# Empirical Formula

**Empirical formula** → **relative number of each type of atom in a molecule.**
**Not the molecular formula.**

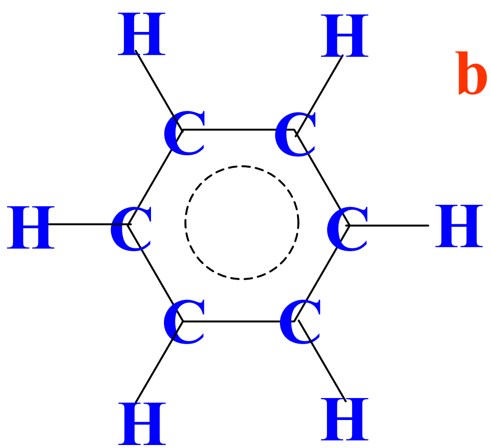
**Empirical formula CH could be for acetylene or benzene.**

**acetylene**

H—C≡C—H

**2C's and 2H's**
**ratio: 1/1**
**empirical formula - CH**

**benzene**




**6C's and 6H's**
**ratio: 1/1**
**empirical formula - CH**

## Empirical formula for hydrocarbons can be determined by combustion.

Hydrocarbon → a molecule with only C's and H's.

**Combustion**

$$\text{Hydrocarbon} + O_2 \longrightarrow CO_2 + H_2O$$

(Note, this is a primary mechanism for the production of the green house gas, carbon dioxide, $CO_2$.)

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**To determine empirical formula:**

Burn hydrocarbon in oxygen; combine completely with $O_2$.

Measure the resulting amount of $CO_2$ and $H_2O$ produced.

**Atoms are conserved in all chemical reactions.**

The C's and H's in the hydrocarbon end up as the C's and H's in the carbon dioxide and water.

## Example

50.000 g of hydrocarbon are burned to give 169.02 g $CO_2$ and 34.593 g $H_2O$.
What are the %s of C and H in the unknown? What is the empirical formula?

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fraction C in $CO_2$ $\quad \dfrac{12.011 \text{ amu}}{44.010 \text{ amu}}$ ← atomic weight C / ← molecular weight $CO_2$ $= 0.2729$ or $27.29\%$

weight $CO_2$ × fraction C → 169.02 g × 0.2729 = $\boxed{46.13 \text{ g C}}$ amount of C in original sample

$$\frac{46.13 \text{ g C}}{50.000 \text{ g sample}} = 0.9226 \longrightarrow \boxed{\% \text{ C} = 92.26 \%}$$

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fraction H in $H_2O$ $\quad \dfrac{2 \times 1.008 \text{ amu}}{18.015 \text{ amu}}$ ← $2 \times$ atomic weight H / ← molecular weight $H_2O$ $= 0.1119$ or $11.19\%$

weight $H_2O$ × fraction H → 34.593 g × 0.1119 = $\boxed{3.871 \text{ g H}}$ amount of H in original sample

$$\frac{3.871 \text{ g H}}{50.000 \text{ g sample}} = 0.07742 \longrightarrow \boxed{\% \text{ H} = 7.742 \%}$$

## Empirical formula

**Determine the number of moles of C and H in the sample.**

$$\frac{46.13\text{ g}}{12.011\text{ g/mol}} = \boxed{3.84\text{ mol}}$$

\# g C in sample → 46.13 g
weight of one mole C → 12.011 g/mol

number of moles C in sample

$$\frac{3.871\text{ g}}{1.008\text{ g/mol}} = \boxed{3.84\text{ mol}}$$

\# g H in sample → 3.871 g
weight of one mole H → 1.008 g/mol

number of moles H in sample

**For each mole of carbon there is one mole of hydrogen.**

**Empirical formula → CH**

**Sample could be acetylene, $C_2H_2$, or benzene, $C_6H_6$.**

**Can't tell from determination of the empirical formula.**

# Chemical Equations

Combustion of benzene, the reaction is

$$\underbrace{C_6H_6 + O_2}_{\text{reactants}} \longrightarrow \underbrace{CO_2 + H_2O}_{\text{products}}$$

Unbalanced Chemical Equation

**Atoms conserved in any chemical reaction!**

Must be same number of each type of atom in reactants and products.

Equation as written

6 C's on left $\longrightarrow$ 1 on right

6 H's on left $\longrightarrow$ 2 on right

2 O's on left $\longrightarrow$ 3 on right

**No good. Doesn't conserve atoms of each kind.**

## Balance the reaction

$$C_6H_6 + 7.5\ O_2 \longrightarrow 6\ CO_2 + 3\ H_2O$$

1. Make same number of C's.
2. Make same number of H's (3×2).
3. Have 2 O's on left, but 15 O's on right
   Multiply $O_2$ by 7.5.

Reaction is now balanced, but sometimes prefer to have all integers.

Multiply each term by 2.

$$2C_6H_6 + 15O_2 \longrightarrow 12CO_2 + 6\ H_2O$$

Not necessary to make all coefficients integers.

Chemical equations tell number of each type of atom involved.

Do not tell how reaction happened.

May be many steps leading to final products.

$C_6H_6$ did not react all at once with 7.5 $O_2$ in combustion of benzene.

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Some chemical equations can provide more information by adding labels in parenthesis.

## Example

Sodium metal reacts with water to form hydrogen and sodium hydroxide.

$$2\,Na + 2\,H_2O \longrightarrow H_2 + 2\,NaOH$$

Not balanced – 2H’s on left, 3H’s on right. Need to balance.

Add labels in parentheses.

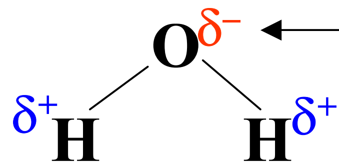
$$\underset{\text{solid}}{2Na(s)} + \underset{\text{liquid}}{2H_2O(l)} \longrightarrow \underset{\text{gas}}{H_2(g)} + \underset{\text{aqueous ion}}{2Na^{+}(aq)} + \underset{\text{aqueous ion}}{2OH^{-}(aq)}$$

# Hydrated ions

**NaOH does not exist in water. It dissolves; forms hydrated ions.**

**$Na^+$ and $OH^-$.**

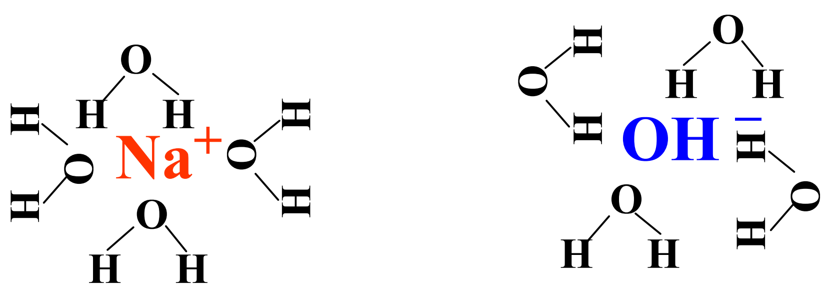
**Water, $H_2O$, bonds partially ionic. Puts partial charge on each atom.**




partial negative charge, delta –

partial positive charge, delta +

## Ions in aqueous solution




**Positive ions surrounded by "negative" side of water.**
**Negative ions surrounded by "positive" side of water.**

**Water is excellent solvent because it can associate with both positive and negative ions and with other molecules that have partial charges.**

## Calculation of reaction yields

**Balanced chemical reaction**

$$xA + yB \longrightarrow zC$$

**x moles A combine with y moles B to give z moles C.**

**Example**

$$C_6H_6 + 71/2 O_2 \longrightarrow 6CO_2 + 3\,H_2O$$

**1 mole benzene combines with 71/2 moles $O_2$ to give 6 moles $CO_2$ and 3 moles $H_2O$.**

**Chemical equation tells the relative amounts of each reacted needed.**

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**However, can start reaction without correct relative amounts as given in equation.**

**Some reactant may be left over when reaction is finished.**

**Can use**

**Balanced equation.**

**Known amounts of starting material.**

**To determine**

**How much of each product is produced.**

**How much starting material is left over.**

**Example:** **Starting with**

10.00 g silver (Ag)

and 1.000 g sulfur (S)

1. How much sliver sulfide ($Ag_2S$) is formed in the reaction

$$2Ag + S \longrightarrow Ag_2S?$$

2. How much starting material is left over?

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The reaction says, 2 moles Ag will combine with 1 mole S, or

2x moles of Ag will combine with x moles of S.

In the problem: Ag $\dfrac{10.00\ \text{g}}{107.868\ \text{g/mol}} = 0.09271$ moles Ag

from atomic weight Ag

S $\dfrac{1.000\ \text{g}}{32.06\ \text{g/mol}} = 0.03119$ moles S

from atomic weight S

Have $3.119 \times 10^{-2}$ mol S.

Need twice this number of moles Ag, $6.238 \times 10^{-2}$ mol Ag to use up all the S.

Have more Ag. Some Ag will be left over.

$3.119 \times 10^{-2}$ mol $Ag_2S$ will be formed.

$3.119 \times 10^{-2}$ mol $Ag_2S$ will be formed.

atomic weight Ag
atomic weight S

**Weight of $Ag_2S$ formed:**

$$3.119 \times 10^{-2} \text{ mol} \times \underbrace{[2(107.868) + 32.06]}_{\text{molecular weight } Ag_2S} \text{ g/mol} = 7.739 \text{ g } Ag_2S$$

number of moles formed

molecular weight $Ag_2S$

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**Amount of Ag left over:**

$$0.09271 \text{ mol} - 2(0.03119) \text{ mol} = 0.03033 \text{ mol Ag}$$

starting amount Ag (mol)

amount Ag used up (mol)

unused

$$0.03033 \text{ mol} \times 107.868 \text{ g/mol} = 3.271\text{g Ag left over}$$

unused Ag (mol)

g/mol Ag

S (sulfur) is the limiting reagent. Determines amount of product.