Chapter 2
Atoms, Molecules and Ions

PRACTICING SKILLS

Atoms:Their Composition and Structure

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**	1.007 u

3. 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

5.

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

Isotopes

7. 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.

Isotope Abundance and Atomic Mass

9. 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.

11. 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.

13. The two stable isotopes of silver are Ag-107 and Ag-109. The masses of the isotopes are respectively: 106.9051 and 108.9047. The atomic weight of Ag on the periodic table is 107.868. Since this weight is a weighted average, a 50:50 mixture of the two would have an atomic weight exactly mid-way between the two isotopic masses. Adding the masses of these two isotopes yields (108.9047 + 106.9051) = 215.8098. One-half this value is 107.9049. Given the proximity of this number to that of the published atomic weight indicates that the two stable isotopes of silver exist in a 50:50 mix.

15. 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:

$$(\mathbf{x})(68.9257) + (1 - \mathbf{x})(70.9249) = 69.723$$

[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 %.

The Periodic Table

17. 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

19. 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: **1**; Period 6 (at.no. 55-86)

21. 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

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

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

Molecular Formulas and Models

25. The formula for sulfuric acid is H_2SO_4. The molecule is **not flat**. The O atoms are arranged around the sulfur at the corners of a tetrahedron—that is the O-S-O angles would be about 109 degrees. The hydrogen atoms are connected to two of the oxygen atoms also with angles (H-O-S) of approximately 109 degrees.

Ions and Ion Charges

27. Most commonly observed ion for:
   (a) Magnesium: 2+ —like all the alkaline earth metals
   (b) Zinc: 2+
   (c) Nickel: 2+
   (d) Gallium: 3+ (an analog of Aluminum)

29. The symbol and charge for the following ions:

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

31. 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**.

Ionic Compounds

33. 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 $BaBr_2$.

35. 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

37. Regarding cobalt oxides:

Cobalt(II) oxide	CoO	cobalt ion :	Co^{2+}
Cobalt(III) oxide	Co_2O_3		Co^{3+}

39. 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

Naming Ionic Compounds

41. 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

43. 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 $AgCH_3CO_2$

45. 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

Coulomb's Law

47. 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.

Naming Binary, Nonmetal Compounds

49. Names of binary nonionic compounds
   (a) NF_3 nitrogen trifluoride
   (b) HI hydrogen iodide
   (c) BI_3 boron triiodide
   (d) PF_5 phosphorus pentafluoride

51. Formulas for:
(a) sulfur dichloride SCl_2
(b) dinitrogen pentaoxide N_2O_5
(c) silicon tetrachloride $SiCl_4$
(d) diboron trioxide B_2O_3

Atoms and the Mole

53. The mass, in grams of:

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

(b) 1.25×10^{-3} mol Fe: $\frac{1.25 \times 10^{-3} \text{mol Fe}}{1} \bullet \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 0.0698 \text{ g Fe (3 sf)}$

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

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

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.

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

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

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

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

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

57. 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:

$$\frac{1.0 \text{ g He}}{1} \bullet \frac{1 \text{ mol He}}{4.0026 \text{ g He}} \bullet \frac{6.0221 \times 10^{23} \text{ atoms He}}{1 \text{ mol He}} = 1.5 \times 10^{23} \text{ atoms He}$$

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 calculating.

Molecules, Compounds, and the Mole

59. 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
   (b) BCl_3 (1)(10.81) + (3)(35.45) = 117.2
   (c) $C_6H_8O_6$ (6)(12.01) + (8)(1.008) + (6)(16.00) = 176.1

61. 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

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

63. 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 we 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:

$$\frac{0.0255 \text{ mol } C_3H_7OH}{1} \bullet \frac{60.10 \text{ g } C_3H_7OH1 \text{ mol } C_3H_7OH}{1 \text{ mol } C_3H_7OH} = 1.53 \text{ g } C_3H_7OH$$

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

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

65. Regarding sulfur trioxide:
   1. Amount of SO_3 in 1.00 kg: $\frac{1.00 \times 10^3 \text{g } SO_3}{1} \bullet \frac{1 \text{ mol } SO_3}{80.07 \text{ g } SO_3} = 12.5 \text{ mol } SO_3$

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

   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

Percent Composition

67. 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)(1.008) = 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

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

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

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

69. 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 = 95.612

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

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

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

Empirical and Molecular Formulas

71. 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

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

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

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

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

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

$$92.26 \text{ g C} \bullet \frac{1 \text{ mol C}}{12.011 \text{g C}} = 7.681 \text{ mol C}$$

$$7.74 \text{ g H} \bullet \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.678 \text{ mol H}$$

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 $\mathbf{C_2H_2}$** .

77. 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.

We can calculate the ratio of mol C: mol H as done in SQ75.

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

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

Calculating the atomic ratio:

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

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}.

79. Empirical and Molecular formula for Mandelic Acid:

$$63.15 \text{ g C} \bullet \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 5.258 \text{ mol C}$$

$$5.30 \text{ g H} \bullet \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 5.28 \text{ mol H}$$

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

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 so 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 $C_8H_8O_3$.

Determining Formulas from Mass Data

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

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

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

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

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

83. Formula of compound formed between zinc and iodine:

Calculate the amount of zinc and iodine present:

$$\frac{2.50 \text{ g Zn}}{1} \bullet \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} = 3.82 \times 10^{-02} \text{mol Zn}$$ and

$$\frac{9.70 \text{ g } I_2}{1} \bullet \frac{1 \text{ mol } I_2}{253.8 \text{ g } I_2} = 3.82 \times 10^{-02} \text{ mol } I_2$$ (recall that the iodine is a diatomic specie, and would be the form of iodine reacting). Note that the amount of Zinc and I_2 combined are identical, making the formula for the compound ZnI_2 . An **alternative** way of solving the problem would be to use the **atomic mass** of iodine (126.9g/ mol I) to represent 7.64 x 10^{-02} mol of I. The ratio of Zn:I would then be 1:2.

GENERAL QUESTIONS

85.

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

87. 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)

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)

Using these solutions, the following letters fit in the boxes:

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

89. (a) The average mass of one copper atom:

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

$$\frac{\$41.70}{7.0\text{ g Cu}} \bullet \frac{1.0552 \times 10^{-22}\text{ g Cu}}{1\text{ Cu atom}} = 6.286 \times 10^{-22}\text{dollars/Cu atom}$$

91. 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: Arsensic

(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

93. 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, 0.5 mol Na, 1.2×10^{22} atoms Na: 0.5 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 $\frac{1.2 \times 10^{22}}{6.0 \times 10^{23}} = 0.02$ mol Na and a mass of (0.02mol)(23 g/mol) = 0.46g Na; 0.5 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 easily 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.

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

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

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

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

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

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

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

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

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

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

97. (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:

$$\frac{11.916}{12.0000} = \frac{X}{6.02214199x10^{23}} \text{ and X } = 5.9802 \times 10^{23} \text{ particles}$$

(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:

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

0.75071 • 16.0000 = 12.011 u

The number of particles associated with one mole is:

$$\frac{12.011}{12.0000} = \frac{X}{6.02214199x10^{23}} \text{ and X } = 6.0279 \times 10^{23} \text{ particles}$$

99. 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.

101. Compound from the list with the highest weight 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.

103. 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.

105. A drop of water has a volume of 0.05 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.

$\frac{0.05\text{ mL}}{1} \bullet \frac{1\text{ cm}^3}{1\text{ mL}} \bullet \frac{1.00\ g}{1\text{ cm}^3} = 0.05\text{ g water}$. The molar mass of water is 18.02 g. The number of moles of water is then: $\frac{0.05\text{ g water}}{1} \bullet \frac{1\text{ mol water}}{18.02\text{ g water}} = 2.77 \times 10^{-3}$ mol.

The number of molecules is then: 2.77×10^{-3} mol x 6.02×10^{23} molecules/mol = 1.7×10^{21} molecules (2 sf).

107. 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 percents are:

Cu: (63.546/245.75) x 100 = 25.86% Cu

N: (56.027/245.75) x 100 = 22.80% N

H: (14.111/245.75) x 100 = 5.742% H

S: (32.066/245.75) x 100 = 13.05% S

O: (79.997/245.75) x 100 = 32.55% O

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

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

For Water: $\frac{10.5 \text{ g compound}}{1} \bullet \frac{18.02 \text{ g H}_2\text{O}}{245.72 \text{ g compound}} = 0.770 \text{ g H}_2\text{O}$

109. 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 ratio of C:H:O that we're given. Multiplying each subscript by 4 (we need to convert the 0.25 to an integer) gives a ratio of $C_4H_6O_5$.

111. 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) 28g. Assuming we have 100 grams, the ratios of masses are then:

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

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

(dividing by 0.5497) 1 Fe: 4.5 CO, so an empirical formula that is $Fe(CO)_{4.5}$. Knowing that we don't typically like fractional atoms, we can express the atomic ratio by multiplying both subscripts by 2: $Fe_2(CO)_9$.

113. For the molecule saccharine:

(a) The formula is $C_7H_5NO_3S$

(b) Mol of saccharine associated with 124 mg:

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

(c) Mass of S in 125 mg saccharine:

$$\frac{125 \times 10^{-3} \text{ g saccharine}}{1} \bullet \frac{32.07 \text{ g S}}{183.19 \text{ g saccharine}} = 0.02188 \text{ g S} \text{ or } 21.9 \text{ mg S}$$

115. Formulas for compounds; identify the ionic compounds

(a) sodium hypochlorite	$NaClO$	ionic
(b) boron triiodide	BI_3	
(c) aluminum perchlorate	$Al(ClO_4)_3$	ionic
(d) calcium acetate	$Ca(CH_3CO_2)_2$	ionic
(e) potassium permanganate	$KMnO_4$	ionic
(f) ammonium sulfite	$(NH_4)_2SO_3$	ionic
(g) potassium dihydrogen phosphate	KH_2PO_4	ionic
(h) disulfur dichloride	S_2Cl_2	
(i) chlorine trifluoride	ClF_3	
(j) phosphorus trifluoride	PF_3	

The ionic compounds are identified by noting the presence of a **metal**.

117. Empirical and molecular formulas:

(a) For Fluorocarbonyl hypofluorite:

In a 100.00 g sample there are:

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

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

$$\frac{46.3 \text{ g F}}{1} \bullet \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 2.437 \text{ mol F}$$

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

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

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

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 $C_{10}H_8$.

119. Molecular formula of cadaverine:

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

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

$$13.81 \text{ g H} \bullet \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 13.70 \text{ mol H}$$

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

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 $C_5H_{14}N_2$.

121. The empirical formula for MMT:

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

$$49.5 \text{ g C} \bullet \frac{1 \text{ mol C}}{12.0115 \text{ g C}} = 4.13 \text{ mol C}$$

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

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

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

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

123. Chromium oxide has the formula Cr_2O_3.

The weight percent of Cr in Cr_2O_3 is: [(2 x 52.00)/((2 x 52.00) + (3 x 16.00))] x 100 or (104.00/152.00) x 100 or 68.42% 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 chromium oxide necessary to produce 850 kg Cr:

$$\frac{850\text{kgCr}}{1} \bullet \frac{100\text{ kg } Cr_2O_3}{68.42\text{ kg Cr}} = 1{,}200\text{ kg } Cr_2O_3 \text{ to 2sf}$$

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.

125.

I_2	+	Cl_2	→	I_xCl_y
0.678 g		(1.246 - 0.678)		1.246 g

Calculate the ratio of I : Cl atoms

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

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

The ratio of Cl:I is: $\frac{1.6\text{x}10^{-2}\text{mol Cl atoms}}{5.34\text{x}10^{-3}\text{mol I atoms}} = 3.00\,\frac{\text{Cl atoms}}{\text{I atoms}}$

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.

127. Mass of Fe in 15.8 kg of FeS_2:

% Fe in FeS_2 = $\frac{55.85\text{gFe}}{119.97\text{g}FeS_2}$ x 100 = 46.55 % Fe

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

129. 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).

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

Moles of the active ingredient:

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

Mass of Bi:

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

$$\frac{3 \text{ mol Bi}}{1 \text{ mol } C_{21}H_{15}Bi_3O_{12}} \bullet \frac{208.98 \text{ g Bi}}{1 \text{ mol Bi}} = 0.346 \text{ g Bi}$$

133. 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**.

135. For what value of *n*, will Br compose 10.46% of the mass of the polymer, $Br_3C_6H_3(C_8H_8)_n$?

Knowing that the 3Br atoms comprises 10.46% of the formula weight of the polymer, we can write the fraction:

$\frac{3 \bullet Br}{\text{formula weight}} = 0.1046$ (where 3*Br) is 3 times the atomic weight of Br. Substituting

we get $\frac{239.7}{\text{formula weight}} = 0.1046$ or $\frac{239.7}{0.1046} =$ formula weight = 2,292 (to 4sf)

Noting that the "fixed" part of the formula contains 3Br 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 (2292) gives 1,977 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 1977 mass remaining:

$\frac{1977 \text{ u}}{104.15\text{u}/C_8H_8} = 18.98\ C_8H_8$ units to give a value for *n* of 19.

137. 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 Πr^3.) Given the desired units, convert the radius to units of cm: $\frac{4.8 \text{ x } 10^{-6}\text{nm}}{1} \bullet \frac{100 \text{ cm}}{1 \text{ x } 10^{9} \text{ nm}} = 4.8 \text{ x } 10^{-13}\text{cm}$

$$V = \frac{4 \bullet 3.1416 \bullet \left(4.8\text{x}10^{-13}\right)^3}{3} = 4.6 \text{ x } 10^{-37} \text{ cm}^3$$

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

(b) The density of the space occupied by the electrons: First express the radius of the atom in cm, as in part (a). 0.125 nm = $1.25 \text{ x } 10^{-8}$ cm.

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

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

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

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

139. Calculate:

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

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

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

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

(b) Formula for the fluoride salt:

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

$$\text{Moles F} = 0.495 \text{ g F} \bullet \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 2.60 \text{ x } 10^{-2} \text{ mol F},$$

so $1.30 \text{ x } 10^{-2}$ mol Ni combines with $2.60 \text{ x } 10^{-2}$ mol F , indicating a formula of NiF_2

(c) Name: Nickel(II) fluoride

IN THE LABORATORY

141. 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:

$$0.824\text{g} \bullet \frac{1 \text{ mol MgSO}_4}{120.36 \text{ g MgSO}_4} = 6.85 \times 10^{-3} \text{ mol of magnesium sulfate}$$

$$0.863 \text{ g H}_2\text{O} \bullet \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 4.79 \times 10^{-2} \text{ mol of water}$$;the ratio of water to $MgSO_4$ is:

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

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

First we need to know the **mass** of 0.125 mol Na:

$$0.125\text{mol Na} \bullet \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.87 \text{ g Na (3 sf)}$$

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

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

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?

$3.0 \text{ cm}^3 = \text{l x l x l}$ so 1.4 cm = length of one edge

145. Number of atoms of C in 2.0000 g sample of C.

$$2.0000 \text{ g C} \bullet \frac{1 \text{ mol C}}{12.011 \text{ g C}} \bullet \frac{6.0221 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} = 1.0028 \times 10^{23} \text{ atoms C}$$

If the accuracy of the balance is +/- 0.0001g then the mass could be 2.0001g C, so we would do a similar calculation to obtain 1.00281×10^{23} atoms C.

147. Using the student data, let's calculate the number of moles of $CaCl_2$ and mole of H_2O:

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

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

The number of moles of water/mol of calcium chloride is then

$$\frac{0.0052 \text{ mol } H_2O}{0.00666 \text{ mol } CaCl_2} = 0.78$$

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

Summary and Conceptual Questions

149. 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:

$$\frac{1.00 \text{ cm}^3}{1} \bullet \frac{7.86 \text{ g Fe}}{1 \text{ cm}^3} \bullet \frac{1 \text{ mol Fe}}{55.845 \text{ g Fe}} \bullet \frac{6.0221 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}}$$

Note that we needed: (b) atomic weight of Fe, (c) Avogadro's number, and (d) density.

151. 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. 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.

153. The hydrated salt loses water upon heating. The anhydrous cobalt salt is a deep blue, and therefore visible.