TOPIC A:
ATOMS AND THE ELEMENTS

Elements, atoms, mixtures and compounds
Chemistry: The Central Science 11th
Readings: 1.2, 2.3-2.4, 3.3 – 3.5

Elements

- An *element* is defined as a substance that cannot be broken down into other substances by chemical means.
- Any single element is comprised of only one type of *atom.*
- The elements are displayed on the periodic table.

Compounds & Mixtures (1.2)

- A **compound** is formed when a number of these elements bond together.
  - Compounds always have a fixed composition of atoms, i.e., they always contain the same, definite amount of each elements atoms. (H_2O, CO_2)
- **Mixtures** have varying composition and are made up of pure substances.
  - **Homogeneous** – uniform composition throughout (salt water)
  - **Heterogeneous** – properties that vary from one part of the mixture to another (chocolate chip cookie)

Chemical Formulas

- **Percent Composition** – express the mass of each element as a percentage of the total mass of the compound.
- **Empirical Formula** – the simplest whole number ratio of the atoms of each element in that compound.
- **Molecular Formula** – tells us exactly how many atoms of each element are present in the compound rather than just the simplest whole number ratio.

Formula of compounds (3.3)

- **The chemical formula** of a compound shows **the exact ratio** of the atoms of the elements that are present in the compound.
- The numbers of each element are recorded using a subscript to the right of the element symbol. When only one atom is present, the subscript of '1' is understood and assumed and therefore not written. For example, H_2O means 2 hydrogen atoms associated with 1 oxygen atom.

Percentage composition in chemical formula (Read 3.3)

- To determine the **percentage by mass composition** of an individual element within a compound, simply express the mass of each element as a percentage of the total mass of the compound.

Empirical formula (Read 3.5)

- **The *empirical formula* of a compound** is the simplest whole number ratio of the atoms of each element in that compound.
- Entirely different and unrelated compounds, with entirely different *molecular formula* (see below), may have the same empirical formula. For example, benzene C_6H_6, ethyne C_2H_2 and 1,3,5,7-cyclooctatetraene C_8H_8, are very different compounds but they each have a 1:1 ratio of C to H atoms in their molecular formula, so all have an empirical formula of CH.

Method to Determine Empirical Formula (3.5)

- a. Take the percentage of each element present in the compound and assume a sample mass of 100 g, thus converting the %'s to a mass in g of each element.
- b. Find the atomic mass of each element on the periodic table.
- c. Divide the mass in grams by the atomic mass number. This gives a quantity known as moles. (carry out values three decimal places)
- d. Find the smallest number of moles calculated in c., and divide all the results of the calculations in c. by that number, i.e., find the ratio of the moles.
- e. The results from d. should be in a whole number ratio and gives the empirical formula, i.e., the empirical formula is a ratio of the moles of the elements present.
- Note: It possible that the ratio includes a recognizable decimal (a fraction) such as .500, .333 or .250 etc. If so, then multiply all numbers by 2, 3 or 4 as appropriate, in order to produce the necessary whole number.]

Example

- For example, to calculate the empirical formula for the compound containing 40.1% carbon,
- 6.60% hydrogen and 53.3% oxygen by mass, following this method.

	C	H	O
[illegible]	40.1	6.60	53.3
[illegible]	12.011	1.0079	16.00
[illegible]	3.34	6.55	3.33
[illegible]	1	2	1
[illegible]	$C_1H_2O_1$ or CH_2O		

Task 1Ai:

- 1. Calculate the empirical formulas of the three oxides of iron shown below.
- (a) 77.78% Fe, 22.22% O

$\boxed{FeO}$

assume 100g

Fe- $77.78/55.845 = 1.39/1.39 = 1$

O - $22.22/15.99 = 1.39/1.39 = 1$

Task 1Ai:

- 1. Calculate the empirical formulas of the three oxides of iron shown below.
- (b) 70.00% Fe, 30.00% O

Fe → $70/55.85 = 1.25/1.25 = 1 \times 2 = 2$

O → $30/15.99 = 1.88/1.25 = 1.5 \times 2 = 3$

$\boxed{Fe_2O_3}$

Task 1Ai:

- 1. Calculate the empirical formulas of the three oxides of iron shown below.
- (c) 72.40% Fe, 27.60% O

Fe $72.40/55.85 = \frac{1.30}{1.30} = 1 \times 3 = 3$

O $27.60/15.99 = \frac{1.73}{1.30} = 1.33 \times 3 = 4$

Fe_3O_4

- Task 1Ai:
- 2. Which of the following are possible empirical formulae? Explain your answer. CH, CH_2, C_4H_8, CH_3O, $C_3H_6O_2$

All are in their lowest whole no. ratio

- 3. Three samples of sodium chloride are analyzed and found to contain differing %'s by mass of chlorine. What does this information alone, tell us about the three samples?

Chlorine has 3 isotopes

Molecular Formula (3.5)

- Once the empirical formula has been established, and given further appropriate data, the ***molecular formula*** of a compound can be calculated.
- The molecular formula tells us **exactly how many atoms of each element** are present in the compound rather than just the simplest whole number ratio.
- The molecular formula is a simple multiple of the empirical formula

Molecular Formula

- The molecular formula is a simple multiple of the empirical formula.
- On a previous slide,
  - an empirical formula of CH_2O has a total mass of 30 $gmol^{-1}$ (12.011 + 1.0079 +1.0079 + 16.00), but could have the molecular formula: CH_2O, $C_2H_4O_2$, $C_3H_6O_3$ etc.
  - Given that the molar mass in this example is found to be 60 g/mol, it is clear that the molecular formula is $C_2H_4O_2$, i.e., two times the empirical formula.
- To find the molecular formula it is necessary to know the molar mass and empirical formula of the compound.

- Practice 1Aii:
- 1. A hydrocarbon (a compound containing only hydrogen & carbon) is found to be 7.690% H and 92.31% C by mass. Calculate its empirical formula.

$$C \quad 92.31/12.01 = \frac{7.68}{7.62} = 1$$

$$H \quad 7.69/1.008 = \frac{7.62}{7.62} = 1$$

$\boxed{CH}$

- Practice 1Aii:
- 2. The same hydrocarbon as in question 1, has a molar mass of 78.00 $gmol^{-1}$. What is the molecular formula of the compound?

$$\frac{M.W.}{F.W.} = \frac{78}{13} = 6$$

$$CH \times 6 = \boxed{C_6H_6}$$

- Practice 1Aii:
- 3. An impure sample of the same hydrocarbon is found to have a % by mass of carbon of 80.00 %. Is this observation consistent with an impurity that contains no arbon? Explain your answer.

yes, an impurity that has carbon would have a mass % that is a multiple of 12.

Avogadro's number and the mole concept (3.4)

- In Chemistry, amounts of substances are measured in a quantity called moles (mols).
- **The mole** is a standard number of particles (atoms, ions, formula units or molecules)
  - defined as the amount of any substance that contains the same number of particles as there are C-12 atoms in 12 g of the C-12 isotope.
- The actual number of particles in a mole, known as the Avogadro constant (or number), is found to be 6.022×10^{23} particles per mole, and has the unit mol^{-1}.
- For example, 12 g of carbon atoms contains one mole of C^{12} atoms, i.e., 6.022×10^{23} atoms. If there are 6.022×10^{23} atoms (a huge number) in just 12 g of carbon, it means that atoms must be very, very tiny!
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Atomic Mass Units (AMU) (2.3)

- **<u>Atomic mass unit (AMU)</u>** is used to express the mass of an individual atom.
  - One amu has a mass of 1.66×10^{-24} g. (That's $1/6.022 \times 10^{23}$)
  - One C-12 atom has a mass of 12 amu and one atom of Cl-35 has a mass of 35 amu.
  - Converting the chlorine atom mass to grams we get, $(35)(1.66 \times 10^{-24} g) = 5.81 \times 10^{-23} g$.
- Since 1 mole contains 6.022×10^{23} particles,
  - if we take 1 mole of Cl-35 atoms they will have a mass of $(5.81 \times 10^{-23} g)(6.022 \times 10^{23}) = 35.0$ g.

Relative Atomic Mass (RAM) (2.4)

(or Atomic Mass)

- **<u>Relative atomic mass</u>** is defined as the weighted average of the masses of all the atoms in a normal isotopic sample of the element based upon the scale where 1 mole of atoms of the C^{12} isotope has a mass of exactly 12.00 g.

Relative Atomic Mass (RAM) (2.4)

(or Atomic Mass)

- Elements occur in nature as a number of different ***<u>isotopes.</u>***
- Atoms with the same number of protons and electrons, ***<u>but different numbers of neutrons</u>* <u>are called isotopes</u>**. This leads to the modification of the postulate in Dalton's atomic theory that claimed all atoms of a given element were identical, to more accurately state;
  - *All atoms of the same element contain the **<u>same number of protons and electrons</u>** but may have different numbers of neutrons.*
- Since it is **<u>the electrons</u>** in atoms that affect the chemical properties of a substance, isotopes of the same element have the **<u>same chemical properties</u>**.

Relative Atomic Mass (RAM)

(or Atomic Mass)

- **Practice 1Aiii:**
- 1. Consider the following pairs. Does either pair represent a pair of isotopes?
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  - (a) $^{40}K_{19}$ and $^{40}Ar_{18}$
  - (b) $^{90}Sr_{38}$ and $^{94}Sr_{38}$

Relative Atomic Mass (RAM)

(or Atomic Mass)

- All periodic tables have atomic mass numbers that are ***<u>not</u>* <u>integers</u>**. What does this mean? A good starting point is to analyze what it does *not* mean.
- For example, the atomic mass of Cl is often quoted on periodic tables as 35.5, and may be represented thus; $^{35.5}Cl_{17}$. This does *not* mean that there are 17 protons, 17 electrons and 18.5 neutrons in an atom of chlorine. It is not possible to have a fraction of a neutron in an atom. So what does it mean, and where does the '0.5' come from?

Relative Atomic Mass (RAM)

(or Atomic Mass)

- The non-integer values mean that there is more than one isotope of chlorine that exists in nature, in the case of chlorine, ^{35}Cl and ^{37}Cl.
- A quick calculation shows that these two species have the **same number of protons and electrons**, but different, whole numbers of neutrons (18 and 20 respectively). That is, they are isotopes of one another.
- These isotopes happen to exist naturally in the approx. abundance, ^{35}Cl, 75 % and ^{37}Cl, 25 %.

- For Chlorine, the 35.5 atomic mass on the periodic table reflects a weighted average of the Cl-35 isotope and the Cl-37 isotope.
- The atomic mass is adjusted for the percentage of the isotopes present (75% Cl-35 and 25% Cl-37).

$$\text{Average atomic mass} = \frac{\Sigma(\text{\% of each isotope})(\text{atomic mass of each isotope})}{100}$$

- $\text{Average atomic mass} = \frac{((75)(35)) + ((25)(37))}{100} = 35.5$

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- **Practice 1Aiv:**
- 1. Neon has three isotopes of masses 22, 21 and 20 amu. If the isotopes have the abundance 8.01 %, 1.99 % and 90.00 % respectively, what is the average atomic mass of neon atoms?

$22(.0801) = 1.7622$
$21(.0199) = .4179$
$20(.9000) = 18.000$
20.18

- **Practice 1Aiv:**
- 2. A naturally occurring sample of an element consists of two isotopes, one of mass 85 amu and one of mass 87 amu. The abundance of these isotopes is 71 % and 29 % respectively. Calculate the average atomic mass of an atom of this element.

$85(.71) = 60.35$
$87(.29) = 25.23$
85.58

- **Practice 1Aiv:**
- 3. If the two isotopes of gallium, Ga-69 and Ga-71 occur in the respective percentages of 62.1 and 37.9, calculate the average atomic mass of gallium atoms.

$69(.621) = 42.849$
$71(.379) = 26.909$
79.758

Molar Mass and Moles (3.4)

- Molar Mass – found by adding all of the individual masses together in one molecule (or formula unit). Expressed in grams per mole (g/mol).
- Since molar mass = g/mol we can find moles by rearranging the equation:

  Moles = mass (g) / molar mass

Molar Mass and Moles 3.4

$$\text{Moles of an element} = \frac{\text{mass of sample}}{\text{RAM}} = \frac{\text{mass of sample}}{\text{Molar Mass}}$$

$$\text{Moles of a molecular compound} = \frac{\text{mass of sample}}{\text{RMM}} = \frac{\text{mass of sample}}{\text{Molar Mass}}$$

$$\text{Moles of an ionic compound} = \frac{\text{mass of sample}}{\text{RFM}} = \frac{\text{mass of sample}}{\text{Molar Mass}}$$

LO 1.4 *Task 1Av*

1. *What is the mass of one mole of sodium chloride, NaCl?*
2. *How many moles of Ca atoms are there in 140. g of calcium?*
3. *How many moles of $CuBr_2$ are there in 0.522 g of copper (II) bromide?*
4. *How many moles of CO_2 molecules are there in 23.0 g of carbon dioxide?*
5. *How many 'particles' are present in each of the chemicals in questions 1-4 above?*

- **Practice 1Av:**
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- 1. What is the mass of one mole of sodium chloride, NaCl?

Na = 22.98
Cl = 35.45
58.43 amu or $\boxed{58.43\text{g/mol}}$

- 2. How many moles of Ca atoms are there in 140. g of calcium?

$$\frac{1.40\text{g Ca}}{} \left| \frac{1\text{mole Ca}}{140.08\text{g}} \right. = \boxed{.035 \text{ mol Ca}}$$

- **Practice 5A:**
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- 3. How many moles of $CuBr_2$ are there in 0.522 g of copper (II) bromide?

$$.522\text{g } CuBr_2 \left| \frac{1\text{mole}}{223.35\text{g}} \right. = \boxed{.0023\text{mole}}$$

- 4. How many moles of CO_2 molecules are there in 23.0 g of carbon dioxide?

$$23.0\text{g } CO_2 \left| \frac{1\text{mole}}{44.01\text{g } CO_2} \right. = \boxed{.523 \text{ moles}}$$

- **Practice 5A:**
- 5. How many 'particles' are present in each of the chemicals in questions 1-4 above?

1) 6.02×10^{23} particles

2) $.035 \text{ mol Ca} \left| \frac{6.02 \times 10^{23}}{1\text{mole}} \right. = 2.107 \times 10^{22}$ particles

3) $.0023 \text{ mole } CuBr_2 \left| \frac{6.02 \times 10^{23}\ p}{1\text{mole } CuBr_2} \right. = 1.385 \times 10^{21}$ particles

4) $.523 \text{ moles} \left| \frac{6.02 \times 10^{23}}{1\text{mole Ca}} \right. = 3.15 \times 10^{23}$ particles