

Introducing the Mole

Answer the questions below.

1. Define relative atomic mass.

The average mass of atoms of an element compared to the mass of carbon-12.

2. Define relative formula mass.

The sum of the relative atomic masses of the elements in a compound.

3. Calculate the Mr of ammonia (NH_3). N=14, H=1

$$14 + (3 \times 1) = 17$$

4. Calculate the Mr of calcium hydroxide ($\text{Ca}(\text{OH})_2$). Ca=40, O=16, H=1

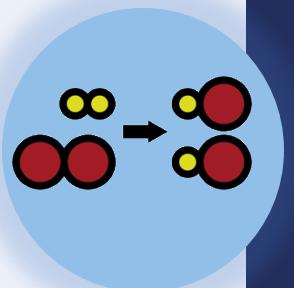
$$40 + 2(16+1) = 74$$

5. Calculate the percentage by mass of hydrogen in ammonia.

$$\% \text{ by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

$$\% \text{ by mass} = \frac{3}{17} \times 100$$

$$\% \text{ by mass} = 17.65 \%$$



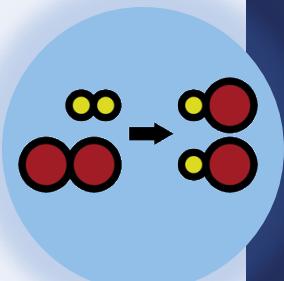
Introducing the Mole

Do Now:

1. Define relative atomic mass.
2. Define relative formula mass.
3. Calculate the Mr of ammonia (NH_3). N=14, H=1
4. Calculate the Mr of calcium hydroxide ($\text{Ca}(\text{OH})_2$). Ca=40, O=16, H=1
5. Calculate the percentage by mass of hydrogen in ammonia.

Drill:

1. Write 58230000 in standard form.
2. Convert 3.07×10^4 from standard form.
3. Write 0.0006 in standard form.

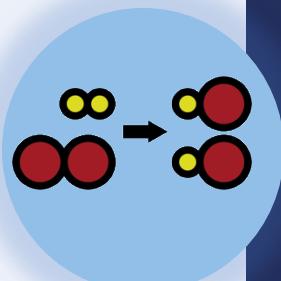


Introducing the Mole

Read Now:

In chemical reactions, there are a huge number (think many, many trillions!) of atoms in reactants that rearrange to form the products. Remember in a chemical reaction, mass is conserved, so no atoms are made or lost during a chemical reaction, only rearranged. So that scientists can keep track of the numbers of atoms involved in different substances without having to write out huge numbers, scientists use a quantity called ‘the mole’. As with many things in chemistry, we use carbon as the reference. 12 g of carbon contains 6.02×10^{23} atoms. This number of atoms is called Avogadro’s number and is the number of atoms in one mole. As a definition, we can say that a mole is the number of atoms contained in 12 g of carbon.

1. State the law of conservation of mass.
2. State Avogadro’s number.
3. Define a mole.
4. State the relative atomic mass of carbon.



(HT) Introducing the Mole

C4.3.2

Science
Mastery

C4.3.1 Prior Knowledge Review

➤ **C4.3.2 (HT) Introducing the Mole**

C4.3.3 (HT) Mole Calculations

C4.3.4 PKR: Concentration

C4.3.5 TIF: Calculating Concentration

C4.3.6 TIF: Calculating an Unknown Concentration

C4.3.7 (HT) Amounts of Substances in Equations

C4.3.8 (HT) Limiting Reactants

C4.3.9 PKR: Reactions of Acids



C4.3.10 Acids, Alkalies and Neutralisation

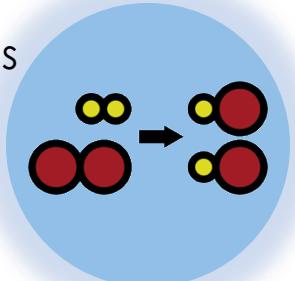
C4.3.11 TIF: Acid-Alkali Titration

C4.3.12 TIF: Acid-Alkali Titration Analysis

C4.3.13 TIF: Titration Calculations

C4.3.14 (HT) Strong and Weak Acids

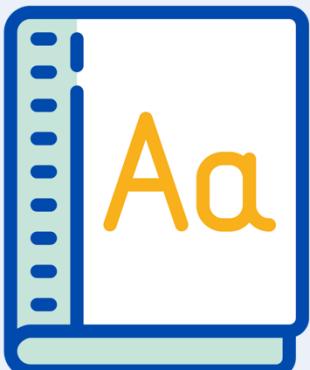
C4.3.15 TIF: Volumes of Gases



Following this lesson, students will be able to:

- State that 1 mole of a substance contains 6.02×10^{23} particles.
- Convert decimal numbers into standard form.
- Calculate the mass of 1 mole of a substance.

Key Words:



mole

mass

amount

relative formula mass

Avogadro's number

This is the fix-it portion of the lesson

The **fix-it** is an opportunity to respond to gaps in knowledge, especially those identified by the **pre-unit quiz**.

- The teacher should customise this slide as needed, to facilitate
 - **reteach, explanation, demonstration or modelling** of ideas and concepts that students have not yet grasped or have misunderstood.
 - **practise** answering specific questions or of key skills.
 - **redrafting** or **improving** previous work.

Answer the questions below.

1. Choose the correct definition of relative formula mass.
 A. The average mass of atoms of an element compared to the mass of carbon-12
 B. The sum of relative atomic masses in a compound
 C. The percentage of a compound is made of a particular element
2. Calculate the relative formula mass of carbon dioxide (CO_2).
 $C = 12, O = 16$
 A. 28
 B. 44
 C. 56
3. Calculate the percentage by mass of oxygen in carbon dioxide.
 A. 27.27%
 B. 36.36%
 C. 72.72 %

Exit ticket

Is it possible to weigh one atom in a lab?



Copper



Sulfur

Symbol	Cu	S
A_r	64	32

Imagine one atom of each element.

What would the ratio of the masses of these atoms be?

64 : 32 or 2 : 1

Is it possible to weigh one atom in our lab?

No, they are far too small. Chemists need to be able to compare masses that can be easily be measured.

Is it possible to weigh one million atoms in a lab?



Copper



Sulfur

Symbol	Cu	S
A_r	64	32
Ratio of masses 1 atom	2	1

What would the ratio of masses be for the following numbers of atoms?

100 atoms	2	1
1000 atoms	2	1
1 000 000 atoms	2	1

Even 1 000 000 atoms cannot be measured out in our lab!
The atoms are just too small.

Is this correct?

64g of Copper contains
the same number of atoms
as 32 g of sulfur.

Yes! There are the same
number of atoms in 64 g
of Cu as in 32 g of S.



Copper



Sulfur

Symbol	Cu	S
A_r	64	32
Ratio of masses	2	1
Mass in grams	64 g	32 g

How small are atoms?

Atoms are **extremely small**.

63.5
Cu
copper
29



Copper

The number of atoms present in a lump of copper is **extremely large**.

The relative atomic mass (63.5 for copper) tells you the mass of copper (in grams) that contains exactly 602,000,000,000,000,000 atoms!

This is the same for every element! It's relative atomic mass tells you the mass of that element that contains exactly 602,000,000,000,000,000 atoms.

12 g of carbon contains 602,000,000,000,000,000,000 atoms.

16 g of oxygen contains 602,000,000,000,000,000,000 atoms.

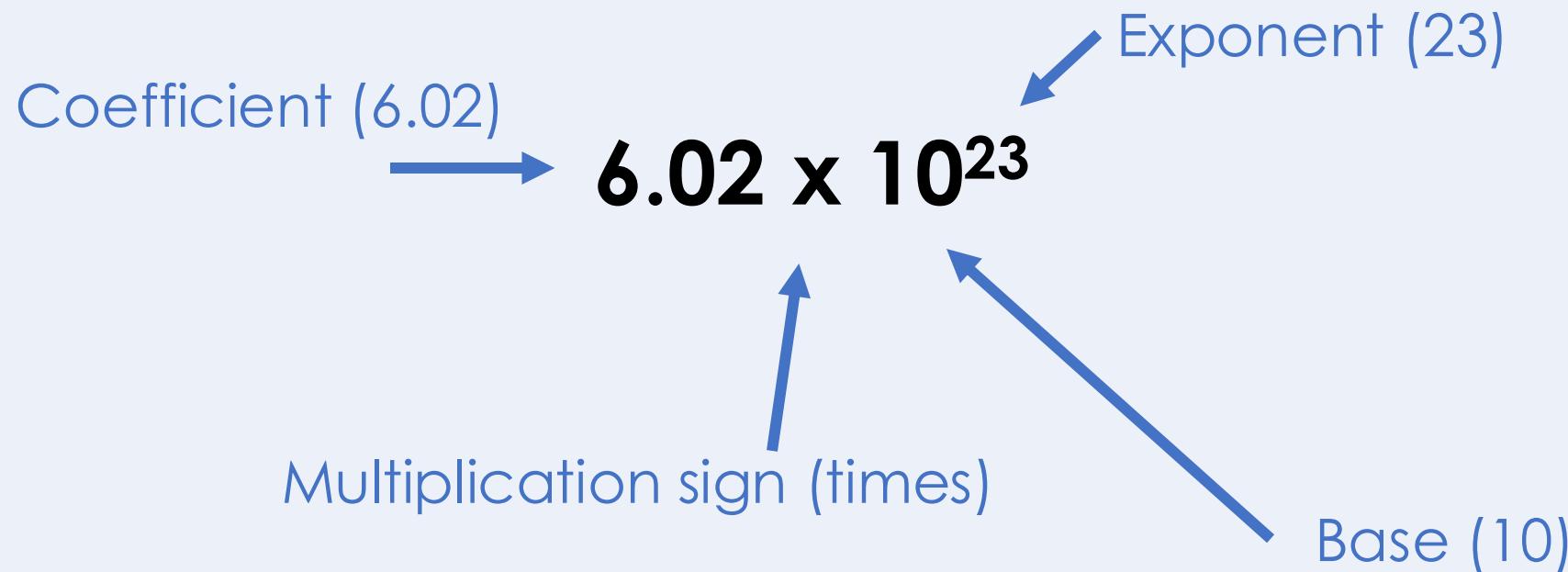
1 g of Hydrogen contains 602,000,000,000,000,000,000 atoms.

Standard Form

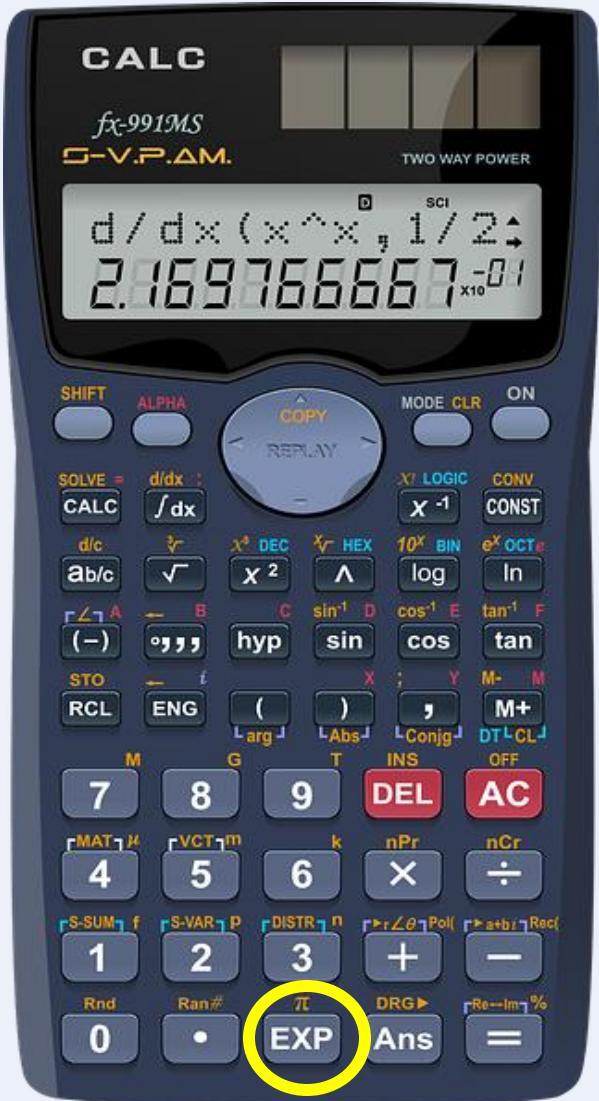
Numbers such as 602,000,000,000,000,000,000 can be written in **standard form** to make them easier to work with.

602,000,000,000,000,000,000 can be written as **6.02×10^{23}**

All numbers in standard form look like this:



Writing standard form into a scientific calculator



For 6.02×10^{23} , type 6.02, then press the button in the picture, then press 23.

Converting a number to standard form

Move the decimal point along until there is just one digit to the left of the decimal point. Start with the **number $\times 10^0$**

If the decimal point moves **left**
The exponent goes **up**

If the decimal point moves **right**
The exponent goes **down**

Examples:

157030.00

1.5703×10^5

849000000

8.49×10^8

0.00000007

7×10^{-9}

Question 1: Which is bigger?

They are both the same size! They both equal 23800000000

✓ 2.38 x 10¹⁰

or

✗ 23.8 x 10⁹

Remember:

Standard form is written in the form of $a \times 10^n$, where a is a number bigger than or equal to 1 and less than 10.

Question 2: Which number shows the correct way to write standard form?

The first number is correct, because there should always be just one number before the decimal point when writing standard form.

Avogadro's number

We can write the number

602,000,000,000,000,000,000 as **6.02×10^{23}**

This number is called Avogadro's number or the
Avogadro constant.

We call this number and the physical amount of this material **1 mole** (unit= mol).

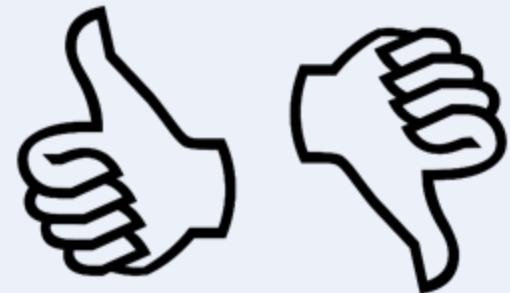
1 mole = 6.02×10^{23} particles in the same way that 1 dozen = 12.

We can say that in 63.5g of Cu there are **6.02×10^{23}** atoms OR 63.5g is 1 mole of Cu.



True or false?

1. One mole of carbon contains the same number of atoms as one mole of copper.
2. 1 gram of carbon contains the same number of atoms as 1 gram of copper.
3. Atoms are very small.
4. 6 g of carbon contains the same number of atoms as 29 g of copper.
5. The symbol for the unit mole is mol.
6. 1 mole of carbon contains 6.02×10^{23} atoms.
7. The number of atoms present in an object are very small.
8. 12 g of carbon contain the same number of atoms as 63.5 g of copper.



True

False

12	C	carbon	6

63.5	Cu	copper	29

How many moles?

The relative formula mass of sulfur is 32.

32
S
sulfur
16

How many atoms would there be in 32 g of sulfur?

How many moles are there in 32 g of sulfur?

How many atoms would there be in 64 g of sulfur?

How many moles would there be in 64 g of sulfur?

How many atoms?

How many atoms would there be in 12 g of carbon?

What is the relative formula mass of carbon dioxide?

How many molecules of carbon dioxide would there be in 44 g?

How many atoms would there be in 44 g of carbon dioxide?

12
C
carbon
6

16
O
oxygen
8

Drill

1. State the relative atomic mass of carbon.
2. State Avogadro's number.
3. State how many atoms would be in 12 g of carbon.
4. State how many moles would be in 12 g of carbon.
5. State the relative atomic mass of calcium.
6. State how many atoms would be in 40 g of calcium.
7. State how many moles would be in 40 g of calcium.

Drill answers

1. 12
2. 6.02×10^{23}
3. 6.02×10^{23}
4. 1 mole
5. 40
6. 6.02×10^{23}
7. 1 mole

Check for understanding

Answer the questions below.

1. Which of the following is **true**?

- A. The symbol for mole is mol.
- B. The symbol for mole is moles.
- C. The symbol for moles is m.

2. How many molecules in 1 mole of CO_2 ?

- A. 44
- B. 6.02×10^{23}
- C. 1

3. Which number is the same as 6×10^3 ?

- A. 60300
- B. 63000
- C. 6000

Lesson C4.3.2

What was good about this lesson?

What can we do to improve this lesson?

[Send us your feedback by clicking this link](#)
or by emailing sciencemastery@arkonline.org
Thank you!