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Chapter 6: Basic Review Worksheet

1. What does the average *atomic mass* of an element represent? What unit is used for average atomic mass?

2. Define *molar mass*. Using K_2O as an example, calculate the molar mass from the atomic masses of the elements.

3. What is meant by the *percent composition* by mass for a compound? Determine the percent composition by mass for water.

4. For 5.00-g samples of each of the following substances, calculate the number of moles of the substance present, as well as the number of atoms of each type present in the sample.
   a. Cu(s)
   b. NH_3(g)
   c. $KClO_3$(s)
   d. $Ca(OH)_2$(s)

5. For the compounds in Question 4, calculate the percent by mass of each element present in the compounds.

6. Define, compare, and contrast what are meant by the *empirical* and *molecular* formulas for a substance. Give an example of each.

7. A 10.00 g sample of a compound that consists of carbon and hydrogen is found to consist of 7.99 g of carbon and 2.01 g of hydrogen. What is the empirical formula for the compound?

8. The molar mass of the compound in question 7 is 30.07 g/mol. What is the molecular formula of the compound?

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Chapter 6: Review Worksheet

1. Express the atomic mass unit in grams. Why is the average atomic mass for an element typically *not* a whole number?

2. What does one mole of a substance represent on a microscopic, atomic basis? What does one mole of a substance represent on a macroscopic, mass basis?

3. Define *molar mass*. Calculate the molar mass of H_3PO_4 from the atomic masses of the elements.

4. Describe in general terms how percent composition by mass is obtained by experiment for new compounds. How can this information be calculated for known compounds?

5. For 5.00-g samples of each of the following substances, calculate the number of moles of the substance present, as well as the number of atoms of each type present in the sample.
   a. $K_2CrO_4(s)$
   b. $AuCl_3(s)$
   c. $SiH_4(g)$
   d. $Ca_3(PO_4)_2(s)$

6. For the compounds in Question 5, calculate the percent by mass of each element present in the compounds.

7. What does an empirical formula tell us about a compound? A molecular formula? What information must be known for a compound to calculate its molecular formula?

8. An oxide of iron is found to be 70.0% iron by mass. Determine the empirical formula for this compound and name it.

9. A compound that consists of nitrogen and oxygen is found to be 30.4% nitrogen by mass. The molar mass of this compound is between 90g/mol and 100 g/mol. Determine the empirical and molecular formulas for this compound.

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Chapter 6: Challenge Review Worksheet

1. How does the text define a mole, and why have chemists defined the mole in this manner?

2. How do we know that 16.00 g of oxygen contains the same number of atoms as 12.01 g of carbon, and 22.99 g of sodium? How do we know that 106.0 g of Na_2CO_3 contains the same number of carbon atoms as does 12.01 g of carbon, but three times as many oxygen atoms as in 16.00 g of oxygen (O), and twice as many sodium atoms as in 22.99 g of sodium.

3. Define *molar mass*. Calculate the molar mass of $Al_2(SO_4)_3$ from the atomic masses of the elements.

4. Why must the molecular formula be an *integer multiple* of the empirical formula?

5. When chemistry teachers prepare an exam question on determining the empirical formula of a compound, they usually take a known compound and calculate the percent composition of the compound from the formula. They then give the students this percent composition data and have the students calculate the original formula. Using a compound of *your* choice, first use the molecular formula of the compound to calculate the percent composition of the compound. Then use this percent composition data to calculate the empirical formula of the compound.

6. How does the percent by mass of each element present in a compound depend on the mass of the sample?

7. A 151.9-mg sample of a new compound has been analyzed and found to contain the following masses of elements: carbon, 82.80 mg; hydrogen, 13.90 mg; oxygen, 55.15 mg. Calculate the empirical formula of this compound.

8. One way of determining the empirical formula of a hydrocarbon (a compound that consists of hydrogen and carbon) is to burn it in the air and measure the mass of carbon dioxide and water vapor that is produced. To do this, we must assume that all of the carbon from the hydrocarbon ends up in the carbon dioxide and all of the hydrogen from the hydrocarbon ends up in the water. Suppose you burn 2.500 g of a hydrocarbon and you collect 3.66 g carbon dioxide and 1.50 g water. Determine the empirical formula for this hydrocarbon.

9. The molar mass of the compound in question 8 is found to be around 60 g/mol. Determine the molecular formula of the compound.

significant (if a decimal point is indicated); (3) exact numbers (e.g., definitions) have an infinite number of significant figures.

When we have to round off an answer to the correct number of significant figures (as limited my whatever measurement was least precise), we do this in a particular manner. If the digit to be removed is equal to or greater than 5, the preceding digit is increased by 1. If the digit to be removed is less than 5, the preceding digit is not changed. To perform a series of calculations involving a set of data, retain all of the digits in the intermediate calculations until arriving at the final answer, and then round off to the appropriate number of significant figures.

When doing arithmetic with experimentally determined numbers, the final answer is limited by the least precise measurement. In doing multiplication or division calculations, the number of significant figures in the result should be the same as the measurement with the fewest significant figures. In performing addition or subtraction, the number of significant figures in the result is limited by the measurement with the fewest decimal places.

3. a. $0.0000009814 = 9.814 \times 10^{-7}$
   b. $14.2 \times 10^{0} = 14.2$

4. a. 351 K – 273 = 78°C
   1.80 (78°C) + 32 = 172.4 = 172°F
   b. [(72°F) – 32] / 1.80 = 22.2°C
   22.2°C + 273 = 295.2 = 295 K

5. –40°C = -40°F; 1.8C + 32 = F; if C = F, 1.8C + 32 = C, thus 32 = -0.8C and -40 = C

6. a. volume = 142.4 g/0.915 g/mL = 156 mL
   b. 4.2 lb = 1.9×10^{3} g
   volume = 1.9×10^{3} g/3.75 g/cm^3 = 507 cm^3 = 5.1×10^{2} cm^3

7. 5.27 ± 0.03 g/mL

   128.1 g/24.3 mL = 5.27 g/mL. We can get the extremes of the densities by dividing the maximum mass by the minimum volume, and the minimum mass by the maximum volume. Doing so we get: 128.2 g/24.2 mL = 5.30 g/mL and 128.0 g/24.4 mL = 5.25 g/mL.

Chapter 6: Basic Review Worksheet

1. The average atomic mass of an element represents the weighted average mass, on the relative atomic scale, of all the isotopes of an element. Average atomic masses are usually given in terms of atomic mass units

2. The molar mass of a compound is the mass in grams of one mole of the compound (6.022×10^{23} molecules of the compound), and is calculated by summing the average atomic masses of all the atoms present in a molecule (or empirical formula unit for an ionic substance) of the compound. For example, a unit of the compound K_2O contains two potassiums and one oxygen: the molar mass is obtained by adding up the average atomic masses of these atoms: molar mass K_2O = 2(39.10 g) + 1(16.00 g) = 94.20g

3. The percent composition (by mass) of a compound shows the relative amount of each element present in the compound on a mass basis. For compounds whose formulas are known (and whose molar masses are therefore known), the percentage of a given element present in the compound is given by

$$\frac{\text{mass of the element present in 1 mol of the compound}}{\text{mass of 1 mol of the compound}} \times 100$$

The percent composition of water, therefore is 11.2% hydrogen and 88.8% oxygen.

4. a. Cu (molar mass = 63.55 g)

$$\text{mol Cu} = 5.00\text{ g} \times \frac{1\text{ mol}}{63.55\text{ g}} = 0.0787\text{ mol}$$

$$\text{atoms Cu} = 0.0787\text{ mol} \times \frac{6.022\times10^{23}\text{ atoms}}{1\text{ mol}} = 4.74 \times 10^{22}\text{ atoms}$$

b. NH_3 (molar mass = 17.03 g)

$$\text{mol } NH_3 = 5.00\text{ g} \times \frac{1\text{ mol}}{17.03\text{ g}} = 0.294\text{ mol}$$

$$\text{molecules } NH_3 = 0.294\text{ mol} \times \frac{6.022\times10^{23}\text{ molecules}}{1\text{ mol}} = 1.77 \times 10^{23}\text{ molecules}$$

$$\text{atoms N} = 1.77 \times 10^{23}\text{ molecules} \times \frac{1\text{ N atom}}{1\text{ molecule}} = 1.77 \times 10^{23}\text{ atoms N}$$

$$\text{atoms H} = 1.77 \times 10^{23}\text{ molecules} \times \frac{3\text{ H atoms}}{1\text{ molecule}} = 5.31 \times 10^{23}\text{ atoms H}$$

c. $KClO_3$ (molar mass = 122.6 g)

$$\text{mol } KClO_3 = 5.00\text{ g} \times \frac{1\text{ mol}}{122.6\text{ g}} = 0.0408\text{ mol } KClO_3$$

$$0.0408\text{ mol } KClO_3 \times \frac{6.022\times10^{23}\text{ form. units}}{1\text{ mol}} = 2.46 \times 10^{22}\text{ formula units } KClO_3$$

$$\text{atoms K} = 2.46 \times 10^{22}\text{ form. un.} \times \frac{1\text{ K atom}}{1\text{ form. un.}} = 2.46 \times 10^{22}\text{ atoms K}$$

$$\text{atoms Cl} = 2.46 \times 10^{22}\text{ form. un} \times \frac{1\text{ Cl atoms}}{1\text{ form. un.}} = 2.46 \times 10^{22}\text{ atoms Cl}$$

$$\text{atoms O} = 2.46 \times 10^{22}\text{ form. units} \times \frac{3\text{ O atoms}}{1\text{ form. un.}} = 7.38 \times 10^{22}\text{ atoms O}$$

d. $Ca(OH)_2$ (molar mass = 74.096 g)

$$\text{mol } Ca(OH)_2 = 5.00 \text{ g} \times \frac{1 \text{ mol}}{74.096 \text{ g}} = 0.0675 \text{ mol } Ca(OH)_2$$

$$0.0675 \text{ mol } Ca(OH)_2 \times \frac{6.022 \times 10^{23} \text{ form. units}}{1 \text{ mol}} = 4.06 \times 10^{22} \text{ formula units } Ca(OH)_2$$

$$\text{atoms Ca} = 4.06 \times 10^{22} \text{ form. un.} \times \frac{1 \text{ Ca atom}}{1 \text{ form. un.}} = 4.06 \times 10^{22} \text{ atoms Ca}$$

$$\text{atoms O} = 4.06 \times 10^{22} \text{ form. un} \times \frac{2 \text{ O atoms}}{1 \text{ form. un.}} = 8.12 \times 10^{22} \text{ atoms O}$$

$$\text{atoms H} = 4.06 \times 10^{22} \text{ form. units} \times \frac{2 \text{ H atoms}}{1 \text{ form. un.}} = 8.12 \times 10^{22} \text{ atoms H}$$

5. a. 100% Cu

b. NH_3: $$\%N = \frac{14.01 \text{ g N}}{17.03 \text{ g}} \times 100\% = 82.27\% \text{ N}$$

$$\%H = \frac{3(1.008 \text{ g H})}{17.03 \text{ g}} \times 100\% = 17.76\% \text{ H}$$

c. $KClO_3$: $$\%K = \frac{39.10 \text{ g K}}{122.6 \text{ g}} \times 100\% = 31.89\% \text{ K}$$

$$\%Cl = \frac{35.45 \text{ g Cl}}{122.6 \text{ g}} \times 100\% = 28.92\% \text{ Cl}$$

$$\%O = \frac{3(16.00 \text{ g O})}{122.6 \text{ g}} \times 100\% = 39.15\% \text{ O}$$

d. $Ca(OH)_2$: $$\%Ca = \frac{40.08 \text{ g Ca}}{74.096 \text{ g}} \times 100\% = 54.09\% \text{ Ca}$$

$$\%O = \frac{2(16.00 \text{ g O})}{74.096 \text{ g}} \times 100\% = 43.19\% \text{ O}$$

$$\%H = \frac{2(1.008 \text{ g H})}{74.096 \text{ g}} \times 100\% = 2.721\% \text{ H}$$

6. The empirical formula of a compound represents the smallest ratio of the relative number of atoms of each type present in a molecule of the compound, whereas the molecular formula represents the actual number of atoms of each type present in a real molecule of the compound. For example, both acetylene (molecular formula C_2H_2) and benzene (molecular formula C_6H_6) have the same relative number of carbon and hydrogen atoms (one hydrogen for each carbon atom), and so have the same empirical formula (CH).

7. $7.99\text{ g C} \times \dfrac{1\text{ mol}}{12.01\text{ g}} = 0.665\text{ mol C}$

   $2.01\text{ g H} \times \dfrac{1\text{ mol}}{1.008\text{ g}} = 1.99\text{ mol H}$

   1.99/0.665 = 2.99. Thus, there are 3 hydrogen atoms for every one carbon atom. The empirical formula is CH_3.

8. The molar mass of the empirical formula CH_3 is 15.034 (12.01 + 3(1.008)). This is half the molar mass of the compound, thus the compound must have the molecular formula C_2H_6.

Chapter 6: Review Worksheet

1. 1 amu = 1.66×10^{-24} g. For example, the average atomic mass of sodium is 22.99 amu, which represents the average mass of all the sodium atoms in the world, (including all the various isotopes and their relative abundances). So that we will be able to use the mass of a sample of sodium to count the number of atoms of sodium present in the sample, we consider that every sodium atom in a sample has exactly the same mass (the *average* atomic mass). The average atomic mass of an element is typically *not* a whole number of amu's because of the presence of the different isotopes of the element, each with its own relative abundance. Since the relative abundance of an element can be any number, when the weighted average atomic mass of the element is calculated, the average is unlikely to be a whole number.

2. On a microscopic basis, one mole of a substance represents Avogadro's number (6.022×10^{23}) of individual units (atoms or molecules) of the substance. On a macroscopic basis, one mole of a substance represents the amount of substance present when the molar mass of the substance in grams is taken (for example 12.01 g of carbon will be one mole of carbon).

3. The molar mass of a compound is the mass in grams of one mole of the compound (6.022×10^{23} molecules of the compound), and is calculated by summing the average atomic masses of all the atoms present in a molecule of the compound. For example, a molecule of the compound H_3PO_4 contains three hydrogen atoms, one phosphorus atom, and four oxygen atoms: the molar mass is obtained by adding up the average atomic masses of these atoms: molar mass H_3PO_4 = 3(1.008 g) + 1(30.97 g) + 4(16.00 g) = 97.99 g

4. When a new compound is prepared (the formula is not known) the percent composition must be determined on an experimental basis. An elemental analysis must be done of a sample of the new compound to see what mass of each element is present in the sample. For example, if a 1.000 g sample of a hydrocarbon is analyzed, and it is found that the sample contains 0.7487 g of C, then the percent by mass of carbon present in the compound is

$$\frac{0.7487 \text{ g C}}{1.000 \text{ g sample}} \times 100\% = 74.87\ \%\text{C}$$

Since we can use the formula of a known compound to calculate the percent composition by mass of the compound, it is not surprising that we can go in the opposite direction – from experimentally determined percent compositions for an unknown compound, we can calculate the formula of the compound.

5. a. K_2CrO_4 (molar mass = 194.2 g)

$$\text{mol } K_2CrO_4 = 5.00 \text{ g} \times \frac{1 \text{ mol}}{194.2 \text{ g}} = 0.0257 \text{ mol } K_2CrO_4$$

$$0.0257 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ formula units}}{1 \text{ mol}} = 1.55 \times 10^{22} \text{ formula units } K_2CrO_4$$

$$\text{atoms K} = 1.55 \times 10^{22} \text{ form. units} \times \frac{2 \text{ K atoms}}{1 \text{ form. un.}} = 3.10 \times 10^{22} \text{ atoms K}$$

$$\text{atoms Cr} = 1.55 \times 10^{22} \text{ form. units} \times \frac{1 \text{ Cr atom}}{1 \text{ form. un.}} = 1.55 \times 10^{22} \text{ atoms Cr}$$

$$\text{atoms O} = 1.55 \times 10^{22} \text{ form. units} \times \frac{4 \text{ O atoms}}{1 \text{ form. un.}} = 6.20 \times 10^{22} \text{ atoms O}$$

b. $AuCl_3$ (molar mass = 303.4 g)

$$\text{mol } AuCl_3 = 5.00 \text{ g} \times \frac{1 \text{ mol}}{303.4 \text{ g}} = 0.0165 \text{ mol } AuCl_3$$

$$0.0165 \text{ mol } AuCl_3 \times \frac{6.022 \times 10^{23} \text{ form. units}}{1 \text{ mol}} = 9.94 \times 10^{21} \text{ formula units } AuCl_3$$

$$\text{atoms Au} = 9.94 \times 10^{21} \text{ form. un.} \times \frac{1 \text{ Au atom}}{1 \text{ form. un.}} = 9.94 \times 10^{21} \text{ atoms Au}$$

$$\text{atoms Cl} = 9.94 \times 10^{21} \text{ form. un} \times \frac{3 \text{ Cl atoms}}{1 \text{ form. un.}} = 2.98 \times 10^{22} \text{ atoms Cl}$$

c. SiH_4 (molar mass = 32.12 g)

$$\text{mol } SiH_4 = 5.00 \text{ g} \times \frac{1 \text{ mol}}{32.12 \text{ g}} = 0.156 \text{ mol } SiH_4$$

$$0.156 \text{ mol SiH}_4 \times \frac{6.022\times10^{23} \text{ molecules}}{1 \text{ mol}} = 9.39 \times 10^{22} \text{ molecules SiH}_4$$

$$\text{atoms Si} = 9.39 \times 10^{22} \text{ molecules} \times \frac{1 \text{ Si atom}}{1 \text{ molecule}} = 9.39 \times 10^{22} \text{ atoms Si}$$

$$\text{atoms H} = 9.39 \times 10^{22} \text{ molecules} \times \frac{4 \text{ H atoms}}{1 \text{ molecule}} = 3.76 \times 10^{23} \text{ atoms H}$$

d. $Ca_3(PO_4)_2$ (molar mass = 310.18 g)

$$\text{mol Ca}_3(\text{PO}_4)_2 = 5.00 \text{ g} \times \frac{1 \text{ mol}}{310.18 \text{ g}} = 0.0161 \text{ mol Ca}_3(\text{PO}_4)_2$$

$$0.0161 \text{ mol Ca}_3(\text{PO}_4)_2 \times \frac{6.022\times10^{23} \text{ form. units}}{1 \text{ mol}} = 9.70 \times 10^{21} \text{ formula units Ca}_3(\text{PO}_4)_2$$

$$\text{atoms Ca} = 9.70 \times 10^{21} \text{ form. un.} \times \frac{3 \text{ Ca atoms}}{1 \text{ form. un.}} = 2.91 \times 10^{22} \text{ atoms Ca}$$

$$\text{atoms P} = 9.70 \times 10^{21} \text{ form. un} \times \frac{2 \text{ P atoms}}{1 \text{ form. un.}} = 1.94 \times 10^{22} \text{ atoms P}$$

$$\text{atoms O} = 9.70 \times 10^{21} \text{ form. units} \times \frac{8 \text{ O atoms}}{1 \text{ form. un.}} = 7.76 \times 10^{22} \text{ atoms O}$$

6. a. K_2CrO_4:

$$\%K = \frac{2(39.10 \text{ g K})}{194.2 \text{ g}} \times 100\% = 40.27\% \text{ K}$$

$$\%C = \frac{52.00 \text{ g Cr}}{194.2 \text{ g}} \times 100\% = 26.78\% \text{ Cr}$$

$$\%O = \frac{4(16.00 \text{ g O})}{194.2 \text{ g}} \times 100\% = 32.96\% \text{ O}$$

b. $AuCl_3$:

$$\%Au = \frac{197.0 \text{ g Au}}{303.4 \text{ g}} \times 100\% = 64.93\% \text{ Au}$$

$$\%Cl = \frac{3(35.45 \text{ g Cl})}{303.4 \text{ g}} \times 100\% = 35.05\% \text{ Cl}$$

c. SiH_4: $\%Si = \frac{28.09\ g\ Si}{32.12\ g} \times 100\% = 87.45\%\ Si$

$\%H = \frac{4(1.008\ g\ H)}{32.12\ g} \times 100\% = 12.55\%\ H$

d. $Ca_3(PO_4)_2$: $\%Ca = \frac{3(40.08\ g\ Ca)}{310.18\ g} \times 100\% = 38.76\%\ Ca$

$\%P = \frac{2(30.97\ g\ P)}{310.18\ g} \times 100\% = 19.97\%\ P$

$\%O = \frac{8(16.00\ g\ O)}{310.18\ g} \times 100\% = 41.27\%\ O$

7. The empirical formula of a compound represents the smallest ratio of the relative number of atoms of each type present in a molecule of the compound, whereas the molecular formula represents the actual number of atoms of each type present in a real molecule of the compound. Once the empirical formula of a compound has been determined, it is also necessary to determine the molar mass of the compound before the actual molecular formula can be calculated.

8. Assume 100.0 g

$70.0\%\ Fe = 70.0\ g\ Fe \times \frac{1\ mol}{55.85\ g} = 1.25\ mol\ Fe$

$30.0\%\ O = 30.0\ g\ O \times \frac{1\ mol}{16.00\ g} = 1.88\ mol\ O$

1.88/1.25 = 1.50 = 3/2. Thus, the empirical formula is Fe_2O_3.

9. Assume 100.0 g

$30.4\%\ N = 30.4\ g\ N \times \frac{1\ mol}{14.01\ g} = 2.17\ mol\ N$

$69.6\%\ O = 69.6\ g\ O \times \frac{1\ mol}{16.00\ g} = 4.35\ mol\ O$

4.35/2.17 = 2. The empirical formula is NO_2. The molar mass of NO_2 is about 46 g/mol, which falls in the range of half the molar mass of the compound. The molecular formula must be N_2O_4.

Chapter 6: Challenge Review Worksheet

1. Chemists have chosen these definitions so that there will be simple relationship between measurable amounts of substances (grams) and the actual number of atoms or molecules present, and so that the number of particles present in samples of *different* substances can easily be compared. For example, it is known that carbon and oxygen react according to the equation

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Chemists understand this equation to mean that one carbon atom reacts with one oxygen molecule to produce one molecule of carbon dioxide, and also that one mole (12.01 g) of carbon will react with one mole (32.00 g) of oxygen to produce one mole (44.01 g) of carbon dioxide.

2. It's all relative. The mass of each substance mentioned in this question happens to be the molar mass of that substance. Each of the three samples of elemental substances mentioned (O, C, and Na) contains Avogadro's number (6.022×10^{23}) of atoms of its respective element. For the compound Na_2CO_3 given, since each unit of Na_2CO_3 contains two sodium atoms, one carbon atom, and three oxygen atoms, then it's not surprising that a sample having a mass equal to the molar mass of Na_2CO_3 should contain one molar mass of carbon, two molar masses of sodium, and three molar masses of oxygen: 12.01 g + 2(22.99 g) + 3(16.00 g) = 106.0 g.

3. The molar mass of a compound is the mass in grams of one mole of the compound (6.022×10^{23} molecules of the compound), and is calculated by summing the average atomic masses of all the atoms present in a molecule of the compound (or empirical formula unit of an ionic substance). For example, a formula unit of the compound $Al_2(SO_4)_3$ contains two aluminum, three sulfur, and twelve oxygen: the molar mass is obtained by adding up the average atomic masses of these atoms: molar mass $Al_2(SO_4)_3$ = 2(26.98 g) + 3(32.07 g) + 12(16.00 g) =342.2 g

4. The subscripts in an empirical formula represent the relative numbers of each type of atom in the molecule. The ratios of these numbers must be the same in the molecular formula (which represents the actual numbers of each type of atom in the molecule). Thus, the molecular formula is always a multiple of the empirical formula. The subscripts must be integers because we cannot have fractions of atoms.

5. Answers will vary, but consider the "known" compound phosphoric acid, H_3PO_4. First we will calculate the percentage composition (by mass) for H_3PO_4, and then we will use our results to calculate the empirical formula.

Molar mass = 3(1.008 g) + 1(30.97 g) + 4(16.00 g) = 97.99 g

$$\%H = \frac{3(1.008\text{ g H})}{97.99\text{ g}} \times 100\% = 3.086\ \%H$$

$$\%P = \frac{30.97\text{ g P}}{97.99\text{ g}} \times 100\% = 31.60\ \%P$$

$$\%O = \frac{4(16.00\text{ g O})}{97.99\text{ g}} \times 100\% = 65.31\ \%O$$

Now we will use this percentage composition data to calculate the empirical formula. We will pretend that we did not know the formula, and were just presented with a question of the type: "a compound contains 3.086 % hydrogen, 31.60 % phosphorus, and 65.31 % oxygen by mass; calculate the empirical formula."

First, we will assume, as usual, that we have 100.0 g of the compound, so that the percentages turn into masses in grams. So our sample will contain 3.086 g H, 31.60 g P, and 65.31 g O. Next we can calculate the number of moles of each element these masses represent.

$$\text{mol H} = 3.086\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 3.062\text{ mol H}$$

$$\text{mol P} = 31.60\text{ g P} \times \frac{1\text{ mol P}}{30.97\text{ g P}} = 1.020\text{ mol P}$$

$$\text{mol O} = 65.31\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 4.082\text{ mol O}$$

To get the empirical formula, divide each of these numbers of moles by the smallest number of moles: this puts things on a relative basis.

$$\frac{3.062\text{ mol H}}{1.020} = 3.002\text{ mol H}$$

$$\frac{1.020\text{ mol P}}{1.020} = 1.000\text{ mol P}$$

$$\frac{4.082\text{ mol O}}{1.020} = 4.002\text{ mol O}$$

Which gives us as the empirical formula—H_3PO_4.

6. It doesn't. The percent by mass of hydrogen in water, for example, is only dependent on the fact that there are 2 hydrogen atoms for every oxygen atom in a molecule of water, and that hydrogen has an average atomic mass of 1.008 compared to 16.00 for oxygen. Whether we have a cup of water, a gallon of water, or an ocean of water does not change the percent by mass of hydrogen and oxygen.

7. $$\text{millimol C} = 82.80\text{ mg} \times \frac{1\text{ millimol}}{12.01\text{ mg}} = 6.894\text{ millimol C}$$

$$\text{millimol H} = 13.90\text{ mg} \times \frac{1\text{ millimol}}{1.008\text{ mg}} = 13.79\text{ millimol H}$$

$$\text{millimol O} = 55.15 \text{ mg} \times \frac{1 \text{ millimol}}{16.00 \text{ g}} = 3.477 \text{ millimol O}$$

Dividing each of these numbers of millimols by the smallest number of millimols (3.447 millimol O) gives the empirical formula as C_2H_4O.

8. $$3.66 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.999 \text{ g C}$$

$$1.50 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.016 \text{ g } H_2O} = 0.168 \text{ g H}$$

$$2.500 \text{ g} - 0.999 \text{ g} - 0.168 \text{ g} = 1.333 \text{ g O}$$

$$0.999 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.0833 \text{ mol C}$$

$$0.168 \text{ g C} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.167 \text{ mol H}$$

$$1.332 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.08331 \text{ mol O}$$

Since 0.0833:0.167:0.08331 = 1:2:1, the empirical formula is CH_2O.

9. The molar mass of CH_2O is about 30 g/mol. This is half of the molar mass of the compound. Thus the molecular formula is $C_2H_4O_2$.

Chapter 7: Basic Review Worksheet

1. There are numerous ways we can recognize that a chemical reaction has taken place.

   In some reactions there may be a *color change*. For example, the ions of many of the transition metals are brightly colored in aqueous solution. If one of these ions undergoes a reaction in which the oxidation state changes, however, the characteristic color of the ion *may* be changed. For example, when a piece of zinc is added to an aqueous copper (II) ion solution (which is bright blue), the Cu^{2+} ions are reduced to copper metal, and the blue color of the solution fades as the reaction takes place.

$$Zn\ (s) + \underset{\text{blue solution}}{Cu^{2+}\ (aq)} \rightarrow Zn^{2}\ (aq) + \underset{\text{red/black solid}}{Cu\ (s)}$$