10.81) + (3)(35.45) = 117.2 g/mol

(c) $C_6H_8O_6$ (6)(12.01) + (8)(1.008) + (6)(16.00) = 176.1 g/mol

2.70. Molar mass of the following:

(a) $Fe(C_6H_{11}O_7)_2$ (1)(55.85) + (12)(12.01) + (22)(1.008) + (14)(16.00) = 446.14 g/mol

(b) $CH_3CH_2CH_2CH_2SH$ (4)(12.01) + (10)(1.008) + (1)(32.07) = 90.19 g/mol

(c) $C_{20}H_{24}N_2O_2$ (20)(12.01) + (24)(1.008) + (2)(14.01) + (2)(16.00) = 324.42 g/mol

2.71. Molar mass of the following: (with atomic weights expressed to 4 significant figures)

(a) $Ni(NO_3)_2\cdot 6H_2O$ (1)(58.69) + (2)(14.01) + 6(16.00) + (12)(1.008) + (6)(16.00) = 290.8 g/mol

(b) $CuSO_4\cdot 5H_2O$ (1)(63.55) + (1)(32.07) + 4(16.00) + (10)(1.008) + (5)(16.00) = 249.7 g/mol

2.72. Molar mass of the following:

(a) $H_2C_2O_4\cdot 2H_2O$ (6)(1.008) + (2)(12.01) + (6)(16.00) = 126.07 g/mol

(b) $MgSO_4\cdot 7H_2O$ (1)(24.31) + (1)(32.07) + (11)(16.00) + (14)(1.008) = 246.48 g/mol

2.73. Mass represented by 0.0255 moles of the following compounds:

Molar masses are calculated as before. To determine the mass represented by 0.0255 moles, recall that 1 mol of a substance has a mass equal to the molar mass expressed in units of grams.

We calculate for (a) the mass represented by 0.0255 moles:

$$\frac{0.0255\text{ mol }C_3H_7OH}{1}\cdot\frac{60.10\text{ g }C_3H_7OH}{1\text{ mol }C_3H_7OH} = 1.53\text{ g }C_3H_7OH$$

Compound	Molar mass	Mass of 0.0255 moles
(a) C_3H_7OH	60.10	1.53 g
(b) $C_{11}H_{16}O_2$	180.2	4.60 g
(c) $C_9H_8O_4$	180.2	4.60 g
(d) C_3H_6O	58.08	1.48 g

Masses are expressed to 3 sf, since the # of moles has 3.

2.74. Mass of 0.123 mol of each of the following:

(a) $0.123 \text{ mol } C_{14}H_{10}O_4 \cdot \dfrac{242.2 \text{ g } C_{14}H_{10}O_4}{1 \text{ mol } C_{14}H_{10}O_4} = 29.8 \text{ g } C_{14}H_{10}O_4$

(b) $0.123 \text{ mol } C_4H_8N_2O_2 \cdot \dfrac{116.2 \text{ g } C_4H_8N_2O_2}{1 \text{ mol } C_4H_8N_2O_2} = 14.3 \text{ g } C_4H_8N_2O_2$

(c) $0.123 \text{ mol } C_5H_{10}S \cdot \dfrac{102.2 \text{ g } C_5H_{10}S}{1 \text{ mol } C_5H_{10}S} = 12.6 \text{ g } C_5H_{10}S$

(d) $0.123 \text{ mol } C_{12}H_{17}NO \cdot \dfrac{191.3 \text{ g } C_{12}H_{17}NO}{1 \text{ mol } C_{12}H_{17}NO} = 23.5 \text{ g } C_{12}H_{17}NO$

2.75. Regarding sulfur trioxide:

1. Amount of SO_3 in 1.00 kg: $\dfrac{1.00 \text{ x } 10^3 \text{ g } SO_3}{1} \cdot \dfrac{1 \text{ mol } SO_3}{80.07 \text{ g } SO_3} = 12.5 \text{ mol } SO_3$

2. Number of SO_3 molecules: $\dfrac{12.5 \text{ mol } SO_3}{1} \cdot \dfrac{6.022 \times 10^{23} \text{ } SO_3}{1 \text{ mol } SO_3} = 7.52 \text{ x } 10^{24} \text{ } SO_3 \text{ molecules}$

3. Number of S atoms: With 1 S atom per SO_3 molecule—7.52×10^{24} S atoms

4. Number of O atoms: With 3 O atoms per SO_3 molecule—$3 \times 7.52 \times 10^{24}$ O atoms

or 2.26×10^{25} O atoms

2.76. Number of ammonium and sulfate ions present:

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{2 \text{ mol } NH_4^+}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ } NH_4^+ \text{ ions}}{1 \text{ mol } NH_4^+} = 2.4 \times 10^{23} \text{ } NH_4^+ \text{ ions}$

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ } SO_4^{2-} \text{ ions}}{1 \text{ mol } SO_4^{2-}} = 1.2 \times 10^{23} \text{ } SO_4^{2-} \text{ ions}$

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{2 \text{ mol N}}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N}} = 2.4 \times 10^{23} \text{ N atoms}$

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{8 \text{ mol H}}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} = 9.6 \times 10^{23} \text{ H atoms}$

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{1 \text{ mol S}}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ S atoms}}{1 \text{ mol S}} = 1.2 \times 10^{23} \text{ S atoms}$

$0.20 \text{ mol } (NH_4)_2SO_4 \cdot \dfrac{4 \text{ mol O}}{1 \text{ mol } (NH_4)_2SO_4} \cdot \dfrac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O}} = 4.8 \times 10^{23} \text{ O atoms}$

2.77. $C_8H_9NO_2$, the molecular formula for acetominophen has a molecular mass of 151.17 g/mol.

Two 500 mg tablets [or $2 \times (500 \times 10^{-3})$ g = 1.00g] would contain:

$$\frac{1.00\text{ g}}{1} \cdot \frac{1\text{ mol acetominophen}}{151.17\text{ g acetominophen}} \cdot \frac{6.022 \times 10^{23}\text{ molecules acetominophen}}{1\text{ mol acetominophen}} = 3.98 \times 10^{21}\text{ molecules}$$

$$= 4 \times 10^{21}\text{ molecules (1 sf)}$$

2.78. In an Alka-Seltzer tablet:

(a) $$324\text{ mg } C_9H_8O_4 \cdot \frac{1\text{ g}}{10^3\text{ mg}} \cdot \frac{1\text{ mol } C_9H_8O_4}{180.2\text{ g } C_9H_8O_4} = 1.80 \times 10^{-3}\text{ mol } C_9H_8O_4$$

$$1904\text{ mg NaHCO}_3 \cdot \frac{1\text{ g}}{10^3\text{ mg}} \cdot \frac{1\text{ mol NaHCO}_3}{84.007\text{ g NaHCO}_3} = 0.02266\text{ mol NaHCO}_3$$

$$1000.\text{ mg } C_6H_8O_7 \cdot \frac{1\text{ g}}{10^3\text{ mg}} \cdot \frac{1\text{ mol } C_6H_8O_7}{192.13\text{ g } C_6H_8O_7} = 5.205 \times 10^{-3}\text{ mol } C_6H_8O_7$$

(b) $$1.80 \times 10^{-3}\text{ mol } C_9H_8O_4 \cdot \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol } C_9H_8O_4} = 1.08 \times 10^{21}\text{ molecules } C_9H_8O_4$$

Percent Composition

2.79. Mass percent for: [4 significant figures]

(a) PbS: (1)(207.2) + (1)(32.06) = 239.3 g/mol

$$\%Pb = \frac{207.2\text{ g Pb}}{239.3\text{ g PbS}} \times 100 = 86.60\ \%$$

%S = 100.00 - 86.60 = 13.40 %

(b) C_3H_8: (3)(12.01) + (8)(l.008) = 44.09 g/mol

$$\%C = \frac{36.03\text{ g C}}{44.09\text{ g } C_3H_8} \times 100 = 81.71\ \%$$

%H = 100.00 - 81.71 = 18.29 %

(c) $C_{10}H_{14}O$: (10)(12.01) + (14)(1.008) + (1)(16.00) = 150.21 g/mol

$$\%C = \frac{120.1\text{ g C}}{150.21\text{ g } C_{10}H_{14}O} \times 100 = 79.96\ \%$$

$$\%H = \frac{14.112\text{ g H}}{150.21\text{ g } C_{10}H_{14}O} \times 100 = 9.394\ \%$$

%O = 100.00 - (79.96 + 9.394) = 10.65 %

2.80. Mass percent of elements in the compounds:

(a) $\frac{(8)(12.01)\text{ g C}}{166.18\text{ g }C_8H_{10}N_2O_2} \cdot 100\% = 57.82\%\text{ C}$ $\frac{(10)(1.008)\text{ g H}}{166.18\text{ g }C_8H_{10}N_2O_2} \cdot 100\% = 6.066\%\text{ H}$

$\frac{(2)(14.01)\text{ g N}}{166.18\text{ g }C_8H_{10}N_2O_2} \cdot 100\% = 16.86\%\text{ N}$ $\frac{(2)(16.00)\text{ g O}}{166.18\text{ g }C_8H_{10}N_2O_2} \cdot 100\% = 19.26\%\text{ O}$

(b) $\frac{(10)(12.01)\text{ g C}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 76.86\%\text{ C}$ $\frac{(20)(1.008)\text{ g H}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 12.90\%\text{ H}$

$\frac{16.00\text{ g O}}{156.26\text{ g }C_{10}H_{20}O} \cdot 100\% = 10.24\%\text{ O}$

(c) $\frac{58.93\text{ g Co}}{237.93\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 24.77\%\text{ Co}$ $\frac{(2)(35.45)\text{ g Cl}}{237.93\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 29.80\%\text{ Cl}$

$\frac{(12)(1.008)\text{ g H}}{237.93\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 5.084\%\text{ H}$ $\frac{(6)(16.00)\text{ g O}}{237.93\text{ g }CoCl_2 \cdot 6\,H_2O} \cdot 100\% = 40.35\%\text{ O}$

2.81. Mass of CuS to provide 10.0 g of Cu:

To calculate the weight percent of Cu in CuS, we need the respective atomic weights:

Cu = 63.546; S = 32.066 adding CuS gives a molecular weight of 95.612

The % of Cu in CuS is then: $\frac{63.546\text{ g Cu}}{95.612\text{ g CuS}} \times 100 = 66.46\ \%\text{ Cu}$

Now with this fraction (inverted) calculate the mass of CuS that will provide 10.0 g of Cu:

$$\frac{10.0\text{ g Cu}}{1} \cdot \frac{95.612\text{ g CuS}}{63.546\text{ g Cu}} = 15.0\text{ g CuS}$$

2.82. Mass percent of Ti in ilmenite; Mass of ilmenite to obtain 750 g of Ti:

$$\frac{47.87\text{ g Ti}}{151.71\text{ g }FeTiO_3} \cdot 100\% = 31.55\%\text{ Ti}$$

$$750\text{ g Ti} \cdot \frac{100.00\text{ g }FeTiO_3}{31.55\text{ g Ti}} = 2.4 \times 10^3\text{ g }FeTiO_3$$

Empirical and Molecular Formulas

2.83. The empirical formula ($C_2H_3O_2$) would have a mass of 59.04 g.

Since the molar mass is 118.1 g/mol we can write

$$\frac{1\text{ empirical formula}}{59.04\text{ g succinic acid}} \cdot \frac{118.1\text{ g succinic acid}}{1\text{ mol succinic acid}} = \frac{2.0\text{ empirical formulas}}{1\text{ mol succinic acid}}$$

So the molecular formula contains 2 empirical formulas (2 x $C_2H_3O_2$) or $C_4H_6O_4$.

2.84. Molecular formula for a compound:

Empirical formula mass = 58.06 g/mol; $\frac{116.1 \text{ g/molecular formula}}{58.06 \text{ g/empirical formula}}$ = 2 empirical formulas/molecular formula; The molecular formula is $C_4H_8N_2O_2$.

2.85. Provide the empirical or molecular formula for the following, as requested:

	Empirical Formula	Molar Mass (g/mol)	Molecular Formula
(a)	CH	26.0	**C_2H_2**
(b)	CHO	116.1	**$C_4H_4O_4$**
(c)	**CH_2**	**112.2**	C_8H_{16}

Note that we can calculate the mass of an empirical formula by adding the respective atomic weights (13 for CH, for example). The molar mass (26.0 for part (a)) is obviously twice that for an empirical formula, so the molecular formula would be 2 x empirical formula (or **C_2H_2** in part (a)).

2.86. Provide the empirical or molecular formula for the following, as requested:

	Empirical Formula	Molar Mass (g/mol)	Molecular Formula
(a)	$C_2H_3O_3$	150.1	**$C_4H_6O_6$**
(b)	C_3H_8	44.1	**C_3H_8**
(c)	**B_2H_5**	**53.3**	B_4H_{10}

2.87. Calculate the empirical formula of acetylene by calculating the atomic ratios of carbon and hydrogen in 100 g of the compound.

$92.26 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.011 \text{g C}} = 7.681 \text{ mol C}$ and $7.74 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.678 \text{ mol H}$

Calculate the atomic ratio: $\frac{7.68 \text{ mol C}}{7.68 \text{ mol H}} = \frac{1 \text{ mol C}}{1 \text{ mol H}}$

The atomic ratio indicates that there is 1 C atom for 1 H atom (1:1). The **empirical formula is then CH**. The formula mass is 13.01. Given that the molar mass of the compound is 26.02 g/mol, there are two formula units per molecular unit, hence the **molecular formula for acetylene is C_2H_2**.

2.88. Empirical formula of a B-H compound:

The compound is 88.5% B and 11.5% H. Assume 100.0 g of compound.

$$88.5 \text{ g B} \cdot \frac{1 \text{ mol B}}{10.81 \text{ g B}} = 8.19 \text{ mol B} \qquad 11.5 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 11.4 \text{ mol H}$$

$$\frac{11.4 \text{ mol H}}{8.19 \text{ mol B}} = \frac{1.39 \text{ mol H}}{1 \text{ mol B}} = \frac{7/5 \text{ mol H}}{1 \text{ mol B}} = \frac{7 \text{ mol H}}{5 \text{ mol B}}$$ So the empirical formula is B_5H_7.

2.89. Determine the empirical and molecular formulas of cumene:

The percentage composition of cumene is 89.94% C and (100.00-89.94) or 10.06%H.

Calculate the ratio of mol C: mol H as done in SQ87.

$$89.94 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.011 \text{g C}} = 7.489 \text{ mol C}$$

$$10.06 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.981 \text{ mol H}$$

Calculating the atomic ratio:

$$\frac{9.981 \text{ mol H}}{7.489 \text{ mol C}} = \frac{1.33 \text{ mol H}}{1.00 \text{ mol C}}$$ or a ratio of 3C : 4H

So the empirical formula for cumene is C_3H_4 , with a formula mass of 40.06.
If the molar mass is 120.2 g/mol, then dividing the "empirical formula mass" into the molar mass gives: 120.2/40.06 or 3 empirical formulas **per** molar mass. The **molecular formula** is then 3 x C_3H_4 or C_9H_{12}.

2.90. Empirical and Molecular formula for sulflower:

$$57.17 \text{ g S} \cdot \frac{1\text{mol S}}{32.065 \text{ g S}} = 1.783 \text{ mol S} \qquad 42.83 \text{ g C} \cdot \frac{1\text{mol C}}{12.011 \text{ g C}} = 3.566 \text{ mol C}$$

$$\frac{3.566 \text{ mol C}}{1.783 \text{ mol S}} = \frac{2 \text{ mol C}}{1 \text{ mol S}}$$ So the empirical formula is C_2S. The molar mass of a C_2S entity is 56.087 g/mol, so divide that into the molar mass:

$$\frac{448.70 \text{ g/mol}}{56.087 \text{ g/mol}} = 8.$$ The molecular formula is $C_{16}S_8$

2.91. Empirical and Molecular formula for Mandelic Acid:

$$63.15 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 5.258 \text{ mol C} \qquad 5.30 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 5.26 \text{ mol H}$$

$$31.55 \text{ g O} \cdot \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.972 \text{ mol O}$$

Using the smallest number of atoms, we calculate the ratio of atoms:

$$\frac{5.258 \text{ mol C}}{1.972 \text{ mol O}} = \frac{2.666 \text{ mol C}}{1 \text{ mol O}} \text{ or } \frac{22/3 \text{ mol C}}{1\text{mol O}} \text{ or } \frac{8/3 \text{ mol C}}{1 \text{ mol O}}$$

So 3 mol O combine with 8 mol C and 8 mol H and the empirical formula is $C_8H_8O_3$.

The formula mass of $C_8H_8O_3$ is 152.15. Given the data that the molar mass is 152.15 g/mL, the molecular formula for mandelic acid is $C_8H_8O_3$.

2.92. Empirical and Molecular formula for Nicotine:

Assume 100.0 g of compound.

$$74.0 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.16 \text{ mol C} \qquad 8.65 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 8.58 \text{ mol H}$$

$17.35 \text{ g N} \cdot \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 1.239$ mol N. Now determine the mole:mole ratio:

$\frac{6.16 \text{ mol C}}{1.239 \text{ mol N}} = \frac{5 \text{ mol C}}{1 \text{ mol N}} \quad \frac{8.58 \text{ mol H}}{1.239 \text{ mol N}} = \frac{7 \text{ mol H}}{1 \text{ mol N}}$ and the empirical formula is C_5H_7N.

Compare the mass of the empirical formula to that of the molecular formula: $\frac{162 \text{ g/mol}}{81.1 \text{ g/mol}} = 2$

so the molecular formula is $C_{10}H_{14}N_2$

Determining Formulas from Mass Data

2.93. Given the masses of xenon involved, we can calculate the number of moles of the element:

$$0.526 \text{ g Xe} \cdot \frac{1\text{mol Xe}}{131.29 \text{ g Xe}} = 0.00401 \text{ mol Xe}$$

The mass of fluorine present is: 0.678 g compound – 0.526 g Xe = 0.152 g F

$$0.152 \text{ g F} \cdot \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 0.00800 \text{ mol F}$$

Calculating atomic ratios:

$\frac{0.00800 \text{ mol F}}{0.00401 \text{ mol Xe}} = \frac{2 \text{ mol F}}{1 \text{ mol Xe}}$ indicating that the empirical formula is XeF_2.

2.94. Value of x in the formula SF_x:

5.722 g compound – 1.256 g S = 4.466 g F; Determine mol of S and mol of F:

$$1.256 \text{ g S} \cdot \frac{1 \text{ mol S}}{32.066 \text{ g S}} = 0.03917 \text{ mol S} \qquad 4.466 \text{ g F} \cdot \frac{1 \text{ mol F}}{18.998 \text{ g F}} = 0.2351 \text{ mol F}$$

$\frac{0.2351 \text{ mol F}}{0.03917 \text{ mol S}} = \frac{6 \text{ mol F}}{1 \text{ mol S}}$, so the empirical formula is SF_6; $x = 6$

2.95. Knowing the mass of the heptahydrate, we can calculate the # of moles in the original

sample: $\frac{1.394 \text{ g heptahydrate}}{1} \cdot \frac{1 \text{ mol heptahydrate}}{246.47 \text{ g heptahydrate}} = 0.005656$ mol heptahydrate

Knowing that any weight loss is attributable to water, we can calculate the # of moles of water

lost due to the heating: $\frac{(1.394\text{g} - 0.885 \text{ g}) \text{ water lost}}{1} \cdot \frac{1 \text{ mol water}}{18.015 \text{ g water}} = 0.02825$ mol water

Calculate the RATIO of mol water lost: mol heptahydrate (initially present)

$\frac{0.02825 \text{ mol water lost}}{0.005656 \text{ mol heptahydrate}} = 4.995$ mol water lost/mol heptahydrate

(or 5 mol water lost/mol heptahydrate). Since the compound BEGAN as a heptahydrate, and LOST 5 waters of hydration/mol, the value of x is (7-5) or 2.

2.96. Formula for Ge_xCl_y:

3.69 g product – 1.25 g Ge = 2.44 g Cl; Determine # mol of Cl and Ge:

$$1.25 \text{ g Ge} \cdot \frac{1 \text{ mol Ge}}{72.61 \text{ g Ge}} = 0.0172 \text{ mol Ge} \qquad 2.44 \text{ g Cl} \cdot \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.0688 \text{ mol Cl}$$

Calculate the mol: mol ratio: $\frac{0.0688 \text{ mol Cl}}{0.0172 \text{ mol Ge}} = \frac{4 \text{ mol Cl}}{1 \text{ mol Ge}}$, so the empirical formula is $GeCl_4$.

Mass Spectrometry

2.97. Regarding the mass spectrum for NO_2:

(a) The four cations are: 14(N^+); 16(O^+); 30(NO^+); 46(NO_2^+). Note the assignments are based upon the masses of the atoms involved.

(b) The mass spectrum does provide evidence for NO bonds as opposed to OO bonds. An OO bond configuration would have produced a peak at 32 m/Z.

2.98. Regarding the mass spectrum for POF_3:

(a) Fragment at m/Z ratio of 85: The peak at 104 is the molecular ion, POF_3^+ , so the fragment at 85 represents (104-19) or POF_2^+

(b) Fragment at m/Z ratio of 69: This fragment is PF_2^+, indicating a loss of O (85-16).

(c) Fragment at m/Z ratio of 47 indicates a PO^+ fragment, and the fragment at m/Z ratio of 50 indicates a PF^+ fragment. The lack of a peak at 35, indicates an absence of OF^+.

2.99. Regarding the mass spectrum of CH_3Cl:

(a) Species responsible for lines at 50 and 52:

$^{12}CH_3{}^{35}Cl$ gives rise to line at 50 (12 + 3 + 35)

$^{12}CH_3{}^{37}Cl$ gives rise to line at 52 (12 + 3 + 37)

The differential in sizes of the lines is due to the relative distribution of isotopes of Cl, with Cl-37 being about 1/3 of the abundance of Cl-35.

(b) The line at 51 is due to C-13 in place of C-12, with the accompanying smaller size of the peak.

2.100. Consider the peaks in the Br_2 mass spectrum at 158, 160, and 162:

(a) The m/Z peak at 158 is: $^{79}Br_2$; while the peak at *m/Z* 160 is: $^{79}Br^{81}Br$. Finally the peak at *m/Z* 162 represents $^{81}Br_2$.

(b) The abundances are close enough to assume an equal abundance of ^{79}Br and ^{81}Br. Two atoms from the two isotopes can be combined in four different manners to form Br_2: $^{79}Br_2$, $^{79}Br^{81}Br$, $^{81}Br^{79}Br$, and $^{81}Br_2$. Thus, the peak at *m/Z* 160 should have twice the intensity of the peaks at *m/Z* 158 and 162.

GENERAL QUESTIONS

2.101.

Symbol	^{58}Ni	^{33}S	^{20}Ne	^{55}Mn
Number of protons	28	16	10	25
Number of neutrons	30	17	10	30
Number of electrons in the neutral atom	28	16	10	25
Name of element	nickel	sulfur	neon	manganese

2.102. As the atomic weight of potassium is 39.0983 u (remember this is a weighted average), the lighter isotope, ^{39}K is more abundant than ^{41}K.

2.103. Crossword puzzle: Clues:

Horizontal
1-2 A metal used in ancient times: tin (Sn)

3-4 A metal that burns in air and is found in Group 5A: bismuth (Bi)

Vertical
1-3 A metalloid: antimony (Sb)

2-4 A metal used in U.S. coins: nickel (Ni)

Using these solutions, the following letters fit

1 **S**	2 **N**
3 **B**	4 **I**

Single squares:
1. A colorful nonmetal: sulfur (S)

2. A colorless gaseous nonmetal: nitrogen (N)

3. An element that makes fireworks green: boron (B)

4. An element that has medicinal uses: iodine (I)

Diagonal:

1-4 An element used in electronics: silicon (Si)

2-3 A metal used with Zr to make wires for superconducting magnets: niobium (Nb)

2.104. Regarding the abundance of elements:

(a) Mg is the most abundant main group metal.

(b) H is the most abundant nonmetal.

(c) Si is the most abundant metalloid.

(d) Fe is the most abundant transition element.

(e) F and Cl are the halogens included, and of these Cl is the most abundant.

2.105. Copper atoms:

(a) The average mass of one copper atom:

One mole of copper (with a mass of 63.546 g) contains 6.0221×10^{23} atoms. So the average mass of **one** copper atom is: $\dfrac{63.546 \text{ g Cu}}{6.0221\text{x}10^{23}\text{atoms Cu}} = 1.0552 \times 10^{-22}\text{g/Cu atom}$

(b) Given the cost data: $41.70 for 7.0 g and the mass of a Cu atom (from part (a)), the cost of one Cu atom is: $\dfrac{\$41.70}{7.0 \text{ g Cu}} \cdot \dfrac{1.0552 \times 10^{-22} \text{ g Cu}}{1 \text{ Cu atom}} = \dfrac{6.286 \times 10^{-22} \text{ dollars}}{\text{Cu atom}}$

or 6.3×10^{-22} dollars/Cu atom (2 sf)

2.106. Identify which of the following is possible:

(a) silver foil that is 1.2×10^{-4} m thick (thick enough for an atom of Ag—possible)

(b) a sample of potassium that contains 1.784×10^{24} atoms (about 3 moles of K—possible)

(c) a gold coin of mass 1.23×10^{-3} kg (massive enough for more than one Au atom—possible)

(d) 3.43×10^{-27} mol S_8 (amount is less than one molecule of S_8—impossible)

2.107. Identify the element that:

(a) Is in Group 2A and the 5th period: Strontium

(b) Is in the 5th period and Group 4B: Zirconium

(c) Is in the second period in Group 4A: Carbon

(d) Is an element in the 4th period of Group 5A: Arsenic

(e) Is a halogen (Group 7A) in the 5th period: Iodine

(f) Is an alkaline earth element (Group 2A) in the 3rd period: Magnesium

(g) Is a noble gas (Group 8A) in the 4th period: Krypton

(h) Is a nonmetal in Group 6A and the 3rd period: Sulfur

(i) Is a metalloid in the 4th period: Germanium or Arsenic

2.108. **Carbon** has three allotropes. Graphite consists of flat sheets of carbon atoms, diamond has carbon atoms attached to four other others in a tetrahedron, and buckminsterfullerene is a 60-atom cage of carbon atoms. **Oxygen** has two allotropes. Diatomic oxygen consists of molecules containing two oxygen atoms and ozone consists of molecules containing three oxygen atoms.

2.109. Which of the following has the greater mass:

(a) 0.5 mol Na, 0.5 mol Si, 0.25 mol U

Easily done by observation and a "mental" calculation. Examine the atomic masses of each element. 0.5 mol of any element has a mass that is one-half the atomic mass. One quarter mol of U (atomic mass approximately 238 g) will have the greatest mass of these three.

(b) 9.0 g of Na, 0.50 mol Na, 1.2×10^{22} atoms Na: 0.50 mol Na will have a mass of approximately 12.5 g Na; One mole of Na will have 6.0×10^{23} atoms, so 1.2×10^{22} atoms

Na will be $\dfrac{1.2 \times 10^{22}}{6.0 \times 10^{23}} = 0.020 \text{ mol H}$ and a mass of (0.020mol)(23 g/mol) = 0.46g Na;

0.50 mol Na will have the greatest mass of these three choices.

(c) 10 atoms of Fe or 10 atoms of K

As in (a), this is done by a visual inspection of atomic masses. Fe has a greater atomic mass, so 10 atoms of Fe would have a greater mass than 10 atoms of K.

2.110. Regarding the RDA of iron for women:

$$18 \text{ mg} \cdot \frac{1 \text{ g}}{1000 \text{ mg}} \cdot \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 3.2 \times 10^{-4} \text{ mol Fe}$$

$$3.2 \times 10^{-4} \text{ mol Fe} \cdot \frac{6.02 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 1.9 \times 10^{20} \text{ atoms Fe}$$

2.111. Arrange the elements from least massive to most massive:

Select a common unit by which to compare the substances (say grams?)

(a) $\frac{3.79 \times 10^{24} \text{ atoms Fe}}{1} \cdot \frac{1 \text{ mol Fe}}{6.0221 \times 10^{23} \text{ atoms Fe}} \cdot \frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}} = 351 \text{ g Fe}$

(b) $\frac{19.921 \text{ mol } H_2}{1} \cdot \frac{2.0158 \text{ g } H_2}{1 \text{ mol } H_2} = 40.157 \text{ g } H_2$

(c) $\frac{8.576 \text{ mol C}}{1} \cdot \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 103.0 \text{ g C}$

(d) $\frac{7.4 \text{ mol Si}}{1} \cdot \frac{28.0855 \text{ g Si}}{1 \text{ mol Si}} = 210 \text{ g Si}$

(e) $\frac{9.221 \text{ mol Na}}{1} \cdot \frac{22.9898 \text{ g Na}}{1 \text{ mol Na}} = 212.0 \text{ g Na}$

(f) $\frac{4.07 \times 10^{24} \text{ atoms Al}}{1} \cdot \frac{1 \text{ mol Al}}{6.0221 \times 10^{23} \text{ atoms Al}} \cdot \frac{26.9815 \text{ g Al}}{1 \text{ mol Al}} = 182 \text{ g Al}$

(g) $\frac{9.2 \text{ mol } Cl_2}{1} \cdot \frac{70.9054 \text{ g } Cl_2}{1 \text{ mol } Cl_2} = 650 \text{ g } Cl_2$

In ascending order of mass: H_2, C, Al, Si, Na, Fe, Cl_2

2.112. Regarding the atomic weights of P and O:

0.744 g phosphorus combined with (1.704 g – 0.744 g) = 0.960 g O

Expressing the mass per atom, divide P by 4 and O by 10: $\frac{(0.744 / 4) \text{ g P}}{(0.960/10) \text{ g O}} = \frac{1.94 \text{ g P}}{1 \text{ g O}}$ telling us that a P atom is 1.94 times more massive than a O atom. Using this ratio, we can determine the mass of a P atom if we define O as 16.000 u: $\frac{16.000 \text{ u}}{1} \cdot \frac{1.94 \text{ g P}}{1 \text{ g O}} = 31.0 \text{ u P}$

2.113. Revised atomic weights:

(a) Using our present atomic weights (based on carbon-12) the relative masses of O:H are:

$$\frac{\text{at. mass O}}{\text{at. mass H}} = \frac{15.9994}{1.00794} = 15.873$$

If H ≡ 1.0000 u, the atomic mass of O would be 15.8729 · 1.0000 = 15.873 u

Similarly for carbon:

$$\frac{\text{at. mass C}}{\text{at. mass H}} = \frac{12.011}{1.00794} = 11.916$$

If H is 1.0000 u, the atomic mass of C would be 11.916 · 1.0000 = 11.916 u

The number of particles associated with one mole is:

$$\frac{11.916}{12.0000} = \frac{X}{6.02214199 \times 10^{23}} \qquad \text{and } X = 5.9802 \text{ x } 10^{23} \text{ particles}$$

(b) Using the ratio from part a

$$\frac{\text{at. mass H}}{\text{at. mass O}} = \frac{1.00794}{15.9994} = 0.0629986$$

If O ≡ 16.0000 u, the atomic mass of H would be

0.0629986 · 16.0000 = 1.00798 u

Similarly for carbon, the ratios of the atomic masses of C to O is:

$$\frac{\text{at. mass C}}{\text{at. mass O}} = \frac{12.011}{15.9994} = 0.75071,$$ and the atomic mass of C is

0.75071 · 16.0000 = 12.011 u

The number of particles associated with one mole is:

$$\frac{12.011}{12.0000} = \frac{X}{6.02214199 \times 10^{23}} \quad \text{and } X = 6.0279 \times 10^{23} \text{ particles}$$

2.114. Percent of K in a Na-K alloy:

$$68 \text{ atoms K} \cdot \frac{1 \text{ mol K}}{6.02 \times 10^{23} \text{ atoms K}} \cdot \frac{39.1 \text{ g K}}{1 \text{ mol K}} = 4.4 \times 10^{-21} \text{g K}$$

$$32 \text{ atoms Na} \cdot \frac{1 \text{ mol Na}}{6.02 \times 10^{23} \text{ atoms Na}} \cdot \frac{23.0 \text{ g Na}}{1 \text{ mol Na}} = 1.2 \times 10^{-21} \text{g Na}$$

$$\text{weight \% K} = \frac{4.4 \times 10^{-21} \text{g K}}{4.4 \times 10^{-21} \text{g K} + 1.2 \times 10^{-21} \text{g Na}} \times 100 = 78 \text{ \% K}$$

2.115. Possible compounds from ions:

	CO_3^{2-}	SO_4^{2-}
NH_4^+	$(NH_4)_2CO_3$	$(NH_4)_2SO_4$
Ni^{2+}	$NiCO_3$	$NiSO_4$

Compounds are electrically neutral—hence the total positive charge contributed by the cation (+ion) has to be equal to the total negative charge contributed by the anion (- ion). Since both carbonate and sulfate are di-negative anions, two ammonium ions are required, while only one nickel(II) ion is needed.

2.116. A strontium atom has 38 electrons (same as atomic number). When an atom of strontium forms an ion, it loses two electrons, forming an ion having the same number of electrons as the noble gas krypton (with 36 electrons).

2.117. Compound from the list with the highest mass percent of Cl:

One way to answer this question is to calculate the %Cl in each of the five compounds. An observation that each compound has the same number of Cl atoms provides a "non-calculator" approach to answering the question.

Since 3 Cl atoms will contribute the same TOTAL mass of Cl to the formula weights, the compound with the highest weight percent of Cl will also have the **lowest** weight percent of the other atom. Examining the atomic weights of the "other" atoms:

B	As	Ga	Al	P
10.81	74.92	69.72	26.98	30.97

B contributes the smallest mass of these five atoms, hence the smallest contribution to the molar masses of the five compounds—so BCl_3 has the highest weight percent of Cl.

2.118. Sample with largest number of ions:

These calculations require (1) converting mass to moles of compound (2) converting moles of compound into moles of ions, and (3) converting moles of ions into numbers of ions. Each of the calculations is identical simply by substituting the appropriate formula weight of the compound and the number of ions formed per molecule. Note especially for (d) and (e), that the polyatomic ions remain intact—giving ONE ion PER polyatomic ion (e.g. CO_3^{2-} is 1 ion, not 4).

(a) $\frac{1.0 \text{ g BeCl}_2}{1} \cdot \frac{1 \text{ mol BeCl}_2}{79.9 \text{ g BeCl}_2} \cdot \frac{3 \text{ mol ions}}{1 \text{ mol BeCl}_2} \cdot \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} = 2.3 \times 10^{22} \text{ ions}$

(b) $\frac{1.0 \text{ g MgCl}_2}{1} \cdot \frac{1 \text{ mol MgCl}_2}{95.2 \text{ g MgCl}_2} \cdot \frac{3 \text{ mol ions}}{1 \text{ mol MgCl}_2} \cdot \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} = 1.9 \times 10^{22} \text{ ions}$

(c) $\frac{1.0 \text{ g CaS}}{1} \cdot \frac{1 \text{ mol CaS}}{72.1 \text{ g CaS}} \cdot \frac{2 \text{ mol ions}}{1 \text{ mol CaS}} \cdot \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} = 1.7 \times 10^{22} \text{ ions}$

(d) $\frac{1.0 \text{ g SrCO}_3}{1} \cdot \frac{1 \text{ mol SrCO}_3}{148 \text{ g SrCO}_3} \cdot \frac{2 \text{ mol ions}}{1 \text{ mol SrCO}_3} \cdot \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} = 8.1 \times 10^{21} \text{ ions}$

(e) $\frac{1.0 \text{ g BaSO}_4}{1} \cdot \frac{1 \text{ mol BaSO}_4}{233 \text{ g BaSO}_4} \cdot \frac{2 \text{ mol ions}}{1 \text{ mol BaSO}_4} \cdot \frac{6.02 \times 10^{23} \text{ ions}}{1 \text{ mol ions}} = 5.2 \times 10^{21} \text{ ions}$

$BeCl_2$ sample has the largest number of ions.

2.119. To determine the greater mass, let's first ask the question about the molar mass of Adenine. The formula for adenine is: $C_5H_5N_5$ with a molar mass of 135.13 g. The number of molecules requested is exactly 1/2 mole of adenine molecules. So 1/2 mol of adenine molecules would have a mass of 1/2(135.13g) or 67.57 g. So 1/2 mol of adenine has a greater mass than 40.0 g of adenine.

2.120. Regarding BaF_2, $SiCl_4$, and $NiBr_2$:

(a) BaF_2: barium fluoride $SiCl_4$: silicon tetrachloride $NiBr_2$: nickel(II) bromide

(b) BaF_2 and $NiBr_2$ are ionic; $SiCl_4$ is molecular

(c) $0.50 \text{ mol BaF}_2 \cdot \frac{175 \text{ g}}{1 \text{ mol BaF}_2} = 88 \text{ g BaF}_2$

$0.50 \text{ mol SiCl}_4 \cdot \frac{170. \text{ g}}{1 \text{ mol SiCl}_4} = 85 \text{ g SiCl}_4$

$1.0 \text{ mol NiBr}_2 \cdot \frac{219 \text{ g}}{1 \text{ mol NiBr}_2} = 219 \text{ g NiBr}_2$ 1.0 mol $NiBr_2$ has the largest mass

2.121. A drop of water has a volume of 0.050 mL. Assuming the density of water is 1.00 g/cm^3, the number of molecules of water may be calculated by first determining the mass of water present. $\frac{0.050 \text{ mL}}{1} \cdot \frac{1 \text{ cm}^3}{1 \text{ mL}} \cdot \frac{1.00 \text{ g water}}{1 \text{ cm}^3} = 0.050 \text{ g water}$

The molar mass of water is 18.02 g.

The number of moles of water is: $\frac{0.050 \text{ g water}}{1} \cdot \frac{1 \text{ mol water}}{18.02 \text{ g water}} = 2.77 \times 10^{-3}\text{mol water}$

The number of molecules is: $2.77 \text{ x } 10^{-3}\text{mol} \cdot \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.7 \times 10^{21}\text{molecules}$

2.122. Regarding capsaicin:

(a) Molar mass = (18)(12.01) + (1)(14.01) + (27)(1.008) + (3)(16.00) = 305.42 g/mol

(b) $55 \text{ mg } C_{18}H_{27}NO_3 \cdot \frac{1 \text{ g}}{10^3 \text{ mg}} \cdot \frac{1 \text{ mol } C_{18}H_{27}NO_3}{305.42 \text{ g}} = 1.8 \times 10^{-4} \text{ mol } C_{18}H_{27}NO_3$

(c) $\frac{(18)(12.01) \text{ g C}}{305.42 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 70.78\% \text{ C}$ $\quad \frac{(27)(1.008) \text{ g H}}{305.42 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 8.911\% \text{ H}$

$\frac{14.01 \text{ g N}}{305.42 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 4.587\% \text{ N}$ $\quad \frac{(3)(16.00) \text{ g O}}{305.42 \text{ g } C_{18}H_{27}NO_3} \cdot 100\% = 15.72\% \text{ O}$

(d) $55 \text{ mg } C_{18}H_{27}NO_3 \cdot \frac{70.78 \text{ mg C}}{100.00 \text{ mg } C_{18}H_{27}NO_3} = 39 \text{ mg C}$

2.123. Molar mass and mass percent of the elements in $Cu(NH_3)_4SO_4 \bullet H_2O$:

Molar Mass: (1)(Cu) + (4)(N) + 12(H) + (1)(S) + (4)(O) + (2)(H) + (1)(O).

Combining the hydrogens and oxygen from water with the compound:

(1)(Cu) + (4)(N) + 14(H) + (1)(S) + (5)(O) =
(1)(63.546) + (4)(14.0067) + 14(1.0079) + (1)(32.066) + (5)(15.9994) = 245.75 g/mol

The mass percentages are:

Cu: (63.546/245.75) × 100 = 25.86% Cu

N: (56.027/245.75) × 100 = 22.80% N

H: (14.111/245.75) × 100 = 5.742% H

S: (32.066/245.75) × 100 = 13.05% S

O: (79.997/245.75) × 100 = 32.55% O

The mass of copper and of water in 10.5 g of the compound:

For Copper: $\frac{10.5 \text{ g compound}}{1} \cdot \frac{25.86 \text{ g Cu}}{100.00 \text{ g compound}} = 2.72 \text{ g Cu}$

For Water: $\frac{10.5 \text{ g compound}}{1} \cdot \frac{18.02 \text{ g } H_2O}{245.72 \text{ g compound}} = 0.770 \text{ g } H_2O$

2.124. Molecular formula and molar mass for the compounds:

(a) Ethylene glycol $C_2H_6O_2$ Molar mass = 62.07 g/mol

$$\frac{(2)(12.01)\text{ g C}}{62.07\text{ g }C_2H_6O_2}\cdot 100\% = 38.70\%\text{ C} \qquad \frac{(2)(16.00)\text{ g O}}{62.07\text{ g }C_2H_6O_2}\cdot 100\% = 51.55\%\text{ O}$$

(b) Dihydroxyacetone $C_3H_6O_3$ Molar mass = 90.08 g/mol

$$\frac{(3)(12.01)\text{ g C}}{90.08\text{ g }C_3H_6O_3}\cdot 100\% = 40.00\%\text{ C} \qquad \frac{(3)(16.00)\text{ g O}}{90.08\text{ g }C_3H_6O_3}\cdot 100\% = 53.29\%\text{ O}$$

(c) Ascorbic acid $C_6H_8O_6$ Molar mass = 176.13 g/mol

$$\frac{(6)(12.01)\text{ g C}}{176.13\text{ g }C_6H_8O_6}\cdot 100\% = 40.91\%\text{ C} \qquad \frac{(6)(16.00)\text{ g O}}{176.13\text{ g }C_6H_8O_6}\cdot 100\% = 54.51\%\text{ O}$$

Ascorbic acid has the largest percentage of carbon **and** of oxygen.

2.125. The empirical formula of malic acid, if the ratio is: $C_1H_{1.50}O_{1.25}$

Since we prefer all subscripts to be integers, we ask what "multiplier" we can use to convert each of these subscripts to integers **while** retaining the given ratio of C:H:O. Multiplying each subscript by 4 (we need to convert the 0.25 to an integer) gives a ratio of $C_4H_6O_5$.

2.126. Substance delivering greater number of Fe atoms:

$$\frac{55.85\text{ g Fe}}{151.92\text{ g }FeSO_4}\cdot 100\% = 36.76\%\text{ Fe} \qquad \frac{55.85\text{ g Fe}}{446.15\text{ g }Fe(C_6H_{11}O_7)_2}\cdot 100\% = 12.52\%\text{ Fe}$$

The tablet containing $FeSO_4$ will deliver more atoms of iron.

2.127. A compound $Fe_x(CO)_y$ is 30.70 % Fe: This implies that the balance of the mass (69.30 %) is attributable to the CO molecules. One approach is to envision CO as **one** atom, with an atomic weight of (12 + 16) or 28g. Assuming we have 100 grams, the ratios of masses are

then: $\frac{30.70\text{ g Fe}}{1}\cdot\frac{1\text{ mol Fe}}{55.845\text{ g Fe}} = 0.5497\text{ mol Fe}$ and for the "element" CO,

$\frac{69.30\text{ g CO}}{1}\cdot\frac{1\text{ mol CO}}{28.010\text{ g CO}} = 2.474\text{ mol CO}$ and the ratio of the particles is:

(dividing by 0.5497) 1 Fe: 4.5 CO, and an empirical formula of $Fe(CO)_{4.5}$.

Knowing that we don't typically like fractional atoms, we can express the atomic ratio by multiplying both subscripts by 2: $Fe_2(CO)_9$.

2.128. Regarding ephedrine:

(a) $C_{10}H_{15}NO$ Molar mass = 165.23 g/mol

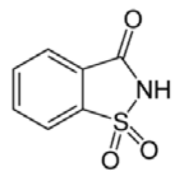
(b) $$\frac{(10)(12.01)\text{ g C}}{165.23\text{ g }C_{10}H_{15}NO} \cdot 100\% = 72.69\%\text{ C}$$

(c) $$0.125\text{ g }C_{10}H_{15}NO \cdot \frac{1\text{ mol }C_{10}H_{15}NO}{165.23\text{ g}} = 7.57 \times 10^{-4}\text{ mol }C_{10}H_{15}NO$$

(d) $$7.57 \times 10^{-4}\text{ mol }C_{10}H_{15}NO \cdot \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol }C_{10}H_{15}NO} = 4.56 \times 10^{20}\text{ molecules}$$

$$4.56 \times 10^{20}\text{ molecules} \cdot \frac{10\text{ C atoms}}{1\text{ molecule}} = 4.56 \times 10^{21}\text{ C atoms}$$

2.129. For the molecule saccharin:



(a) The formula is $C_7H_5NO_3S$

(b) Mol of saccharin associated with 125 mg:

$$\frac{125\text{ mg saccharin}}{1} \cdot \frac{1\text{ g saccharin}}{1000\text{ mg saccharin}} \cdot \frac{1\text{ mol saccharin}}{183.19\text{ g saccharin}} = 6.82 \times 10^{-4}\text{ mol saccharin}$$

(c) Mass of S in 125 mg saccharin:

$$\frac{125\text{ mg saccharin}}{1} \cdot \frac{32.07\text{ mg S}}{183.19\text{ mg saccharin}} = 21.9\text{ mg S}$$

2.130. Name the compounds and indicate which are **ionic**:

(a) chlorine trifluoride
(b) nitrogen trichloride
(c) **strontium sulfate**
(d) **calcium nitrate**
(e) xenon tetrafluoride
(f) oxygen difluoride
(g) **potassium iodide**
(h) **aluminum sulfide**
(i) phosphorus trichloride
(j) **potassium phosphate**

2.131. Formulas for compounds; identify the **ionic** compounds

(a) sodium hypochlorite	**NaClO**
(b) boron triiodide	BI_3
(c) aluminum perchlorate	**$Al(ClO_4)_3$**
(d) calcium acetate	**$Ca(CH_3CO_2)_2$**
(e) potassium permanganate	**$KMnO_4$**
(f) ammonium sulfite	**$(NH_4)_2SO_3$**
(g) potassium dihydrogen phosphate	**KH_2PO_4**
(h) disulfur dichloride	S_2Cl_2

(i) chlorine trifluoride ClF_3
(j) phosphorus trifluoride PF_3
The ionic compounds are identified by looking for a *metal*.

2.132. Supply the indicated information:

Cation	Anion	Name	Formula
NH_4^+	Br^-	ammonium bromide	NH_4Br
Ba^{2+}	S^{2-}	barium sulfide	BaS
Fe^{2+}	Cl^-	iron(II) chloride	$FeCl_2$
Pb^{2+}	F^-	lead(II) fluoride	PbF_2
Al^{3+}	CO_3^{2-}	aluminum carbonate	$Al_2(CO_3)_3$
Fe^{3+}	O^{2-}	iron(III) oxide	Fe_2O_3

2.133. Empirical and molecular formulas:

(a) For Fluorocarbonyl hypofluorite:

In a 100.00 g sample there are:

$$\frac{14.6 \text{ g C}}{1} \cdot \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 1.215 \text{ mol C}$$

$$\frac{39.0 \text{ g O}}{1} \cdot \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 2.438 \text{ mol O}$$

$$\frac{46.3 \text{ g F}}{1} \cdot \frac{1 \text{ mol F}}{15.9994 \text{ g F}} = 2.437 \text{ mol F}$$

Dividing all three terms by 1.215 gives a ratio of O:C of 2:1. Likewise F:C is 2:1

The empirical formula would be $C_1O_2F_2$ with an "empirical mass" of 82.0 g/mol Since the molar mass is also 82.0 g/mol, the molecular formula is also CO_2F_2.

(b) For Azulene:

Given the information that azulene is a hydrocarbon, if it is 93.71 % C, it is also (100.00 - 93.71) or 6.29 % H.

In a 100.00 g sample of azulene there are

$$\frac{93.71 \text{ g C}}{1} \cdot \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 7.802 \text{ mol C} \text{ and}$$

$$\frac{6.29 \text{ g H}}{1} \cdot \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 6.241 \text{ mol H}$$

The ratio of C to H atoms is: 1.25 mol C : 1 mol H or a ratio of 5 mol C:4 mol H (C_5H_4). The mass of such an empirical formula is ≈ 64. Given that the molar mass is ~128 g/mol, the molecular formula for azulene is $C_{10}H_8$.

2.134. Determine the empirical formula:

Assume 100.00 g of compound. Calculate moles of C, H, and As:

$$22.88 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 1.905 \text{ mol C} \qquad 5.76 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.71 \text{ mol H}$$

$$71.36 \text{ g As} \cdot \frac{1 \text{ mol As}}{74.922 \text{ g As}} = 0.9525 \text{ mol As}$$

Determine a ratio of atoms: $\frac{1.905 \text{ mol C}}{0.9525 \text{ mol As}} = \frac{2 \text{ mol C}}{1 \text{ mol As}}$ and $\frac{5.71 \text{ mol H}}{0.9525 \text{ mol As}} = \frac{6 \text{ mol H}}{1 \text{ mol As}}$

The empirical formula is C_2H_6As $\frac{210 \text{ g/mol}}{105.0 \text{ g/mol}} = 2$

With a molar mass of 210 g/mol, the molecular formula is $C_4H_{12}As_2$

2.135. Molecular formula of cadaverine:

Calculate the amount of each element in the compound (assuming that you have 100 g)

$$58.77 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 4.893 \text{ mol C}$$

$$13.81 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 3.70 \text{ mol H}$$

$$27.40 \text{ g N} \cdot \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 1.956 \text{ mol N}$$

The ratio of C:H:N can be found by dividing each by the smallest amount (1.956):

to give $C_{2.50}H_7N_1$ and converting each subscript to an integer (multiplying by 2)

$C_5H_{14}N_2$. The weight of this "empirical formula" would be approximately 102, hence the molecular formula is also $C_5H_{14}N_2$.

2.136. For $Ni(CO)_x$, determine the value of x:

The mass of CO is: 0.364 g $Ni(CO)_x$ – 0.125 g Ni = 0.239 g CO

Determine moles of each:

$$0.239 \text{ g CO} \cdot \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} = 0.00853 \text{ mol CO}; \; 0.125 \text{ g Ni} \cdot \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} = 0.00213 \text{ mol Ni}$$

The ratio of these is: $\frac{0.00853 \text{ mol CO}}{0.00213 \text{ mol Ni}} = \frac{4 \text{ mol CO}}{1 \text{ mol Ni}}$. The compound formula is $Ni(CO)_4$.

2.137. The empirical formula for MMT:

Moles of each atom present in 100. g of MMT:

$$49.5 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 4.12 \text{ mol C}$$

$$3.2 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 3.2 \text{ mol H}$$

$$22.0 \text{ g O} \cdot \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.38 \text{ mol O}$$

$$25.2 \text{ g Mn} \cdot \frac{1 \text{ mol Mn}}{54.938 \text{ g Mn}} = 0.459 \text{ mol Mn}$$

The ratio of C:H:O:Mn can be found by dividing each by the smallest amount (0.459): to give $MnC_9H_7O_3$.

2.138. Mass % of P; Mass of calcium phosphate to produce 15.0 kg P:

Mass % of P: $\frac{(2)(30.97) \text{ g P}}{310.18 \text{ g } Ca_3(PO_4)_2} \cdot 100\% = 19.97\% \text{ P}$

With this percentage, we calculate the amount of the phosphate to produce the requested amount of P: $15.0 \text{ kg P} \cdot \frac{100.00 \text{ kg } Ca_3(PO_4)_2}{19.97 \text{ kg P}} = 75.1 \text{ kg } Ca_3(PO_4)_2$

2.139. Chromium oxide has the formula Cr_2O_3.

The weight percent of Cr in Cr_2O_3 is: $\frac{(2 \cdot 52.00)}{\left[(2 \cdot 52.00) + (3 \cdot 16.00)\right]} \cdot 100 = 68.42\% \text{ Cr.}$

[The numerator is the sum of the mass of 2 atoms of Cr, while the denominator is the sum of the mass of 2 atoms of Cr and 3 atoms of O.]

The weight of Cr_2O_3 necessary to produce 850 kg Cr:

$$\frac{850 \text{ kg Cr}}{1} \cdot \frac{100 \text{ kg } Cr_2O_3}{68.42 \text{ kg Cr}} = 1200 \text{ kg } Cr_2O_3 \text{ (2 sf)}$$

The second fraction represents the %Cr in the oxide. Dividing the desired mass of Cr by the percent Cr (or multiplying by the reciprocal of that percentage) gives the mass of oxide needed.

2.140. Mass percent of Sb in the sulfide; Mass of stibnite in 1.00 kg of ore:

Percent of Sb: $\frac{(2)(121.8)\text{ g Sb}}{339.8\text{ g Sb}_2\text{S}_3} \cdot 100\% = 71.69\%\text{ Sb}$

Mass of Stibnite in ore: $\frac{1.00\text{ kg ore}}{1} \cdot \frac{1000\text{ g ore}}{1\text{ kg ore}} \cdot \frac{10.6\text{ g Sb}}{100.0\text{ g ore}} \cdot \frac{100.0\text{ g Sb}_2\text{S}_3}{71.69\text{ g Sb}} = 148\text{ g Sb}_2\text{S}_3$

2.141. Empirical and Molecular formula for I_xCl_y:

I_2	+	Cl_2	$\rightarrow$	I_xCl_y
0.678 g		(1.246 - 0.678)		1.246 g

Calculate the ratio of I:Cl atoms

$0.678\text{ g I} \cdot \frac{1\text{ mol I}}{126.9\text{ g I}} = 5.34 \times 10^{-3}\text{ mol I atoms}$

$0.568\text{ g Cl} \cdot \frac{1\text{ mol Cl}}{35.45\text{ g Cl}} = 1.6 \times 10^{-2}\text{ mol Cl atoms}$

The ratio of Cl:I is: $\frac{1.6\times10^{-2}\text{ mol Cl atoms}}{5.34\times10^{-3}\text{ mol I atoms}} = 3.00\ \frac{\text{Cl atoms}}{\text{I atoms}}$

The empirical formula is ICl_3 (FW = 233.3)

Given that the molar mass of I_xCl_y was 467 g/mol, we can calculate the number of empirical formulas per mole:

$\frac{467\text{g/mol}}{233.3\text{g/empiricalformula}} = 2\ \frac{\text{empirical formulas}}{\text{mol}}$ for a molecular formula of I_2Cl_6.

2.142. Empirical formula of the vanadium and sulfur compound:

Calculate the moles of each element: $2.04\text{ g V} \cdot \frac{1\text{ mol V}}{50.94\text{ g V}} = 0.0400\text{ mol V}$

$1.93\text{ g S} \cdot \frac{1\text{ mol S}}{32.065\text{ g S}} = 0.0602\text{ mol S}$. Now we can calculate the ratios of V:S

$\frac{0.0602\text{ mol S}}{0.0400\text{ mol V}} = \frac{1.5\text{ mol S}}{1\text{ mol V}} = \frac{3\text{ mol S}}{2\text{ mol V}}$ The empirical formula is V_2S_3

2.143. Mass of Fe in 15.8 kg of FeS_2:

$\%\text{ Fe in FeS}_2 = \frac{55.85\text{ g Fe}}{119.97\text{ g FeS}_2} \times 100 = 46.55\ \%\text{ Fe}$

and in 15.8 kg FeS_2: $15.8\text{ kg FeS}_2 \cdot \frac{46.55\text{ kg Fe}}{100.00\text{ kg FeS}_2} = 7.35\text{ kg Fe}$

2.144. Statements about octane that are not true:

(a) True. $0.500 \text{ mol } C_8H_{18} \cdot \dfrac{114.2 \text{ g } C_8H_{18}}{1 \text{ mol } C_8H_{18}} = 57.1 \text{ g } C_8H_{18}$

(b) True. $\dfrac{(8)(12.01) \text{ g C}}{114.2 \text{ g } C_8H_{18}} \cdot 100\% = 84.1\% \text{ C}$

(c) True. Dividing the molecular formula by 2 gives the simplest integral ratio of C and H.

(d) False. $57.1 \text{ g } C_8H_{18} \cdot \dfrac{(18)(1.008) \text{ g H}}{114.2 \text{ g } C_8H_{18}} = 9.07 \text{ g H}$

2.145. The formula of barium molybdate is $BaMoO_4$. What is the formula for sodium molybdate?

This question is easily answered by observing that the compound indicates ONE barium ion. Since the barium ion has a 2+ charge, 2 Na^+ cations would be needed, making the formula for sodium molybdate Na_2MoO_4 or choice (d).

2.146. Identity of the metal that forms MCl_4: The compound is 74.75% Cl, and we know that this mass represents 4 Cl atoms, so we determine the molar mass of the compound:

$\dfrac{74.75 \text{ g Cl}}{100.00 \text{ g } MCl_4} = \dfrac{(4)(35.453) \text{ g Cl}}{\text{molar mass } MCl_4}$; Solving for molar mass: 189.7 g/mol.

Determine the mass of the metal by difference: 189.7 g MCl_4 – (4)(35.453) g Cl = 47.9 g.

M is Ti, titanium.

2.147. Mass of Bi in two tablets of Pepto-Bismol™ ($C_{21}H_{15}Bi_3O_{12}$):

Moles of the active ingredient:

$$\frac{2 \text{ tablets}}{1} \cdot \frac{300 \times 10^{-3} \text{ g } C_{21}H_{15}Bi_3O_{12}}{1 \text{ tablet}} \cdot \frac{1 \text{ mol } C_{21}H_{15}Bi_3O_{12}}{1086 \text{ g } C_{21}H_{15}Bi_3O_{12}} = 5.52 \times 10^{-4} \text{ mol } C_{21}H_{15}Bi_3O_{12}$$

Mass of Bi: $$\frac{5.52 \times 10^{-4} \text{ mol } C_{21}H_{15}Bi_3O_{12}}{1} \cdot \frac{3 \text{ mol Bi}}{1 \text{ mol } C_{21}H_{15}Bi_3O_{12}} \cdot \frac{208.98 \text{ g Bi}}{1 \text{ mol Bi}} = 0.346 \text{ g Bi}$$

2.148. Molar mass of MO_2; Possible elements for M:

We know that the compound is 15.2% O, and the molecule has 2 O atoms. Calculate the molar mass: $\dfrac{15.2 \text{ g O}}{100 \text{ g } MO_2} = \dfrac{(2)(16.00) \text{ g O}}{\text{molar mass } MO_2}$ for which the molar mass MO_2 = 211 g/mol.

Determine the mass of the metal by difference: 211 g MO_2 – (2)(16.00) g O = 179 g.

An examination of the periodic table lets us know that M is Hf, hafnium.

2.149. What is the molar mass of ECl_4 and the identity of E?

2.50 mol of ECl_4 has a mass of 385 grams. The molar mass of ECl_4 would be:

$$\frac{385 \text{ g } ECl_4}{2.50 \text{ mol } ECl_4} = 154 \text{ g/mol } ECl_4.$$

Since the molar mass is 154, and we know that there are 4 chlorine atoms per mole of the compound, we can subtract the mass of 4 chlorine atoms to determine the mass of E.

154 - 4(35.5) = 12. The element with an atomic mass of 12 g/mol is **carbon**.

2.150. Atomic weights of A and Z:

$$\frac{15.9 \text{ g } A_2Z_3}{0.15 \text{ mol } A_2Z_3} = \frac{x}{1 \text{ mol } A_2Z_3} \text{ and } x = 106 \text{ g/mol } A_2Z_3$$

$$\frac{9.3 \text{ g } AZ_2}{0.15 \text{ mol } AZ_2} = \frac{y}{1 \text{ mol } AZ_2} \text{ and } y = 62 \text{ g/mol } AZ_2$$

For AZ_2: (atomic mass A) + (2)(atomic mass Z) = 62; and

For A_2Z_3: (2)(atomic mass A) + (3)(atomic mass Z) = 106

Rearrange the AZ_2 equation to solve for (atomic mass A), and substitute this into the equation for A_2Z_3: (2)[62 – (2)(atomic mass Z)] + (3)(atomic mass Z) = 106; The Atomic mass Z = 18 g/mol and substituting that value into either of the two original equations gives an atomic mass A = 26 g/mol.

2.151. For what value of *n*, will Br compose 0.105% of the mass of the polymer, $Br_3C_6H_3(C_8H_8)_n$?

Knowing that the 3Br atoms comprises 0.105% of the formula weight of the polymer, we can write the fraction:

$\frac{3 \cdot \text{ Br}}{\text{formula weight}} = 0.00105$ (where $3 \cdot$Br) is 3 times the atomic weight of Br. Substituting we get $\frac{239.7}{\text{formula weight}} = 0.00105$ or $\frac{239.7}{0.00105} = $ formula weight $= 2.283 \text{ x } 10^5$ g.

Noting that the "fixed" part of the formula contains 3 Br atoms, 6 C atoms, and 3 H atoms, we can calculate the mass associated with this part of the molecule (239.7 + 75.09 = 314.79).

Subtracting this mass from the total ($2.283 \text{ x } 10^5$) gives 2.280×10^5 u as the mass corresponding to the "C_8H_8" units.

Since each such unit has a mass of 104.15 (8C + 8H), we can divide the mass of *one* C_8H_8 unit into the 2.280×10^5 mass remaining:

$$\frac{2.280 \times 10^5 \text{u}}{104.15 \text{ u/C}_8\text{H}_8} = 2189 \text{ C}_8\text{H}_8 \text{ units}$$ to give a value for *n* of 2.19×10^3 (3 sf).

2.152. The molar mass of hemoglobin represents 100% of the mass of the molecule, while Fe represents 33.5% of the molecule. Set up a ratio of mass of Fe: mass of hemoglobin, which parallels this percentage of Fe in the molecule: $\frac{(4)(55.85) \text{ g Fe}}{\text{molar mass hemoglobin}} = \frac{0.335}{100}$ and solving for (molar mass hemoglobin) gives 6.67×10^4 g/mol

2.153. For the Zn-64 atom:

(a) Calculate the density of the nucleus: We can do so provided we make two assumptions:

1. The mass of the Zn atom is identical to the mass of the nucleus of the Zn atom. Given the very small masses of the electrons, this isn't a bad assumption.

2. Assume the nucleus of the Zn atom is a sphere (whose volume would be $4/3\ \pi r^3$.) Given the desired units, express the radius in units of cm:

$$\frac{4.8 \times 10^{-6}\text{nm}}{1} \cdot \frac{100 \text{ cm}}{1 \times 10^9\text{nm}} = 4.8 \times 10^{-13} \text{ cm} \quad V = \frac{4 \cdot 3.1416 \cdot (4.8 \times 10^{-13})^3}{3} = 4.6 \times 10^{-37} \text{cm}^3$$

$$D = \frac{1.06 \times 10^{-22} \text{ g}}{4.6 \times 10^{-37} \text{ cm}^3} = 2.3 \times 10^{14} \text{ g/cm}^3$$

(b) The density of the space occupied by the electrons: Express the radius of the atom in cm, as in part (a): 0.125 nm = 1.25×10^{-8} cm

Calculate the volume of that sphere: $V = \frac{4 \cdot 3.1416 \cdot (1.25 \times 10^{-8})^3}{3} = 8.18 \times 10^{-24} \text{ cm}^3$

Mass of 30 electrons (It is zinc, yes?): 30 electrons $\times 9.11 \times 10^{-28}$ g

$$D = \frac{2.733 \times 10^{-26} \text{ g}}{8.18 \times 10^{-24} \text{ cm}^3} = 3.34 \times 10^{-3} \text{ g/cm}^3$$

(c) As we've learned, the mass of the atom is concentrated in the nucleus—borne out by these densities.

2.154. Estimate the radius of a Pb atom:

(a) Volume of cube = $(1.000 \text{ cm})^3 = 1.000 \text{ cm}^3$

$$1.000 \text{ cm}^3 \text{ Pb} \cdot \frac{11.35 \text{ g Pb}}{1 \text{ cm}^3} \cdot \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \cdot \frac{6.0221 \times 10^{23} \text{atoms Pb}}{1 \text{ mol Pb}} = 3.299 \times 10^{22} \text{ atoms Pb}$$

(b) Volume of one lead atom = $\frac{0.60 \cdot 1.000 \text{ cm}^3}{3.299 \times 10^{22} \text{ atoms Pb}} = 1.819 \times 10^{-23} \text{ cm}^3$

$1.819 \times 10^{-23} \text{ cm}^3 = (^4/_3)(\pi)(\text{Pb radius})^3$ and solving, Pb radius = 1.631×10^{-8} cm

2.155. Calculate:

(a) moles of nickel—found by density **once** the volume of foil is calculated.

V = 1.25 cm × 1.25 cm × 0.0550 cm = $8.59 \times 10^{-2} \text{ cm}^3$

Mass = $\frac{8.908 \text{ g}}{1 \text{ cm}^3} \cdot 8.59 \times 10^{-2} \text{ cm}^3 = 0.766$ g Ni

$0.766 \text{ g Ni} \cdot \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} = 1.30 \times 10^{-2}$ mol Ni

(b) Formula for the fluoride salt:

Mass F = (1.261 g salt - 0.766 g Ni) = 0.495 g F

Moles F = $0.495 \text{ g F} \cdot \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 2.60 \times 10^{-2}$ mol F, so 1.30×10^{-2} mol Ni combines with 2.60×10^{-2} mol F, indicating a formula of NiF_2

(c) Name: Nickel(II) fluoride

2.156. Concerning uranium:

(a) 0.199 g U_xO_y – 0.169 g U = 0.030 g O

$0.169 \text{ g U} \cdot \frac{1 \text{ mol U}}{238.0 \text{ g U}} = 7.10 \times 10^{-4}$ mol U and $0.030 \text{ g O} \cdot \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 1.9 \times 10^{-3}$ mol O

$\frac{1.9 \times 10^{-3} \text{ mol O}}{7.10 \times 10^{-4} \text{ mol U}} = \frac{2.68 \text{ mol O}}{1 \text{ mol U}} = \frac{8 \text{ mol O}}{3 \text{ mol U}}$ so the empirical formula is U_3O_8, a mixture of uranium(IV) oxide and uranium(VI) oxide.

$7.10 \times 10^{-4} \text{ mol U} \cdot \frac{1 \text{ mol } U_3O_8}{3 \text{ mol U}} = 2.37 \times 10^{-4} \text{ mol } U_3O_8$

(b) The atomic weight of U is 238.029 u, implying that the isotope ^{238}U is the most abundant.

(c) 0.865 g – 0.679 g = 0.186 g H_2O lost upon heating

$0.186 \text{ g } H_2O \cdot \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.0103 \text{ mol } H_2O$

$0.679 \text{ g } UO_2(NO_3)_2 \cdot \frac{1 \text{ mol } UO_2(NO_3)_2}{394.0 \text{ g } UO_2(NO_3)_2} = 0.00172 \text{ mol } UO_2(NO_3)_2$

$$\frac{0.0103 \text{ mol } H_2O}{0.00172 \text{ mol } UO_2(NO_3)_2} = \frac{6 \text{ mol } H_2O}{1 \text{ mol } UO_2(NO_3)_2}$$

The formula of the hydrated compound is $UO_2(NO_3)_2 \cdot 6\ H_2O$.

2.157. The volume of a cube of Na containing 0.125 mol Na:

First we need to know the **mass** of 0.125 mol Na: $0.125\text{mol Na} \cdot \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.87 \text{ g Na}$ (3 sf)

Now we can calculate the volume that contains 2.87 g Na: (using the density given)

$$2.87\text{g Na} \cdot \frac{1 \text{ cm}^3}{0.97 \text{ g Na}} = 3.0 \text{ cm}^3$$

If the cube is a perfect cube (that is each side is equivalent in length to any other side), what is the length of one edge? $3.0 \text{ cm}^3 = \ell \text{ x } \ell \text{ x } \ell$ so 1.4 cm = length of one edge

2.158. Calculate the atomic weight:

Determine the contribution of each isotope by multiplying the abundance x isotopic mass:

(0.00193)(135.9090) + (0.00250)(137.9057) + (0.8848)(139.9053) + (0.1107)(141.9090) =

140.1046044

[Note that we convert Abundances from percentages first!]

IN THE LABORATORY

2.159. Molecules of water per formula unit of $MgSO_4$:

From 1.687 g of the hydrate, only 0.824 g of the magnesium sulfate remain.

The mass of water contained in the solid is: (1.687- 0.824) or 0.863 grams

Use the molar masses of the solid and water to calculate the number of moles of each substance present:

$$\frac{0.824 \text{ g}}{1} \cdot \frac{1 \text{ mol } MgSO_4}{120.36 \text{ g } MgSO_4} = 6.85 \times 10^{-3} \text{ mol } MgSO_4$$

$\frac{0.863 \text{ g } H_2O}{1} \cdot \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 4.79 \times 10^{-2} \text{ mol } H_2O$; the ratio of H_2O to $MgSO_4$ is:

$\frac{4.79 \times 10^{-2} \text{ mol } H_2O}{6.85 \times 10^{-3} \text{ mol magnesium sulfate}} = 6.99$ So we write the formula $MgSO_4 \cdot 7\ H_2O$.

2.160. Number of waters of hydration in alum:

Moles of anhydrous alum: 4.74 g hydrated compound – 2.16 g H_2O = 2.58 g $KAl(SO_4)_2$

$$2.58 \text{ g } KAl(SO_4)_2 \cdot \frac{1 \text{ mol } KAl(SO_4)_2}{258.2 \text{ g } KAl(SO_4)_2} = 0.00999 \text{ mol } KAl(SO_4)_2$$

Moles of water: $2.16 \text{ g } H_2O \cdot \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.120 \text{ mol } H_2O$. Now we determine the number of moles of water per mol of alum: $\frac{0.120 \text{ mol } H_2O}{0.00999 \text{ mol } KAl(SO_4)_2} = \frac{12.0 \text{ mol } H_2O}{1 \text{ mol } KAl(SO_4)_2}$

There are 12 water molecules per formula unit of $KAl(SO_4)_2$; $x = 12$

2.161. We can calculate the mass of Sn used in the compound by the **difference** of masses in the original mixture and the mass recovered after reaction: = 1.056 g – 0.601 g = 0.455 g Sn

The # mol of Sn is: $\frac{0.455 \text{ g Sn}}{118.71 \text{ g/mol}} = 0.00383 \text{ mol Sn}$

Similarly for iodine, the mass of iodine can be converted into moles of **atomic iodine.**

$\frac{1.947 \text{ g I}}{1} \cdot \frac{1 \text{ mol I}}{126.90 \text{ g I}} = 0.01534 \text{ mol I}$ The empirical formula is determined by calculating the RATIO of **iodine** to **tin**. $\frac{0.01534 \text{ mol I}}{0.00383 \text{ mol Sn}} = 4.00 \text{ mol I/mol Sn}$.

The empirical formula for the compound: Sn_1I_4, or more commonly expressed SnI_4.

2.162. Determine the empirical formula:

Assume 100.0 g of sample.

$$54.0 \text{ g C} \cdot \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.50 \text{ mol C} \qquad 6.00 \text{ g H} \cdot \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.95 \text{ mol H}$$

$$40.0 \text{ g O} \cdot \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.50 \text{ mol O;}$$

Determine the mol:mol ratios for C:O $\frac{4.50 \text{ mol C}}{2.50 \text{ mol O}} = \frac{1.8 \text{ mol C}}{1 \text{ mol O}} = \frac{9 \text{ mol C}}{5 \text{ mol O}}$

Determine mol:mol ratios for H:O $\frac{5.95 \text{ mol H}}{2.50 \text{ mol O}} = \frac{2.38 \text{ mol H}}{1 \text{ mol O}} = \frac{12 \text{ mol H}}{5 \text{ mol O}}$

Answer (d) $C_9H_{12}O_5$ is correct. Several reasons exist for an incorrect empirical formula. The students apparently did not correctly calculate the number of moles of material in 100.0 g or they improperly calculated the ratio of those moles in determining their empirical formula.

2.163. Using the student data, calculate the number of moles of $CaCl_2$ and moles of H_2O:

$$0.739 \text{ g CaCl}_2 \cdot \frac{1 \text{ mol CaCl}_2}{111.0 \text{ g CaCl}_2} = 0.00666 \text{ mol CaCl}_2$$

$$(0.832 \text{ g} - 0.739 \text{ g}) \text{ or } 0.093 \text{ g H}_2\text{O} \cdot \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0052 \text{ mol H}_2\text{O}$$

The number of moles of water/mol of calcium chloride is $\frac{0.0052 \text{ mol H}_2\text{O}}{0.00666 \text{ mol CaCl}_2} = 0.78$

This is a sure sign that they should **(c) heat the crucible again, and then reweigh it**.

2.164. Empirical formula of tin oxide: 14.710 g crucible & Sn – 13.457 g crucible = 1.253 g Sn

$$1.253 \text{ g Sn} \cdot \frac{1 \text{ mol Sn}}{118.710 \text{ g Sn}} = 0.01056 \text{ mol Sn}$$

The O present is:15.048 g (crucible & Sn & O) – 14.710 g (crucible & Sn) = 0.338 g O

$$0.338 \text{ g O} \cdot \frac{1 \text{ mol}}{15.9994 \text{ g O}} = 0.0211 \text{ mol O}$$

Now the ratio of O/Sn is: 0.0211 mol O/0.01056 mol Sn = 2 mol O/1 mol Sn

The formula is SnO_2.

SUMMARY AND CONCEPTUAL QUESTIONS

2.165. Necessary information to calculate the number of atoms in one cm^3 of iron:

A sample calculation to arrive at an exact value is shown below:

$$\frac{1.00 \text{ cm}^3}{1} \cdot \frac{7.86 \text{ g Fe}}{1 \text{ cm}^3} \cdot \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \cdot \frac{6.0221 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 8.49 \times 10^{22} \text{ Fe atoms}$$

Note that we needed: (b) molar mass of Fe, (c) Avogadro’s number, and (d) density of iron.

2.166. Relationship between abundance and atomic number:

Element abundance generally decreases with increasing atomic number (with exceptions at Li–B and Sc–Fe). Elements with an even atomic number appear to be slightly more abundant than those with an odd atomic number.

2.167. Reactivity of Mg, Ca, and Ba:

(a) Given the greater reactivity of Ca over Mg with water, one would anticipate that Ba would be even more reactive than Ca or Mg—with a more vigorous release of hydrogen gas.

(b) Reactivity of these metals increases down the group. Mg is in period 3, Ca in period 4, and Ba is in period 6. This trend is noted for Group IA as well.

2.168. One possible method involves the following steps:

(1) weigh a representative sample of jelly beans (say 10 or so) in order to determine the average mass of a jelly bean;

(2) weigh the jelly beans in the jar (subtract the mass of the empty jar from the mass of the jar filled with jelly beans);

(3) use the average mass per jelly bean and the total mass of the jelly beans in the jar to determine the approximate number of jelly beans in the jar.