HCM CITY, VIETNAM 2016

Welcome to

CHeM$_g$IST$_a$R$_h$Y

The Atom

*What is an **atom**?*

- Atoms are the smallest unit. They are the building blocks of matter.
- Everything in the world is made of atom !
- Atoms are too small in size : Size of Earth : Soda can = Soda can : Atom

*Atoms are made up of **Protons, Neutrons and Electrons***

- The mass and charge of these subatomic particles is **really small**, so **relative mass** and **relative charge** are used instead.

  1 unit of charge is 1.602 x 10^{-19} coulombs

  1 unit of mass is 1.661 x 10^{-27} kg

- *The symbol and the charge of proton, neutron, electron*

Atomic Structure	Particle	Symbol	Actual charge/C	Relative charge
Nucleus	proton	p	+1.60 x 10^{-19}	1+
	neutron	n	0	0
Electron cloud	electron	e	−1.60 x 10^{-19}	1–

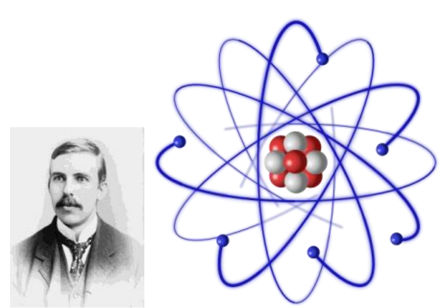
- *The mass of proton, neutron, electron*

Atomic Structure	Particle	Symbol	Actual mass/kg	Relative mass
Nucleus	proton	p	1.67 x 10^{-27}	1
	neutron	n	1.67 x 10^{-27}	1
Electron cloud	electron	e	9.11 x 10^{-31}	$0.00055 = \frac{1}{2000}$ (which is usually ignored)

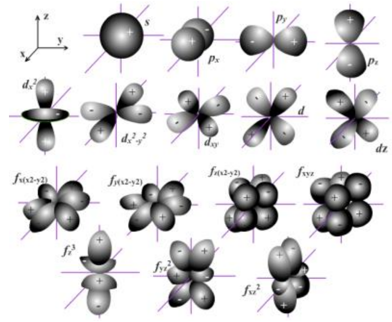
The mass of an electron is negligible compared to a prton or a neutron – this means you an usually ignore it.

1 a) Draw a diagram showing the structure of the atom, labelling each part.

b) Where is the mass concentrated in an atom, and what makes up most of the volume of an atom?



RUTHERFORD model

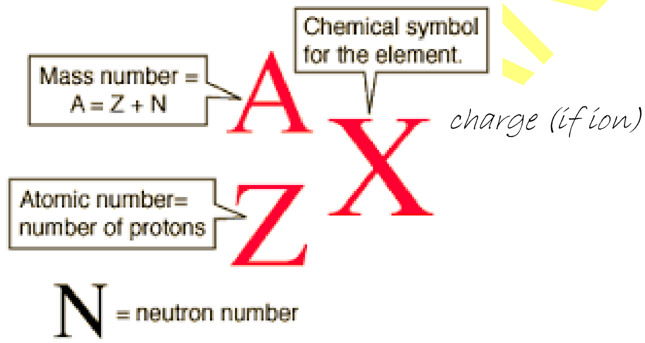


NIELS BOHR and . . . model 

The shapes of **s, p, d, f – orbitals**

Niels Bohr proposed that the electrons aren't on any random orbit around the nucleus, they are on "special" orbits. The size and shape of an orbital is drawn so that there is a 90% probability of finding the electron within its boundary.

Atomic (Nuclear) Symbols** show Numbers of **Subatomic Particles

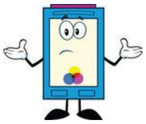


The atomic number (**Z**) of an element (on the periodic table) is the number of **protons** = the number of **electrons** in a **neutral atom** (an atom with zero charge).
The mass number (**A**) is the **sum** of the number of **protons and neutrons** in the nucleus.
Number of neutrons = A – Z

 2

Nuclear symbol	Atomic number, Z	Mass number, A	Protons	Electrons	Neutrons
$^{7}_{3}Li$					7 – 3 = 4
$^{24}_{12}Mg$					
$^{80}_{35}Br$					

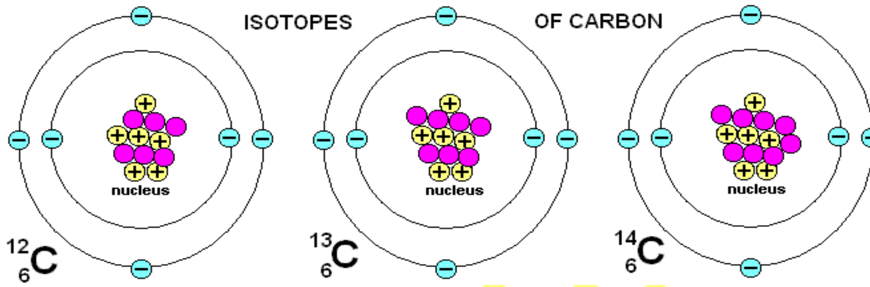
 3 Identify the particle that has 13 protons, 14 neutrons and 10 electrons. *[ASChemistry/page39]*



4 Use the periodic table to write symbols for the following species:

a) 19 protons, 20 neutrons, 18 electrons.
b) 8 protons, 8 neutrons, 10 electrons.
c) 1 proton, 2 neutrons, 1 electron.
d) 82 protons, 126 neutrons, 80 electrons.
e) 53 protons, 74 neutrons, 54 electrons.

***Isotopes** of An Element are atoms with the **Same** Number of **Protons** (atomic number - Z) but **Different** Numbers of **Neutrons** (mass numbers - A).*



5 Here's another example: **magnesium** (atomic number 12), **oxygen** (atomic number 8) has 3 naturally occurring isotopes.

${}^{24}_{12}Mg$ (79%) : 12 protons, 12 electrons, . . . neutrons
${}^{25}_{12}Mg$ (10%) : 12 protons, 12 electrons, . . . neutrons
${}^{26}_{12}Mg$ (11%) : 12 protons, 12 electrons, . . . neutrons
${}^{16}_{8}O$ (99.76%) : . . . protons, . . . electrons, . . . neutrons
${}^{17}_{8}O$ (0.04%) : . . . protons, . . . electrons, . . . neutrons
${}^{18}_{8}O$ (0.20%) : . . . protons, . . . electrons, . . . neutrons

6 Hydrogen, deuterium and tritium are all isotopes of each other.

a) Identify one similarity and one difference between these isotopes. [2 marks]
b) Deuterium can be written as ^{2}H. Determine the number of protons, neutrons and electrons in neutral deuterium. [3 marks]
c) Write a nuclear symbol for tritium, given that it has 2 neutrons. [1 mark]

Calculating ***Relative Atomic Mass*** *of* ***An Element (A_r)***

The relative atomic mass can be calculated by the formula:

$$A_r = \frac{\Sigma \text{ (perentage abundance of each isotope x relative mass of each isotope)}}{100}$$

7 Lithium has two isotopee, ^{6}Li and ^{7}Li. Use the date below to calculate the relative atomic mass of lithium. *[ASChemistry/page20]*

Isotope	Relative isotopic mass m/e	% abundance
^{6}Li	6.015	7.42
^{7}Li	7.016	92.58

8 Use the following isotpic abundance data for Ar and Mg to calculate their relative atomic mass (A_r)

Isotope	$^{40}_{18}Ar$	$^{36}_{18}Ar$	$^{38}_{18}Ar$
Relative abundance (%)	99.60	0.34	0.06

Isotope	^{24}Mg	^{25}Mg	^{26}Mg
Relative abundance (%)	78.99	10.00	11.01

9 Use the following isotpic abundance data for Ni to calculate their relative atomic mass (A_r)

Isotope	$^{58}_{28}Ni$	$^{60}_{28}Ni$	$^{61}_{28}Ni$	$^{62}_{28}Ni$	$^{64}_{28}Ni$
Relative abundance (%)	68.27%	26.10%	1.13%	3.59%	0.91%

10 The relative atomic mass of copper is 63.5. Calculate the relative abundance of the two copper isotopes with relative isotopic masses of 63.0 and 65.0. *[ASChemistry/page39]*

Ions *have* ***Different*** *Numbers of* ***Protons*** *and* ***Electrons***

- **Negative** ions have **more electrons** than protons

E.g. $_{35}Br^{-}$ The 1– charge means that there's . . . more electron than there are protons. Br has . . . . protons, so Br^{-} must have . . . . electrons.
The overall charge = +. . . . . – . . . . . = ...

- **Positive** ions have **fewer electrons** than protons

E.g. $_{12}Mg^{2+}$ The 2+ charge means that there's . . . fewer electrons than there are protons. Mg has . . . . protons, so Mg^{2+} must have . . . . electrons.
The overall charge = +. . . . . – . . . . . = ...

11 $^{23}_{11}Na^{+}$: . . . protons, . . . electrons, . . . neutrons

$^{24}_{12}Mg^{2+}$: . . . protons , . . . electrons, . . . neutrons

$^{24}_{12}Mg^{+}$: . . . protons, . . . electrons, . . . neutrons

$^{27}_{13}Al^{3+}$: . . . protons, . . . electrons, . . . neutrons

$^{35}_{17}Cl^{-}$: . . . protons, . . . electrons, . . . neutrons

$^{37}_{17}Cl^{-}$: . . . protons, . . . electrons, . . . neutrons

$^{16}_{8}O^{2-}$: . . . protons, . . . electrons, . . . neutrons

$^{17}_{8}O^{2-}$: . . . protons, . . . electrons, . . . neutrons

$^{14}_{7}N^{3-}$: . . . protons, . . . electrons, . . . neutrons

12 Potassium was first isolated by Sir Humphrey Davy in 1807. It has two main isotopes, ^{39}K and ^{41}K. Both isotopes are able to form a positive ion with a single charge.
Complete the table below.

Particle	Number of protons	Number of neutrons	Number of electrons
^{39}K	. . . .	. . . .	. . . .
$^{41}K^{+}$	. . . .	. . . .	. . . .

13 Deduce the number of protons, neutrons and electrons in the following species:

$^{1}_{1}H$ $^{14}_{6}C$ $^{17}_{8}O$ $^{132}_{54}Xe$ $^{235}_{92}U$

$^{4}_{2}He^{2+}$ $^{27}_{31}Al^{3+}$ $^{1}_{1}H^{+}$ $^{45}_{21}Sc^{3+}$ $^{37}_{17}Cl^{-}$ $^{32}_{16}S^{2-}$ $^{15}_{7}N^{3-}$

14 Study the table below and answer the queations that follow.

a) Identify the two elements of A – F that are isotopes.
b) Identify which of A – F are neutral atoms.
c) Identify which of A – F are cations.
d) Identify which of A – F are anions.
e) Identify which of A – F have the same electron configuration. *[ASChemistry/page39]*

Element	Number of protons	Number of neutrons	Number of electrons
A	19	20	18
B	17	18	17
C	12	12	10
D	17	20	16
E	35	44	36
F	18	22	18

15 This question relates to the atoms or ions A to D:

A. ${}^{32}_{16}S^{2-}$,   B. ${}^{40}_{18}Ar$,   C. ${}^{30}_{16}S$,   D. ${}^{42}_{20}Ca$.

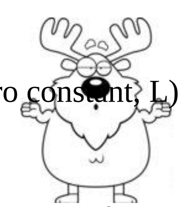
a) Identify the similarity for each of the following pairs, justifying your answer in each case.
   (i) A and B. [2 marks]
   (ii) A and C. [2 marks]
   (iii) B and D. [2 marks]
b) Which two of the atoms or ions are isotopes of each other? Explain your reasoning. [2 marks]

Atoms and Moles

Relative Masses** are Masses Compared to **Carbon – 12

- The mass of ${}^{12}C$ = 1.661 x 10^{-27} kg
- E.g: A natural sample of chlorine contains a mixture of ${}^{35}Cl$ (75%) and ${}^{37}Cl$ (25%), so
  - The **Relative isotopic mass** are 35 and 37
  - The **Relative atomic mass, A_r** is 35.5 (the **average mass** of an atom of an element)
- E.g: The **Relative molecular mass, M_r** (H_2O) = (2x1) + 16 = 18

***A Mole** is just a **Very Large Number** of Particles*



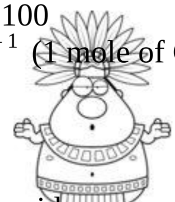
- One mole is roughly **6 x 10^{23} particles** (the Avogadro constant, L)
- Number of moles = $\dfrac{Number\ of\ \ particles\ you\ have}{6\,x10^{23}}$
- *Example:* I have 1.5x10^{24} carbon atoms. How many moles of carbon is this?

Number of moles = $\frac{1.5x10^{24}}{6\,x\,10^{23}} = 2.5$ moles

*Molar Mass (M) is the Mass of **One Mole** of something*

- **Molar mass** is just the **same as** the **relative mass** (the only difference is you stick a "**g mol^{-1}**" for grams per mole on the end

  Example: Find the molar mass of $CaCO_3$?
  - M_r of $CaCO_3$ = 40 + 12 + (3 x 16) = 100
  - So the molar mass, M, is 100 g mol^{-1} (1 mole of $CaCO_3$ weighs 100 gam)



- Number of moles = $\frac{Mass\ of\ substance}{M}$

  Example: How many moles of aluminium oxide are present in 1.5 gam of Al_2O_3?
  - M of Al_2O_3 = (2 x 27) + (3 x 16) = 102 g mol^{-1}
  - Number of moles of Al_2O_3 = $\frac{5.1g}{102\ g\ mol^{-1}}$ = 0.05 moles

  Worked example: 0.0222 mol of an oxide of sulfur has a mass of 1.42 g. Calculate its molar mass.
  - Molar mass of the sulfur oxide $= \frac{mass}{moles} = \frac{1.42\ g}{0.0222\ mol} = 64.0\ \ g\ mol^{-1}$

- Moles and mass to identity

  Worked example: 0.0250 mol of a group 2 sulfate has a mass of 4.60 g. Calculate the molar mass of the sulfate and hence identify the group 2 metal ion in the compound.

  Answer
  - Molar mass $= \frac{mass}{moles} = \frac{4.60g}{0.0250mol} = 184\ g\ mol^{-1}$
  - The formula of group 2 sulfates is of the form MSO_4, where M represents the group 2 metal. Of the 184 g mol^{-1}, 32.1 + (4x16.0) = 96.1 g comes from the SO_4 group. Therefore, the molar mass of the group 2 metal = 184 g mol^{-1} – 96.1 g mol^{-1} = 87.9 g mol^{-1}
  - From the periodic table, the group 2 metal that has a molar mass nearest to 87.9 g mol^{-1} is strontium, **Sr**.

  Worked example: 0.100 mol of hydrated sodium carbonate, $Na_2CO_3.xH_2O$, has a mass of 28.6 g. Calculate its molar mass and hence the number of molecules of water of crystallisation.

  Answer
  - Molar mass of $Na_2CO_3.xH_2O$ $= \frac{mass}{moles} = \frac{28.6g}{0.100mol} = 286\ g\ mol^{-1}$
  - Mass of Na_2CO_3 = (2x23)+12.0+(3x16.0) = 106.0 g
  - Mass of water = 286 – 106 = 180 g
  - There are 180 g of water in 1 mol of the hydrated solit
  - Mmount (moles) of water $= \frac{180\ g}{18.0\ g\ mol^{-1}} = 10.0\ mol\ in\ 1\ mol\ of\ solid$
  - The number of molecules of water of crystallisation is 10.

***One mole** of **any Gas** always has the **Same Volume** (if **Temperature** and **Pressure** stay **the Sam**)*

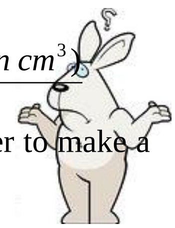
- At **room temperature and pressure** (r.t.p is 298 K (25 ^{0}C) and 101.3 kPa), this happens to be **24 dm^3**

- Number of moles = $\dfrac{Volume\,in\,cm^3}{24\,000} = \dfrac{Volume\,in\,dm^3}{24}$

  Example: How many moles are there in 6 dm^3 of oxygen gas at r.t.p?

  Number of moles = $\dfrac{6}{24}$ = 0.25 moles of oxygen molecules

*The **Concentration** of a Solution is how may **moles** are dissolved per **1** dm^3 **of solution**. The units are **mol dm^{-3}***



- Number of moles = $Concentration\ x\,Volume\ (in\ dm^3) = \dfrac{Concentration\ x\,Volume\ (in\ cm^3)}{1000}$
- *Example:* What mass of sodium hydroxide needs to be dissolved in 50 cm^3 of water to make a 2 mol dm^{-3} solution?
  - Number of moles = $\dfrac{2x50}{1000} = 0.1$ moles of NaOH
  - M of NaOH = 23 + 16 + 1 = 40 g mol^{-1}
  - Mass = number of moles x M = 0.1 x 40 = 4 g

Parts Per Million** is used for **Really Small Quantities

- **Xenon** makes up only **0.000 009%** of the **atmosphere**. Numbers this small are a pain to work with
- So if there's **0.000 009 parts** of Xe in every **one hundred parts of air**, you can multiply both quantities by **10 000** to make the quantity **large enough** to work with, like this:

  0.000 009% = $\dfrac{0.000\ 009}{100} = \dfrac{0.000\ 009\ \ x\,10\ 000}{100\ \ x\,10\ 000} = \dfrac{0.09}{1\,000\ 000} = 0.09\ \ parts\ \ per\ million$

  So there's **0,09 ppm xenon**. The atmosphere also contains **0.01 ppm carbon** and **0.3 ppm** nitrous oxide.

Question 16

1. Calculate the amount (in moles) of water in 4.56 g.
2. Calculate the amount (in moles) of calcium chloride, $CaCl_2$ in 7.89 g.
3. Calculate the amount (in moles) of sucrose, $C_{12}H_{22}O_{11}$ in 3.21 g. The molar mass of sucrose is 342.0 g mol^{-1}.
4. Calculate the amount (in moles) of 1.11 g of calcium carbonate, $CaCO_3$.
5. Calculate the amount (in moles) of 2.22 g of barium hydroxide, $Ba(OH)_2$.

ASChemistry.page 56 ; 60

Question 17 Calculate the mass of substance present in the following cases:

a) 100 mol of sodium metal.
b) 0.05 moles of Cl_2
c) 250 moles of Fe_2O_3
d) 0.0333 mol of sulfuric acid, H_2SO_4.
e) 0.36 moles of ethanoic acid, CH_3COOH. [2 marks]
f) What mass of H_2SO_4 is needed to produce 60 cm^3 of a 0.25 mol dm^{-3} solution? [2 marks]

Question 18

1. 0.0500 mol of an organic acid had a mass of 3.00 g. Calculate the molar mass of the acid. *ASChemistry. page 57*

2. Calculate the relative molecular mass of the following substances and suggest a possible identity of each substance:

a) 0.015 moles, 0.42 g c) 0.55 moles, 88 g b) 0.0125 moles, 0.50 g
d) 2.25 moles, 63 g e) 0.00125 moles, 0.312 g

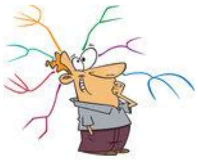
Question 19

1. Calculate the number of:
a) molecules in 1.2 mol of water ; in 1.2 g of water.
b) nitrogen atoms in 100 g of N_2.
c) oxygen atoms in 0.0100 mol of carbon dioxide, CO_2.
d) molecules in 3.33 g of methane, CH_4 (molar mass = 16.0 g mol^{-1}).
e) sodium ions in 2.22 g of sodium sulfate, Na_2SO_4 (molar mass = 142.1 g mol^{-1}).
f) hydroxide ions in 10.0 g of barium hydroxide.

2. Calculate the mass of the following substances:
a) 2.5×10^{23} molecules of N_2 b) 1.5×10^{24} molecules of CO_2 c) 2×10^{20} atoms of Mg

3. Calculate:
a) the number of atoms in 1 mol of CO_2
b) the number of chloride ions in 1 mol of calcium chloride, $CaCl_2$

(The Avogadro constant, $N_A = 6.02 \times 10^{23}$ mol^{-1}) *[ASChemistry/page12]*



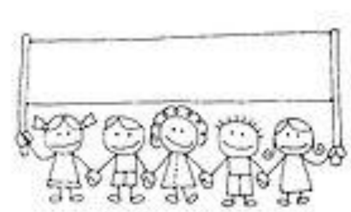
Question 20 0.0185 mol of hydrated magnesium sulfate, $MgSO_4.xH_2O$, has a mass of 4.56 g. Calculate the number of molecules of water of crystallisation in the hydrated salt. *[ASChemistry/page60]*

Empirical and Molecular Formulas

The Empirical formula	The Molecular formula
gives the ***smallest whole number ratio*** of atoms in a molecule.	▫ gives the ***actual numbers of atoms*** in a molecule. ▫ is made up of a whole number of empirical units.

***Example 1*:**
- The molecular formula: $C_8H_6O_4$
- The empirical formula: $C_4H_3O_2$
- So, you can understand $(C_4H_3O_2)_2$

***Example 2*:**
- The molecular formula: $C_3H_7O_2N$
- The empirical formula: $C_3H_7O_2N$
- So, the empirical and molecular formulas are the same



***Empirical Formulas** are calculated from **Experimental Results** (given as **masses or percentages**)*

***Example 1*:** When a hydrocarbon is burnt in excess oxygen, **4.4 gam** of carbon dioxide and **1.8 gam** of water are made. What is the **empirical formula** of the hydrocarbon?

Answer
- No. of moles of $CO_2 = \frac{mass}{M} = \frac{4.4}{44} = 0.1$ mol ; so no. of moles of C = 0.1
- No. of moles of $H_2O = \frac{mass}{M} = \frac{1.8}{18} = 0.1$ mol ; so no. of moles of H = 0.2
- Ratio C **:** H = 0.1 **:** 0.2 *(divide each number of moles by the smallest number – in this case it's 0.1)* = 1 **:** 2
- So the empirical formula must be CH_2.

***Example 2*:** A compound is found to have percentage composition 56.5% potassium, 8.7% carbon and 34.8% oxygen by mass. Calculate its empirical formula?

Answer
- In **100 gam** of compound there are:

$\frac{56.5}{39} = 1.449$ moles of K $\frac{8.7}{12} = 0.725$ moles of C $\frac{34.8}{16} = 2.175$ moles of O

(divide each number of moles by the smallest number – in this asea it's 0.725)

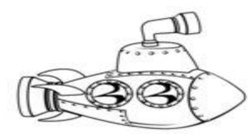
- Ratio K : C : O = 2 : 1 : 3
- So the empirical formula must be K_2CO_3.

***Molecular Formulas** are calculated from **Experimental Data** too*

Example 1: 4.6 gam of an alcohol, with molar mass 46 g mol $^{-1}$, is burnt in excess oxygen. It produces 8.8 gam of carbon dioxide and 5.4 gam of water. Calculate the empirical formula for the alcohol and then its molecular formula?

Answer

- No. of moles of CO_2 = $\frac{mass}{M} = \frac{\ldots\ldots\ldots}{\ldots\ldots\ldots}$ = ... mol ; so no. of **moles of C** = ...
- No. of moles of H_2O = $\frac{mass}{M} = \frac{\ldots\ldots\ldots}{\ldots\ldots\ldots}$ = ... mol ; so no. of **moles of H** = ...
- Mass of C = ... x .... = ... g
  Mass of H = ... x .... = ... g
  Mass of O = ... – (........... + ...) = ... g
  No. of **moles of O** = $\frac{mass}{M} = \frac{\ldots\ldots\ldots}{\ldots\ldots\ldots}$ = ... mol



- Ration C : H : O = ... = ...
- **Empirical formula** = ...
- Mass of empirical formula = ... = ...g
- In this example, the **mass of the empirical formula** equals the **molecular mass**, so the empirical and molecular formulas are the same
- **Molecular formula** = ...

***Example 2*:** An alkene has the empirical formula CH_2. 0.075 mol of the alkene has a mass of 2.1 g. Calculate its molar mass and hence its molecular formula.

Answer

- Mass of empirical formula = ... = ...g
- Molar mass = $\frac{mass}{moles} = \frac{2.1g}{0.075mol}$ = 28 g mol $^{-1}$
- So there are $\frac{\ldots\ldots\ldots}{\ldots\ldots\ldots}$ = ... empirical units in the molecule
- The molecular formula must be the empirical formula x ...., so the molecular formula = ...

Question 21

a) In an experiment to determine the formula of an oxide of copper, 2.8 g of the oxide was heated in a stream of hydroben gas until there was no further mass change, 2.5 g of copper remained. Calculate the empirical formula of the oxide? *[4 marks]*

b) When 1.2 g of magnesium ribbon is heated in air, it burns to form a white powder, which has a mass of 2 g. What the empirical formular of the powder? *[2 marks]*

c) 3.36 g of iron join with 1.44 g of oxygen in an oxide of iron. What is the empirical formula of the oxide?

d) A compound of rubidium and oxygen contains 72.6% rubidium by mass. Calculate its empirical formula.

e) An organic compound contains the following by mass: carbon 17.8% ; hydrogen 3.0% ; bromine 79.2%. Calculate its empirical formula.

f) Find the empirical formula of the compound containing C 22.02%, H 4.59% and Br 73.39% by mass.

Question 22

a) Analysis of a hydrocarbon showed that 7.8 g of the hydrocarbon contained 0.6 g of hydrogen and that the relative molecular mass was 78. Find the molecular formula of the hydrocarbon.

b) When 19.8 g of an organic acid, A, is burnt in excess oxygan, 33 g of carbon dioxide and 10.8 g of water are produced. Calculate the empirical formula for A and hence its molecular formula, if $M_r(A) = 132$. *[4 marks]*

c) Hydrocarbon X has a molecular mass of 78 g mol^{-1}. It is found to have 92.3% carbon and 7.7% hydrogen by mass. Calculate the empirical and molecular formula of X. *[3 marks]*

d) An organic compound contains the following by mass: carbon 36.4% ; hydrogen 6.1% ; fluorine 57.5%.

a) Calculate its empirical formula.

b) The molar mass of the compound is 66 g mol^{-1}. Deduce its molecular formula. *[ASChemistry/page60]*

e) A compound contains C 62.08%, H 10.34% and O 27.58% by mass. Find its empirical formula and its molecular formula given that its relative molecular mass is 58.

f) A compound containing 85.71% C and 14.29% H has a relative molecular mass of 56. Find its molecular formula.

g) A compound containing 84.21% carbon and 15.79% hydrogen by mass has a relative molecular mass of 114. Find its molecular formula.