EXAM 1
PRACTICE PROBLEMS

CHEMISTRY 161A // FALL 2019

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PRACTICE PROBLEM 1

Complete the following table:

– answer –

Recognize that the atomic symbol ($^{A}_{Z}X$) gives us the mass number (A) as well as the atomic number (Z).

Symbol:	$^{90}_{40}Zr$	$^{91}_{40}Zr^{4+}$	$^{32}_{16}S^{-}$
# Protons	40	40	16
# Neutrons	50	51	16
# Electrons	40	36	17
Mass Number (A)	90	91	32
Net charge	0	4+	–1

For $^{91}_{40}Zr^{4+}$, the atomic number is Z = 40, which means that there are 40 protons in the nucleus. There would be 40 electrons in the neutral atom, but since we have Zr^{4+} there are four fewer electrons, so 36 electrons in total. Because the mass number (A = 91) is the sum of the number of protons and neutrons, we can determine there are 51 neutrons in the nucleus ($91 = 40 + n$).

PRACTICE PROBLEM 2

For each of the following entries, write the chemical name or the chemical formula.

— answer —

Name	Formula
Chromium(III) phosphate	$CrPO_4$
Manganese(IV) oxide	MnO_2
Nitrogen monoxide	NO
Aluminum sulfide	Al_2S_3
Iron(III) nitrate	$Fe(NO_3)_3$
Sodium azide	NaN_3
Sulfur trioxide	SO_3
Barium sulfite	$BaSO_3$
Vanadium(IV) chlorate	$V(ClO_3)_4$
Zinc(II) nitrate	$Zn(NO_3)_2$
Gallium acetate	$Ga(CH_3COO)_3$
Molybdenum(IV) thiocyanate	$Mo(SCN)_4$
Ammonium sulfate	$(NH_4)_2SO_4$

PRACTICE PROBLEM 3

How many hydrogen atoms are in a 50.0 g sample of ammonium carbonate?

— answer —

The molar mass of $(NH_4)_2CO_3$ is 96.09 g/mol.

Remember that:

- 1 molecule of $(NH_4)_2CO_3$ = 2 atoms N + 8 atoms H + 1 atom C + 3 atoms O
- 1 mole $(NH_4)_2CO_3$ = 2 mole N + 8 mole H + 1 mole C + 3 mole O

So:

$$50.0 \text{ g } (\mathrm{NH_4})_2\mathrm{CO_3} \times \frac{1 \text{ mol } (\mathrm{NH_4})_2\mathrm{CO_3}}{96.09 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } (\mathrm{NH_4})_2\mathrm{CO_3}} = 3.13_4 \times 10^{23} \text{ molecules } (\mathrm{NH_4})_2\mathrm{CO_3}$$

$$3.13_4 \times 10^{23} \text{ molecules } (\mathrm{NH_4})_2\mathrm{CO_3} \times \frac{8 \text{ atoms H}}{1 \text{ molecule } (\mathrm{NH_4})_2\mathrm{CO_3}} = 2.51 \times 10^{24} \text{ atoms H}$$

PRACTICE PROBLEM 4.1

A 3.25 g sample of a sugar containing only carbon, hydrogen, and oxygen was burned in excess oxygen. The mass of carbon dioxide collected was 4.76 g and the mass of water collected was 1.95 g. What is the empirical formula of the sugar?

— answer —

First, realize that the combustion of this sugar can be expressed as the unbalanced chemical equation:

$$C_xH_yO_z\ (?) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g).$$

Second, understand that all of the carbon in the sugar must be converted into the 4.76 g of CO_2 and all of the hydrogen in the sugar must be converted into the 1.95 g of H_2O. This means that we can determine the mass of C and H in the sugar as:

$$4.76\ \mathrm{g\ CO_2} \times \frac{1\ \mathrm{mol\ CO_2}}{44.01\ \mathrm{g}} \times \frac{1\ \mathrm{mol\ C}}{1\ \mathrm{mol\ CO_2}} \times \frac{12.01\ \mathrm{g}}{1\ \mathrm{mol\ C}} = 1.30\ \mathrm{g\ C} \qquad 1.95\ \mathrm{g\ H_2O} \times \frac{1\ \mathrm{mol\ H_2O}}{18.02\ \mathrm{g}} \times \frac{2\ \mathrm{mol\ H}}{1\ \mathrm{mol\ H_2O}} \times \frac{1.008\ \mathrm{g}}{1\ \mathrm{mol\ H}} = 0.218\ \mathrm{g\ H}$$

This means that the mass of O in the sugar is: $3.25\ \mathrm{g} - (1.30 + 0.218)\mathrm{g} = 1.73\ \mathrm{g\ O}$

Third, we can determine the relative number of moles of C, H, and O in the sugar as:

$$\mathrm{C} \rightarrow 1.30\ \mathrm{g\ C} \times \frac{1\ \mathrm{mol\ C}}{12.01\ \mathrm{g}} = 0.108\ \mathrm{mol\ C} \rightarrow \frac{0.108\ \mathrm{mol\ C}}{0.108} = 1\ \mathrm{mol\ C}$$

$$\mathrm{H} \rightarrow 0.218\ \mathrm{g\ H} \times \frac{1\ \mathrm{mol\ H}}{1.008\ \mathrm{g}} = 0.216\ \mathrm{mol\ H} \rightarrow \frac{0.216\ \mathrm{mol\ H}}{0.108} = 2\ \mathrm{mol\ H}$$

$$\mathrm{O} \rightarrow 1.73\ \mathrm{g\ O} \times \frac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g}} = 0.108\ \mathrm{mol\ O} \rightarrow \frac{0.108\ \mathrm{mol\ O}}{0.108} = 1\ \mathrm{mol\ O}$$

So, the empirical formula of the sugar is CH_2O.

PRACTICE PROBLEM 4.2

A 3.25 g sample of a sugar containing only carbon, hydrogen, and oxygen was burned in excess oxygen. The mass of carbon dioxide collected was 4.76 g and the mass of water collected was 1.95 g. What is the empirical formula of the sugar?

The molecular mass of the sugar is 180.6 g/mol. What is the molecular formula of the sugar?

— answer —

Previously, we found the empirical formula of the sugar is CH_2O, which has an empirical mass of 30.03 g/mol.

Because we know the molecular mass of the sugar, we can determine the molecular formula $(CH_2O)_n$:

$$n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

$$= \frac{180.6\ \text{g}}{30.03\ \text{g}}$$

$$n \approx 6$$

So, the molecular formula of the sugar is:

$$(\mathrm{CH_2O})_6 \text{ or } \mathrm{C_6H_{12}O_6}$$

PRACTICE PROBLEM 5

You perform a reaction between 0.200 g of cesium metal and 0.824 g of chlorine gas, and obtain 0.167 g of cesium chloride as a product. What is the percent yield of cesium chloride?

— answer —

$$\underline{\ 2\ }\ Cs\,(s) + \underline{\ 1\ }\ Cl_2\,(g) \rightarrow \underline{\ 2\ }\ CsCl\,(s)$$

First, determine the limiting reactant. There are several ways to determine the limiting reactant.

Method 1: *Figure out which reactant produces less product*

$$0.200\text{ g Cs}\times\frac{1\text{ mol Cs}}{132.9\text{ g}}\times\frac{2\text{ mol CsCl}}{2\text{ mol Cs}} = 0.00150\text{ mol CsCl} \qquad 0.824\text{ g Cl}_2\times\frac{1\text{ mol Cl}_2}{70.90\text{ g}}\times\frac{2\text{ mol CsCl}}{1\text{ mol Cl}_2} = 0.0232\text{ mol CsCl}$$

∴ Cs is limiting because it produces less CsCl.

Method 2: *Compare how much you need vs how much you have of each reactant*

$$0.200\text{ g Cs}\times\frac{1\text{ mol Cs}}{132.9\text{ g}}\times\frac{1\text{ mol Cl}_2}{2\text{ mol Cs}}\times\frac{70.90\text{ g}}{1\text{ mol Cl}_2} = 0.0533\text{ g Cl}_2\text{ needed}$$

∴ Cs is limiting because we have more Cl_2 gas than we need (0.053 g vs 0.824 g).

Method 3: *Compare expected and actual mole ratios*

$$0.200\text{ g Cs}\times\frac{1\text{ mol Cs}}{132.9\text{ g Cs}} = 0.00150\text{ mol Cs} \qquad 0.824\text{ g Cl}_2\times\frac{1\text{ mol Cl}_2}{70.90\text{ g Cl}_2} = 0.0116\text{ mol Cl}_2$$

∴ Cs is limiting because Cs:Cl_2 mole ratio should be 2:1, but it's actually 0.13:1.

Now, determine the theoretical yield of CsCl, and then the percent yield:

$$0.200\text{ g Cs}\times\frac{1\text{ mol Cs}}{132.9\text{ g}}\times\frac{2\text{ mol CsCl}}{2\text{ mol Cs}}\times\frac{168.35\text{ g}}{1\text{ mol CsCl}} = 0.253_3\text{ g CsCl}$$

$$\frac{0.167\text{ g CsCl}}{0.253_3\text{ g CsCl}}\times 100\% = 65.9\%$$

PRACTICE PROBLEM 6

First, balance the following chemical equation. If you start the reaction below with 1.665 g of phosphoric acid (H_3PO_4) and 2.000 g of sodium carbonate, how much (in grams) of each reactant remain after the reaction is over.

You may assume 100% for the reaction.

— answer —

$$\underline{\ 2\ }\ H_3PO_4\,(aq) + \underline{\ 3\ }\ Na_2CO_3\,(s) \rightarrow \underline{\ 2\ }\ Na_3PO_4\,(aq) + \underline{\ 3\ }\ CO_2\,(g) + \underline{\ 3\ }\ H_2O\,(l)$$

First, determine the limiting reactant. I will show only one method but see Problem 5 for alternative methods.

<u>Method 2</u>: *Compare how much you need vs how much you have of each reactant*

$$2.000\text{ g }Na_2CO_3 \times \frac{1\text{ mol }Na_2CO_3}{105.99\text{ g}} \times \frac{2\text{ mol }H_3PO_4}{3\text{ mol }Na_2CO_3} \times \frac{97.99\text{ g}}{1\text{ mol }H_3PO_4} = 1.233\text{ g }H_3PO_4\text{ needed}$$

$\therefore$ Na_2CO_3 is limiting because we have more H_3PO_4 gas than we need (1.665 g vs 1.233 g).

Because Na_2CO_3 is the limiting reactant, we will have none of it leftover (0 g Na_2CO_3).

From the above calculation, we know we need 1.233 g of H_3PO_4 to react with all 2.000 g of Na_2CO_3. So, the amount of H_3PO_4 leftover is:

$$1.665\text{ g }H_3PO_4 - 1.233\text{ g }H_3PO_4 = 0.432\text{ g }H_3PO_4$$

PRACTICE PROBLEM 7

A metallic oxide (an ionic compound) has the formula M_xO_y. The molar mass of the compound is 250.2 g/mol and the charge on the metal ion is 3+. Identity the metal ion and write the name of the ionic compound.

— answer —

If M^{3+}, then the ionic compound has the empirical formula M_2O_3.

If we assume we have 1 mol of M_2O_3, then we can subtract the mass of 3 mol O^{2-} anions from molar mass to give the mass of the remaining 2 mol M^{3+} cations:

$$250.2\text{ g} - 48.00\text{ g} = 202.2\text{ g}$$

Divide this mass of 2 mol M^{3+} to get the molar mass of just 1 mol M^{3+} (101.1 g/mol).

We can locate on periodic table as that Ru^{3+} has this molar mass.

Therefore, Ru_2O_3 is ruthenium(III) oxide.

PRACTICE PROBLEM 8

Silicone exists in three stable isotopes, as listed in the table to the right.

Calculate the average atomic mass (in amu) of a sample of natural silicon.

Isotope	^{28}Si	^{29}Si	^{30}Si
Mass (amu)	27.97693	28.97650	29.97377
Abundance	92.23%	4.67%	3.10%

— answer —

The average atomic mass of any natural sample is an average of the masses of that sample's stable isotopes, weighted by their natural abundances.

In other words,

$$\begin{aligned}\bar{M}_{Si} &= a_1 m_1 + a_2 m_2 + a_3 m_3 \\ &= (0.9223)(27.97693\ \text{amu}) + (0.0467)(28.97650\ \text{amu}) + (0.0310)(29.97377\ \text{amu}) \\ &= 25.80\ \text{amu} + 1.35\ \text{amu} + 0.929\ \text{amu} \\ \bar{M}_{Si} &= 28.09\ \text{amu}\end{aligned}$$

PRACTICE PROBLEM 9

An unknown hydrocarbon has an empirical formula of CH and its mass spectrum is shown below. What is the molecular formula of the hydrocarbon?

— answer —

In mass spectrometry, the peak farthest to the right (with the largest intensity) is called the molecular ion peak. This peak gives us the molecular mass of the molecule.

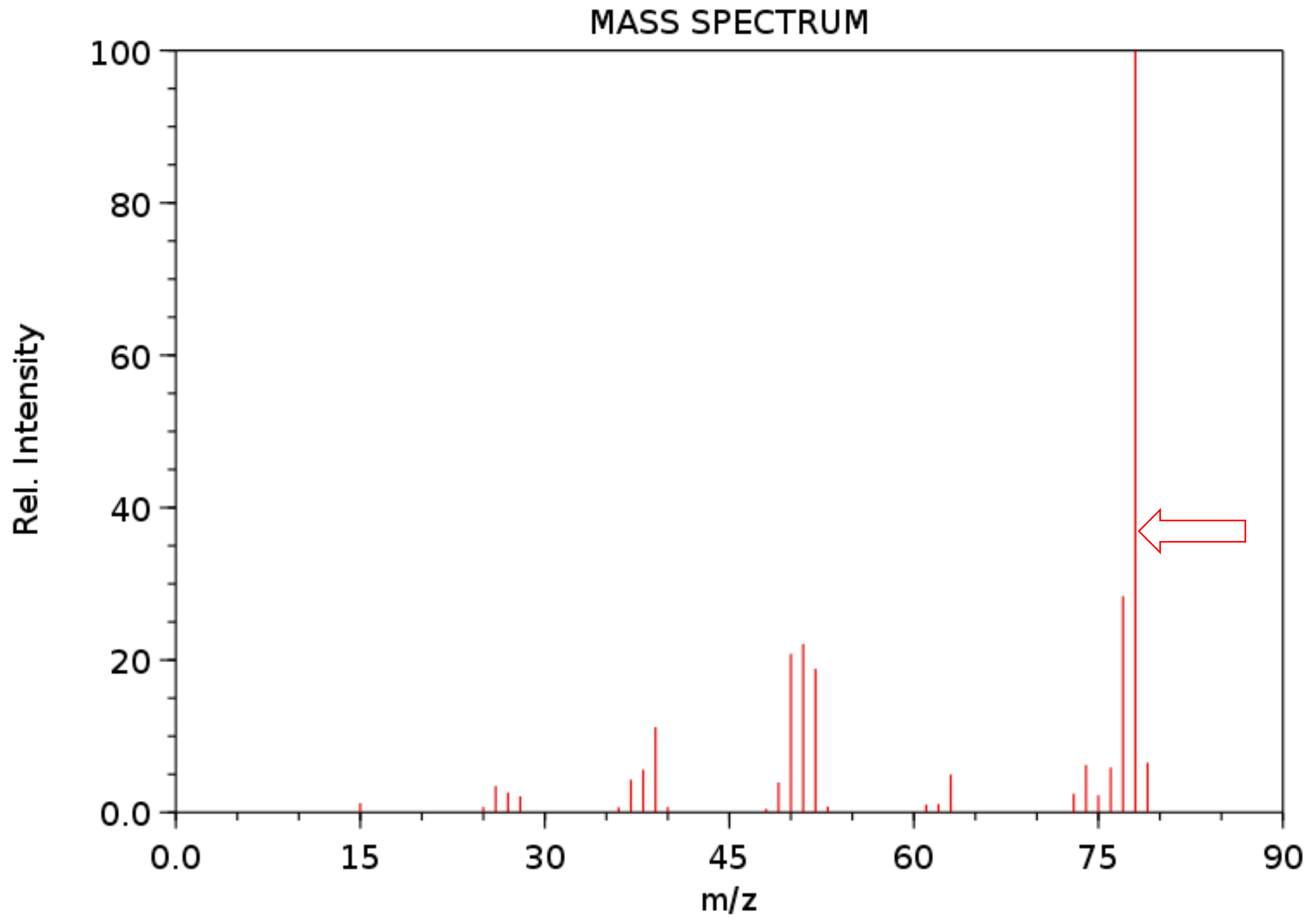
In the spectrum to the right, the molecular ion peak is ~78, so the molecular mass if ~78 g.

The empirical mass (CH) is 13.02 g.

So, the molecular formula $(CH)_n$ is:

$$n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}} = \frac{78\text{ g}}{13.02\text{ g}} \approx 6$$

$(CH)_6$ or C_6H_6



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