Some definitions:

**Nominal mass**: The integer mass/mass number of the most abundant naturally occurring stable isotope.

Molecules: The sum of the nominal masses of the elements in its molecular formula

**Molecular mass** is the mass of a given molecule, it is measured in units of daltons (Da) or amu (u).

\[ 1 \text{ Da} = 1 \text{ u} = 1.660 \, 538 \, 782(83) \times 10^{-27} \, \text{kg} \]

**Molar mass** is the mass of one mole of a substance. Molar mass is closely related to the molecular weight or the relative molar mass. Measured in grams (g/mol).

Numerically molecular mass and molar mass are very close.

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**Monoisotopic mass**

The monoisotopic mass is the sum of the masses of the atoms in a molecule where mass of the principal (most abundant) isotope for each element is used in the calculation. For typical organic compounds this also results in the lightest isotope being selected.

For some heavier atoms the principle isotope is not the lightest isotope.

Ions made up of the principal isotopes of atoms making up the molecule.

Monoisotopic mass is typically expressed in atomic mass units (u), also called Daltons (Da).

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**Determinations of the molecular formula of a compound is fundamental to chemists and material scientists. Three steps involved;**

1. Qualitative analysis (for elements)
   
   Na fusion followed by chemical tests
   Combustion followed by GC (quantification)

2. Quantitative analysis (to find the elemental composition)
   
   Empirical formula (= molecular formula, sometimes)

3. Molecular mass
   
   - Vapor density (gaseous compounds)
   
   \[ P = \frac{w}{(MW)V}RT \]

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**Cryoscopic methods**

\[ \Delta T = Km \quad m = \text{molality} \]

\[ \pi = iMRT \quad M = \text{molarity} \]

\[ i = \text{number of moles of solute particles} \]

Vapor pressure osmometry
Use of Raoult’s Law
Neutralization reactions (titrations) for acids and bases

**Mass Spectrometry**

**Molecular ion**

Mass spectrometry (MS) is an analytical technique that can measure the mass of a molecule. The molecule is ionized (parent ion) usually by EI. Often it yields a radical univalent positive ion (OE ionic species) and detected. The mass of the un-fragmented molecular ion is the mass of the molecule.

The parent ion composition could change for CI (EE molecular ion).

MS is also used as any other spectroscopic method to elucidate the chemical structures of molecules.

The MS principle consists of ionizing chemical compounds to generate charged molecules or molecular fragments and measurement of their mass/charge ratio.

(MS measures the masses of all isotopic combinations)
With the molecular mass, specially the mono-isotopic mass of the parent ion of the molecule in hand, a plausible molecular formula can be predicted.

These are available in tabulated form (e.g. Chapter 1 – Appendix A) and electronically:

http://library.med.utah.edu/masspec/elcomp.htm
http://www.colby.edu/chemistry/NMR/NMR.html

The electronic outputs must be accepted with caution. However they are extremely useful.

Fragment Finder.
Formula Finder

Caution: With EI, in some cases, the highest mass peak in a mass spectrum is not the molecular ion. This is due to the instability of the high energy parent ion.

For example, alcohols sometimes give a parent peak that has a molar mass of M-18.

"Molar Mass Finder" provides a list of possible molecular ion molar masses from given fragment masses.

Molecular Property: Degree of Unsaturation, U

Also referred to as ‘index of hydrogen deficiency, unsaturation index, double bond equivalents’ is the number of pi bonds and rings in organic molecules and odd electron molecular ions. U is calculated from the molecular formula.

\[
U = n + 1 - \frac{m - t}{2}
\]

\(n\) = # of tetravalent atoms in formula
\(m\) = # monovalent atoms in formula
\(t\) = # trivalent atoms in formula
Example:

\[ C_6H_7CINO_4 \]

\[ U = n + 1 - \frac{m - t}{2} \]

\[ U = 8 + 1 - \frac{7 - 1}{2} = 6 \]

U<0: impossible combination for OE molecular ion/molecule, indicates O or N atom presence.

Fractional U values are unlikely.

U is also a useful tool to ascertain whether a molecule is an odd or even electron molecular ion.

For EE molecular ions, calculate 'U' as above & subtract \( \frac{1}{2} \).

**Nitrogen Rule**

(for molecules and OE molecular ions, \( M^{+*} \))

Nearly all molecules are even electron systems.

If the nominal mass is even, the compound contains an even number of N atoms; 0, 2, 4, 6, ...  
If the nominal mass is odd, the compound contains an odd number of N atoms; 1, 3, 5, 7, ...

**Nitrogen Rule**

(for EE molecular ions)

For EE precursor ions \((M+H)^+, (M+Na)^+, (M+Cl)^-\) etc; the N rule is:

if the nominal mass is odd, the compound contains an even number of N atoms; 0, 2, 4, 6, ...  
If the nominal mass is even, the compound contains an odd number of N atoms; 1, 3, 5, 7, ...

**Molecular Formula from The Rule of Thirteen**

With nominal molecular mass, (say from EI-MS; \( M^{+} \)) a base formula (algebraically valid and physically meaningful if compound contains C, H only) can be generated from the following relationship:

\[ \frac{M}{13} = n + \frac{r}{13} \]

to give the base formula: \( C_nH_{r+1} \)

For this base formula, the \( U \) would be:

\[ U = \frac{n - r + 2}{2} \]

Nominal mass = Mass of the fragment with lightest/abundant isotopes.

Example:

If \( M^{+} = 142 \) what are the possible molecular formulae.

a. only C and H are present
b. only C, H, and O are present
c. only C, H, and N, are present

With the base formula in hand one can calculate the other possible molecular formulae that would correspond to the nominal molecular mass.

To obtain the other possible molecular formulae, add atoms and subtract an equivalent mass comprising of C and H atoms from the 'base formula'.

<table>
<thead>
<tr>
<th>Element added</th>
<th>Subtract</th>
<th>( \Delta U )</th>
<th>Element added</th>
<th>Subtract</th>
<th>( \Delta U )</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>H_{12}</td>
<td>7</td>
<td>^{35}Cl</td>
<td>C_{2}H_{13}</td>
<td>3</td>
</tr>
<tr>
<td>H_{12}</td>
<td>C</td>
<td>7</td>
<td>^{79}Br</td>
<td>C_{2}H_{13}</td>
<td>-3, 4</td>
</tr>
<tr>
<td>O</td>
<td>CH_{3}</td>
<td>1</td>
<td>F</td>
<td>CH_{3}</td>
<td>2</td>
</tr>
<tr>
<td>N</td>
<td>CH_{3}</td>
<td>1/2</td>
<td>Si</td>
<td>CH_{2}</td>
<td>1</td>
</tr>
<tr>
<td>S</td>
<td>C_{2}H_{5}</td>
<td>2</td>
<td>P</td>
<td>C_{2}H_{5}</td>
<td>2</td>
</tr>
</tbody>
</table>

Calculate the base formula

Base formula: 142 / 13 = 10 + 12 / 13

\( C_{10}H_{22} \) with calculated \( U = 0 \)

Some of other possible formula

add 1 C:

\[ C_{10}H_{22} + C - H_{12} = C_{11}H_{18} \]

\( U = 7 \) because \( \Delta U = 7 \)

add 12 H:

\[ C_{10}H_{22} + H_{12} - C = C_{12}H_{34} \]

\( U = -7 \) because \( \Delta U = -7 \) (impossible)
add O:
\[ \text{C}_{10}\text{H}_{22} + \text{O} \rightarrow \text{C}_{9}\text{H}_{18}\text{O} \]
\[ U = 1 \]

add another O:
\[ \text{C}_{9}\text{H}_{18}\text{O} + \text{O} \rightarrow \text{C}_{8}\text{H}_{16}\text{O}_2 \]
\[ U = 2 \]

add another O:
\[ \text{C}_{8}\text{H}_{16}\text{O}_2 + \text{O} \rightarrow \text{C}_{7}\text{H}_{14}\text{O}_3 \]
\[ U = 3 \]

add another O:
\[ \text{C}_{7}\text{H}_{14}\text{O}_3 + \text{O} \rightarrow \text{C}_{6}\text{H}_{12}\text{O}_4 \]
\[ U = 4 \]

add N:
\[ \text{C}_{10}\text{H}_{22} + \text{N} \rightarrow \text{C}_{9}\text{H}_{20}\text{N} \]
\[ U = 0.5 \]
(unlikely)

add another N:
\[ \text{C}_{9}\text{H}_{20}\text{N} - \text{CH}_2 = \text{C}_{8}\text{H}_{18}\text{N}_2 \]
\[ U = 1 \]

Isotope distribution of a very large molecule.