Chapter 5 Practice Problems

Counting Atoms and Molecules: The Mole

Solutions for Practice Problems

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1. **Problem**

   The two stable isotopes of boron exist in the following proportions: 19.78% $^{10}_{5}B$ (10.01 u) and 80.22% $^{11}_{2}B$ (1.01 u). Calculate the average atomic mass of boron.

   What is Required?

   You need to find the average atomic mass of boron.

   What is Given?

   You are given the relative abundance (proportions) and the atomic mass of each stable isotope.

   Plan Your Strategy

   Express each isotopic proportion as a decimal and multiply this by its atomic mass. The average mass of boron is the sum of these products.

   Act on Your Strategy

   $$\begin{aligned}\text{Average atomic mass of B} &= 10.01\text{ u }(0.1978) + 11.01\text{ u }(0.8022)\\ &= 10.81\text{ u}\end{aligned}$$

   Check Your Solution

   In this case, the greater proportion of isotope was that of 11.01 u. Therefore, the atomic mass of the naturally occurring boron should be closer to this value than to 10.01 u. The calculated atomic mass matches this reasoning. A check in the periodic table confirms the value.

2. **Problem**

   In nature, silicon is composed of three isotopes. These isotopes (with their isotopic abundances and atomic masses) are $^{28}_{14}Si$ (92.235, 27.98 u), $^{29}_{14}Si$ (4.67%, 28.97 u), and $^{30}_{14}Si$ (3.10%, 29.97 u). Calculate the average mass of silicon.

   What is Required?

   You need to find the atomic mass of silicon.

   What is Given?

   You are given the relative abundance and the atomic mass of each isotope.

   Plan Your Strategy

   Express each isotopic proportion as a decimal and multiply this by its atomic mass. The average mass of silicon is the sum of these products.

   Act on Your Strategy

   $$\begin{aligned}\text{Average atomic mass of Si} &= 27.98\text{ u }(0.9223) + 28.97\text{ u }(0.0467)\\ &\quad + 29.97\text{ u }(0.0310)\\ &= 28.09\text{ u}\end{aligned}$$

Check Your Solution
In this case, the greatest proportion of isotope was that of 27.98 u. Therefore, the atomic mass of the naturally occurring silicon should be closest to this value than to 28.97 u or 29.97 u. The calculated atomic mass matches this reasoning and a check in the periodic table confirms the value.

3. **Problem**
Copper is a corrosion-resistant metal that is used extensively in plumbing and wiring. Copper exists as two naturally occurring isotopes: ${}^{63}_{29}Cu$ (62.93 u) and ${}^{65}_{29}Cu$ (64.93 u). These isotopes have isotopic abundances of 69.1% and 30.9%, respectively. Calculate the average atomic mass of copper.

What is Required?
You need to find the atomic mass of copper.

What is Given?
You are given the relative abundance and the atomic mass of each isotope.

Plan Your Strategy
Express each isotopic proportion as a decimal and multiply this by its atomic mass. The average mass of copper is the sum of these products.

Act on Your Strategy
$$\text{Average atomic mass of Cu} = 62.93\text{ u }(0.691) + 64.93\text{ u }(0.309)$$
$$= 63.55\text{ u}$$

Check Your Solution
In this case, the greater proportion of isotope was that of 62.93 u. Therefore, the atomic mass of the naturally occurring copper should be closer to this lower value than to 64.93 u. The mean of the two isotopic atomic mass units is 63.93 u. So, the calculated atomic mass of copper, which is slightly lower than this mean value is a reasonable answer. A check in the periodic table confirms the atomic mass calculated.

4. **Problem**
Lead occurs naturally as four isotopes. These isotopes (with their isotopic abundances and atomic masses) are ${}^{204}_{82}Pb$ (1.37%, 204.0 u), ${}^{206}_{82}Pb$ (26.26%, 206.0 u), ${}^{207}_{82}Pb$ (20.82%, 207.0 u), and ${}^{208}_{82}Pb$ (51.55%, 208.0 u). Calculate the average atomic mass of lead.

What is Required?
You need to find the atomic mass of lead.

What is Given?
You are given the relative abundance and the atomic mass of each isotope.

Plan Your Strategy
Express each isotopic proportion as a decimal and multiply this by its atomic mass. The average mass of lead is the sum of these products.

Act on Your Strategy
Average atomic mass of
$$\text{Pb} = 204\text{ u }(0.0137) + 206\text{ u }(0.2626) + 207\text{ u }(0.2082) + 208\text{ u }(0.5155)$$
$$= 207.2\text{ u}$$

Check Your Solution
In this case, although the greatest proportion of isotope was that of 208 u, it contributed to only half of the atomic mass, and was therefore equally counterbalanced by the combined atomic masses of the other three isotopes. So, one would expect the atomic mass of lead to be in the region of between 206-208 u (since the 204 u

contribution is almost negligible). The calculated atomic mass of 207.2 is reasonable. A check in the periodic table confirms the atomic mass calculated.

Solutions to Practice Problems

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1. Problem

Hydrogen is found primarily as two isotopes in nature: ${}^{1}_{1}H$ (1.0078 u) and ${}^{2}_{1}H$ (2.0140 u). Calculate the percentage abundance of each isotope based on hydrogen's average atomic mass.

What is Required?

You need to find the isotopic abundance of hydrogen.

What is Given?

The atomic mass of each isotope is given. From the periodic, the average atomic mass of hydrogen is 1.01 u.

Plan Your Strategy

First express the abundance of each isotope as a decimal instead of a percentage, giving a total abundance of both isotopes as 1. Let the abundance of hydrogen-1 be x and hydrogen-2 be $1 - x$. Set up the equation and solve for x. Then multiply both decimal answers by 100% to express it as percentage abundance.

Act on Your Strategy

$$\begin{aligned}
\text{Average atomic mass} &= x(\text{mass hydrogen-1}) + (1 - x)(\text{mass hydrogen-2})\\
1.01 &= x(1.0078) + (1 - x)(2.0140)\\
1.01 &= 1.0078x + 2.0140 - 2.0140x\\
1.0062x &= 1.004\\
x &= 0.9978
\end{aligned}$$

The percentage abundance of hydrogen-1 is $0.9978 \times 100\% = 99.78\%$.
The percentage abundance of hydrogen-2 $= (1 - 0.9978) \times 100\% = 0.022\%$.

Check Your Solution

Since the atomic mass of hydrogen-1 is very close to the average atomic mass of hydrogen, one would expect its percentage abundance to be very high. The 99.78% result confirms this reasoning.

6. Problem

Lanthanum is composed of two isotopes: ${}^{138}_{57}La$ (137.91 u) and ${}^{139}_{57}La$ (138.91 u). Look at the periodic table. What can you say about the abundance of ${}^{138}_{57}La$?

What is Required?

You are asked to predict the abundance of the ${}^{138}_{57}La$.

What is Given?

The atomic mass of each isotope is given. From the periodic, the average atomic mass of lanthanum is 138.91 u.

Plan Your Strategy

Compare the atomic masses of each isotope to the average atomic mass from the periodic table. The closer the value of La-138 or La-139 is to the average atomic mass, the greater is its abundance. The further the value of La-138 or La-139 is from the average atomic mass, the lower is the abundance of that particular isotope.

Act on Your Strategy
Because the Periodic Table gives the average atomic mass of lanthanum as 138.91 u, the same as the atomic mass of La-139 to five significant digits, you know that the abundance of La-138 must be very low.

Check Your Solution
Calculate the abundances of the isotopes, by expressing La-138 as x and La-139 as $1 - x$, and solving for x.

$$\begin{aligned}\text{Average atomic mass} &= x(\text{mass La-138}) + (1 - x)(\text{mass La-139})\\ 138.91 &= x(137.91) + (1 - x)(138.91)\\ 138.91 &= 137.91x + 138.91 - 138.91x\\ 0 &= 137.91x - 138.91x\end{aligned}$$

From this calculation, x is zero, implying there is no naturally occurring La-138 and that lanthanum is 100% La-139. In fact, La-138 does exist naturally, at an almost negligible percentage of 0.0902%, while the abundance of La-139 is 99.9098%. The reason these values were not calculated here is because we only took the average atomic mass to 5 significant digits. To seven significant digits, the average atomic mass of La is 138.9055 u, which would have yielded results closer to computed values.

7. Problem
Rubidium ignites spontaneously when exposed to oxygen to form rubidium oxide, RbO_2. Rubidium exists as two isotopes: ${}^{85}_{37}Rb$ (84.91 u) and ${}^{87}_{37}Rb$ (86.91 u). If the average atomic mass of rubidium is 85.47 u, determine the percentage abundance of ${}^{85}_{37}Rb$.

What is Required?
You have to find the percentage abundance of ${}^{85}_{37}Rb$.

What is Given?
The atomic mass of each isotope is given, as well as the average atomic mass of rubidium.

Plan Your Strategy
First express the abundance of each isotope as a decimal instead of a percentage, giving a total abundance of both isotopes as 1. Let the abundance of Rb-85 be x and Rb-87 be $1 - x$. Set up the equation and solve for x. Then multiply x by 100% to express it as percentage abundance.

Act on Your Strategy

$$\begin{aligned}\text{Average atomic mass} &= x(\text{mass Rb-85}) + (1 - x)(\text{mass Rb-87})\\ 85.47\text{ u} &= x(84.91) + (1 - x)(86.91)\\ 84.57 &= 84.91x + 86.91 - 86.91x\\ 2x &= 1.44\\ x &= 0.72\end{aligned}$$

Therefore, the percentage abundance of Rb-85 is $0.72 \times 100\% = 72\%$.

Check Your Solution
The mean of the atomic masses of the two isotopes is 85.91 u. The average atomic mass of rubidium (85.47 u) is slightly less than this mean, and closer to the value of Rb-85 (84.91 u). So you would expect it to comprise of a greater abundance of Rb-85, which matches the result.

8. Problem
Oxygen is composed of three isotopes: ${}^{16}_{8}O$ (15.995 u), ${}^{17}_{8}O$ (16.999 u), and ${}^{18}_{8}O$ (17.999 u). One of these isotopes, ${}^{17}_{8}O$, comprises 0.037% of oxygen. Calculate the

percentage abundance of the other two isotopes, using the average atomic mass of 15.9994 u.

What is Required?
You need to find the percentage abundances of ${}^{16}{}_{8}O$ and ${}^{18}{}_{8}O$.

What is Given?
The atomic mass of each isotope is given, as well as the average atomic mass of oxygen. The percentage abundance of ${}^{17}{}_{8}O$ is given as 0.037%.

Plan Your Strategy
For a simpler calculation, express the abundance of each isotope as a decimal instead of a percentage, giving a total abundance of all isotopes as 1. Let the abundance of oxygen-16 be x. The abundance of oxygen-17 is 0.00037. Then the abundance of oxygen-18 would be $1 - 0.00037 + x$. Set up the equation and solve for x. Then multiply the calculated values by 100% to express as percentage abundances.

Act on Your Strategy

$$\text{Average atomic mass} = x(\text{mass O-16}) + 0.00037(\text{mass O-17}) + (1 - 0.00037 + x)(\text{mass O-18})$$

$$16.00 = x(15.9949) + (0.00037)(16.9991) + (0.99963 - x)(17.9992)$$

$$16.00 = 15.9949x + 0.006289 + 17.99254 - 7.9992x$$

$$2.0043x = 1.998829$$

$$\text{so } x = 0.9972$$

The percentage abundance of O-16 is $0.9972 \times 100\% = 99.72\%$.
The percentage abundance of O-17 is 0.037% (given).
The percentage abundance of O-18 is $(1 - 0.9972 + 0.00037) \times 100\% = 0.243\%$.

Check Your Solution
From the given atomic masses of the three isotopes of oxygen, it is very clear that O-16 has the closest atomic mass to the average atomic mass of naturally occurring oxygen. Its abundance would be expected to be very high, as was calculated to be so.

Solutions for Practice Problems

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9. Problem
The length of British Columbia's coastline is 17 856 km. If you laid 6.02×10^{23} metre sticks end to end along the coast of BC, how many rows of metre sticks would you have?

What is Required?
You need to find the number of rows that 6.02×10^{23} metre sticks would generate for a length as long as BC's coastline.

What is Given?
The length of the coastline is 17 856 km. The number of metre sticks is 6.02×10^{23}.

Plan Your Strategy
First convert the km to metres since you are working with metre sticks. The total length of all the metre sticks end to end is 6.02×10^{23} times 1 m. The number of rows of these sticks along BC's coastline would therefore be the total length of the sticks divided by the length of the coastline.

Act on Your Strategy
Length of coastline in metres $= 1.7856 \times 10^4 \text{ km} \times 10^3 = 1.7856 \times 10^7$ m
Total length of the sticks $= (6.02 \times 10^{23})$ sticks $\times 1 \text{ m} = 6.02 \times 10^{23}$ m

Number of rows = Total length of sticks / Length of coastline

$$= \frac{6.02 \times 10^{23}\ \cancel{m}}{1.7856 \times 10^{7}\ \cancel{m}}$$

$$= 3.37 \times 10^{16}$$

Check Your Solution
Looking at the magnitude of the numbers, $10^7 \times 10^{16}$ accounts for the 10^{23} of the metre sticks. Hence the magnitude of the answer is correct.

10. Problem
The area of Nunavut is 1 936 113 km^2. Suppose that you had 6.02×10^{23} sheets of pastry, each with the dimensions 30 cm × 30 cm. How many times could you cover Nunavut completely with the pastry?

What is Required?
You need to find the number of times that 6.02×10^{23} sheets of a given dimension of pastry would cover the area of Nunavut.

Plan Your Strategy
First convert the km^2 to cm^2 since you are working with the pastry in cm^2. Multiply the pastry dimension by 6.02×10^{23} to get the total area of pastry. The number of times this covers Nunavut is the total area of the pastry divided by the area of Nunavut.

Act on Your Strategy
Area of Nunavut $= (1.936113 \times 10^6)$ km$^2 \times (10^5)^2$ cm^2 $= 1.936\ 10^{16}$ cm^2
Total area of pastry $= (6.02 \times 10^{23})(30)^2$ cm^2 $= 5.418 \times 10^{26}$ cm^2

$$\text{Number of layers of pastry} = \frac{\text{Total area of pastry}}{\text{Area of Nunavut}}$$

$$= \frac{5.418 \times 10^{26}}{1.936 \times 10^{16}}$$

$$= 2.8 \times 10^{10}$$

Check Your Solution
Looking at the magnitude of the numbers, $\frac{10^{26}}{10^{16}}$ accounts for the 10^{10} of the result. Hence the magnitude of the answer is correct.

11. Problem
If you drove for 6.02×10^{23} days at a speed of 100 km/h, how far would you travel?

What is Required?
You need to find the distance travelled at a speed of 100 km/h over 6.02×10^{23} days.

What is Given?
The number of days of travel is 6.02×10^{23} days. The speed of travel is 100 km/h. You assume there are no stops along the way.

Plan Your Strategy
For every hour, 100 km is traversed. Therefore, you have to determine how many hours there are in 6.02×10^{23} days. There are 24 hours in a day, therefore multiply 6.02×10^{23} days by 24. Multiply this answer by 100 km for the total distance travelled.

Act on Your Strategy
Total number of hours $= 6.02 \times 10^{23}$ days $\times$ 24 hours/day $= 1.44 \times 10^{25}$ hours
Total distance travelled $= 1.44 \times 10^{25}$ hours $\times$ 100 km $= 1.44 \times 10^{27}$ km

Check Your Solution
Work backwards. The total distance divided by the total number of hours travelled should give the speed of travel. $\frac{1.44 \times 10^{27} \text{ km}}{1.44 \times 10^{25} \text{ hours}} = 100$ km/h, as given in the question.

12. Problem
If you spent 6.02×10^{23} at a rate of \$1.00/s, how long, in years, would the money last? Assume that every year has 365 days.

What is Required?
You need to find the number of years for which 6.02×10^{23} would last, at a spending rate of \$1.00/s.

What is Given?
The amount to be spent is 6.02×10^{23}. The rate of spending is \$1.00/s, assuming nonstop spending. You assume there are 365 days in every year.

Plan Your Strategy
Since the answer is to be given in years, first convert the spending rate to that of per year. This is done via multiplying the rate by 3600 s/h multiplied by 24 h/day multiplied by 365 days/year. Divide this yearly rate into the full amount to obtain the number of years this amount will last.

Act on Your Strategy

$$\begin{aligned}\text{Spend rate per year} &= (\$1.00\ \cancel{\text{s}}) \times (3600\ \cancel{\text{s/h}}) \times (24\ \cancel{\text{h/day}}) \times (365\ \cancel{\text{day}}/\text{yr}) \\ &= \$3.1536 \times 10^{7}/\text{yr}\end{aligned}$$

$$\begin{aligned}\text{The number of years the amount will last} &= \frac{\$6.02 \times 10^{23}}{(\$3.1536 \times 10^{7}/\text{yr})} \\ &= 1.91 \times 10^{16}\ \text{yr}\end{aligned}$$

Check Your Solution
Looking at the magnitude of the numbers, $\frac{10^{23}}{10^{7}}$ accounts for the 10^{16} of the result. Hence the magnitude of the answer is correct.

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13. Problem
A small pin contains 0.0178 mol of iron, Fe. How many atoms of iron are in the pin?

What is Required?
You need to find the number of atoms of iron in the pin.

What is Given?
The pin contains 0.017 mol of Fe. Avogadro's number, $N_A = 6.02 \times 10^{23}$ atoms/mol.

Plan Your Strategy
The number of atoms would be Avogadro's number multiplied by the number of mol of Fe in the pin.

Act on Your Strategy

$$\begin{aligned}\text{The number of atoms of Fe} &= 0.0178\ \cancel{\text{mol}} \times (6.02 \times 10^{23}\ \text{atoms}/\cancel{\text{mol}}) \\ &= 1.07 \times 10^{22}\ \text{atoms}\end{aligned}$$

Therefore, there are 1.07×10^{22} iron atoms in the pin.

Check Your Solution
Work backwards. Dividing the number of atoms by Avogadro's number should give you the number of mol of Fe.

14. **Problem**
A sample contains 4.70×10^{-4} mol of gold, Au. How many atoms of gold are there in the sample?

What is Required?
You need to find the number of atoms of gold in the sample.

What is Given?
The sample contains 4.70×10^{-4} mol of Au. Avogadro's number, $N_A = 6.02 \times 10^{23}$ atoms/mol.

Plan Your Strategy
The number of atoms would be Avogadro's number multiplied by the number of mol of gold in the sample.

Act on Your Strategy
$$\begin{aligned}\text{The number of atoms of Au} &= 4.70 \times 10^{-4}\ \cancel{\text{mol}} \times (6.02 \times 10^{23}\ \text{atoms}/\cancel{\text{mol}}) \\ &= 2.83 \times 10^{20}\ \text{atoms}\end{aligned}$$
Therefore, there are 2.83×10^{20} gold atoms in the sample.

Check Your Solution
Work backwards. Dividing the number of atoms by Avogadro's number should give you the number of mol of Au.

15. **Problem**
How many formula units are contained in 0.21 mol of magnesium nitrate, $Mg(NO_3)_2$?

What is Required?
You need to find the number of formula units in the magnesium nitrate.

What is Given?
The sample is 0.21 mol. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy
The number of formula units would be Avogadro's number multiplied by the number of mol of $Mg(NO_3)_2$.

Act on Your Strategy
The number of formula units of
$$\begin{aligned}Mg(NO_3)_2 &= 0.21\ \cancel{\text{mol}} \times (6.02 \times 10^{23}\ \text{formula units}/\cancel{\text{mol}}) \\ &= 1.3 \times 10^{23}\ \text{formula units}\end{aligned}$$
Therefore, there are 1.3×10^{23} formula units of $Mg(NO_3)_2$ in the sample.

Check Your Solution
Work backwards. Dividing the number of formula units by Avogadro's number should give you the number of mol of $Mg(NO_3)_2$.

Problem
A litre of water contains 55.6 mol of water. How many molecules of water are in this sample?

What is Required?
You need to find the number of molecules of water in the 1 litre volume.

What is Given?
There is 55.6 mol of water in the 1 litre volume. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy
The number of molecules would be Avogadro's number multiplied by the number of mol of water.

Act on Your Strategy
The number of molecules of H_2O $= 55.6 \cancel{\text{mol}} \times (6.02 \times 10^{23} \text{ molecules}/\cancel{\text{mol}})$
$= 3.35 \times 10^{25}$ molecules
Therefore, there are 3.35×10^{25} water molecules in one litre of water.

Check Your Solution
Work backwards. Dividing the number of molecules by Avogadro's number should give you the number of mol of water.

17. Problem
Ethyl acetate, $C_4H_8O_2$, is frequently used in nail polish remover. A typical bottle of nail polish remover contains about 2.5 mol of ethyl acetate.
(a) How many molecules are in the bottle of nail polish remover?
(b) How many atoms are in the bottle?
(c) How many carbon atoms are in the bottle?

What is Required?
(a) You need to find the number of molecules in the sample.
(b) You need to find the number of atoms in the bottle.
(c) You need to find the number of carbon atoms in the bottle.

What is Given?
The sample consists of 2.5 mol. The molecular formula of ethyl acetate is $C_4H_8O_2$. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy
(a) The number of molecules would be Avogadro's number multiplied by the number of mol of ethyl acetate.
(b) The number of atoms per molecule, from the molecular formula of ethyl acetate, is 14. The total number of atoms in the bottle would be 14 multiplied by the number of molecules.
(c) The number of carbon atoms in the molecular formula of ethyl acetate is 4. The total number of atoms in the bottle would be 4 multiplied by the number of molecules.

Act on Your Strategy
(a) Number of molecules of $C_4H_8O_2$ $= 2.5 \cancel{\text{mol}} \times (6.02 \times 10^{23} \text{ molecules}/\cancel{\text{mol}})$
$= 1.5 \times 10^{24}$ molecules
Therefore, there are 1.5×10^{24} molecules in the bottle.
(b) The number of atoms $= 14 \text{ atoms}/\cancel{\text{molecule}} \times (1.5 \times 10^{24} \cancel{\text{molecules}})$
$= 2.1 \times 10^{25}$ atoms
Therefore, there are 2.1×10^{25} atoms in the bottle.
(c) The number of C atoms $= 4 \text{ atoms}/\cancel{\text{molecule}} \times (1.5 \times 10^{24} \cancel{\text{molecules}})$
$= 6.0 \times 10^{24}$ atoms
Therefore, there are 6.0×10^{24} atoms of carbon in the bottle.

Check Your Solution
(a) Work backwards. Dividing the number of molecules by Avogadro's number should give you the number of mol of ethyl acetate.

(b) Work backwards. Dividing the number of atoms by the number of molecules in the bottle should give you the number of atoms in each molecule.

(c) Work backwards. Dividing the number of carbon atoms in the bottle by the number of molecules in the bottle should give you the number of carbon atoms in each molecule.

18. Problem

Consider a 0.829 mol sample of sodium sulfate, Na_2SO_4.

(a) How many formula units are in the sample?

(b) How many sodium ions, Na^+, are in the sample?

What is Required?

(a) You need to find the number of formula units in the sample.

(b) You need to find the number of sodium ions in the bottle.

What is Given?

The sample consists of 0.829 mol. The molecular formula of sodium sulfate is Na_2SO_4. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

(a) The number of formula units would be Avogadro's number multiplied by the number of mol of sodium sulfate.

(b) There are 2 sodium ions per formula unit of sodium sulfate. The total number of sodium ions would be 2 multiplied by the number of formula units in the sample.

Act on Your Strategy

(a) Number of formula units $= 0.829\ \cancel{\text{mol}} \times (6.02 \times 10^{23}\ \text{formula units}/\cancel{\text{mol}})$
$= 4.99 \times 10^{23}$ formula units

Therefore, there are 4.99×10^{23} formula units in the sample.

(b) The number of sodium ions $= 2\ \text{ions}/\cancel{\text{formula unit}} \times (4.99 \times 10^{23}\ \cancel{\text{formula units}})$
$= 9.98 \times 10^{23}$ sodium ions.

Therefore, there are 9.98×10^{23} sodium ions in the sample.

Check Your Solution

(a) Work backwards. Dividing the number of formula units by Avogadro's number should give you the number of mol of sodium sulfate.

(b) Work backwards. Dividing the number of sodium ions by the number of formula units in the sample should give you the number of ions per formula unit.

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19. Problem

A sample of bauxite ore contains 7.71×10^{24} molecules of aluminum oxide, Al_2O_3. How many moles of aluminum oxide are in the sample?

What is Required?

You need to find the number of moles in 7.71×10^{24} molecules of aluminum oxide.

What is Given?

The number of molecules is 7.71×10^{24}. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy

Divide the number of molecules by Avogadro's number to get the number of mol of aluminum oxide.

Act on Your Strategy
Number of moles $= \dfrac{7.71 \times 10^{24} \cancel{\text{molecules}}}{6.02 \times 10^{23}/ \cancel{\text{molecules}}/\text{mol}}$
$= 12.8$ mol
Therefore, there are 12.8 moles of aluminum oxide in the sample.

Check Your Solution
Work backwards. Dividing the number of molecules by the number of moles should give you Avogadro's number.

20. Problem
A vat of cleaning solution contains 8.03×10^{26} molecules of ammonia, NH_3. How many moles of ammonia are in the vat?

What is Required?
You need to find the number of moles in 8.03×10^{26} molecules of ammonia.

What is Given?
The number of molecules is 8.03×10^{26}. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy
Divide the number of molecules by Avogadro's number to get the number of mol of ammonia.

Act on Your Strategy
Number of moles $= \dfrac{8.03 \times 10^{26} \cancel{\text{molecules}}}{(6.02 \times 10^{23} \cancel{\text{molecules}}/\text{mol})}$
$= 1.33 \times 10^{3}$ mol
Therefore, there are 1.33×10^{3} moles of ammonia in the vat.

Check Your Solution
Work backwards. Dividing the number of molecules by the number of moles should give you Avogadro's number.

21. Problem
A sample of cyanic acid, HCN, contains 3.33×10^{22} atoms. How many moles of cyanic acid are in the sample? **Hint:** Find the number of molecules of HCN first.

What is Required?
You need to find the number of moles in 3.33×10^{22} atoms of HCN.

What is Given?
The number of atoms is 3.33×10^{22}. The formula of cyanic acid is given. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy
From the formula, you know there are 3 atoms in every molecule of cyanic acid. The number of molecules of cyanic acid is the number of atoms divided by 3. Then, divide the number of molecules by Avogadro's number to get the number of mol of HCN.

Act on Your Strategy
Number of molecules of HCN $= \dfrac{3.33 \times 10^{22} \cancel{\text{atoms}}}{3 \cancel{\text{atoms}}/\text{molecule}}$
$= 1.11 \times 10^{22}$ molecules
Number of moles of HCN $= \dfrac{1.11 \times 10^{22} \cancel{\text{molecules}}}{(6.02 \times 10^{23} \cancel{\text{molecules}}/\text{mol})}$
$= 1.84 \times 10^{-2}$ mol
Therefore, there are 1.84×10^{-2} mol or 0.0184 mol of cyanic acid in the sample.

Check Your Solution
Work backwards. Dividing the number of molecules by the number of moles should give you Avogadro's number.

22. Problem
A sample of pure acetic acid, CH_3COOH, contains 1.40×10^{23} carbon atoms. How many moles of acetic acid are in the sample?

What is Required?
You need to find the number of moles of acetic acid that carries 1.40×10^{23} carbon atoms.

What is Given?
The number of carbon atoms is 1.40×10^{23}. The molecular formula of acetic acid is CH_3COOH. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy
From the formula, you know there are 2 carbon atoms in every molecule of acetic acid. The number of molecules of acetic acid is the number of carbon atoms divided by 2. Then, divide the number of molecules by Avogadro's number to get the number of mol of acetic acid.

Act on Your Strategy

$$\text{Number of molecules of Acetic acid} = \frac{1.40 \times 10^{23} \cancel{\text{C-atoms}}}{2/\cancel{\text{C-atoms}}/\text{molecule}}$$
$$= 7.0 \times 10^{22} \text{ molecules}$$
$$\text{Number of moles of Acetic acid} = \frac{7.0 \times 10^{22} \cancel{\text{molecules}}}{(6.02 \times 10^{23}/\cancel{\text{molecules}}/\text{mol})}$$
$$= 1.16 \times 10^{-1} \text{ mol}$$

Therefore, there are 1.16×10^{-1} mol or 0.116 mol of acetic acid in the sample.

Check Your Solution
Work backwards. Dividing the number of molecules by the number of moles should give you Avogadro's number.

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23. Problem
State the molar mass of each element.
(a) xenon, Xe
(b) osmium, Os
(c) barium, Ba
(d) tellurium, Te

What is Required?
You have to find the molar mass of the elements listed.

What is Given?
The chemical symbols and chemical names are given.

Plan Your Strategy
With no other information given, the only source of the answer is a Periodic Table. The average atomic mass of the element is numerically equivalent to its molar mass.

Act on Your Strategy
From the average atomic mass in the Periodic Table, the numerically equivalent molar masses are:

(a) Xe: 131.29 g/mol.
(b) Os: 190.23 g/mol.
(c) Ba: 137.33 g/mol.
(d) Te: 127.60 g/mol.

Check Your Solution
Periodic Table values can vary in some publications, depending on the number of significant digits recorded. Check with other books or online sources to verify the most commonly accepted atomic mass units (and hence the molar mass) of the elements.

24. Problem
Find the molar mass of each compound.
(a) ammonia, NH_3
(b) glucose, $C_6H_{12}O_6$
(c) potassium dichromate, $K_2Cr_2O_7$
(d) iron(III) sulfate, $Fe_2(SO_4)_3$

What is Required?
You have to find the molar masses of the compounds listed.

What is Given?
The chemical formulas of the compounds are given.

Plan Your Strategy
The molar mass of a compound is the sum of all the molar masses of its component elements. Find the atomic mass (which is numerically equivalent to the molar mass) of the elements from the Periodic Table and add them together.

Act on Your Strategy
The molar masses of the compounds are:
(a) NH_3: $14.01\text{ g/mol} + 3(1.01\text{ g/mol}) = 17.04\text{ g/mol}$.
(b) $C_6H_{12}O_6$: $6(12.01\text{ g/mol}) + 12(1.01\text{ g/mol}) + 6(16.00\text{ g/mol}) = 180.2\text{ g/mol}$.
(c) $K_2Cr_2O_7$: $2(39.10\text{ g/mol}) + 2(52.00\text{ g/mol}) + 7(16.00\text{ g/mol}) = 294.2\text{ g/mol}$.
(d) $Fe_2(SO_4)_3$: $2(55.85\text{ g/mol}) + 3[32.07\text{ g/mol} + 4(16.00\text{ g/mol})] = 399.91\text{ g/mol}$.

Check Your Solution
Using round numbers for a quick check, for example in the case of (a), $14.00 + 3.00 = 17.00$. This estimate is close to the answer of 17.04 g/mol.

25. Problem
Strontium may be found in nature as celestite, $SrSO_4$. Find the molar mass of celestite.

What is Required?
You have to find the molar mass of celestite.

What is Given?
The chemical formula of celestite is $SrSO_4$. The average atomic mass can be obtained from any Periodic Table.

Plan Your Strategy
The molar mass of celestite is the sum of all the molar masses of its component elements. Find the atomic mass of 1 mol of Sr, 1 mol of S, and 4 mol of O from the Periodic Table and add them together.

Act on Your Strategy
$$\begin{aligned}\text{Molar mass of } SrSO_4 &= 87.62\text{ g/mol} + 32.07\text{ g/mol} + 4(16.00\text{ g/mol})\\ &= 183.69\text{ g/mol}\end{aligned}$$

Check Your Solution
Using round numbers for a quick check, you get $87 + 32 + (4 \times 16) = 183$. This estimate is close to the answer of 183.69 g/mol.

26. Problem
What is the molar mass of the ion $[Cu(NH_3)_4]^{2+}$?

What is Required?
You have to find the molar mass of the $[Cu(NH_3)_4]^{2+}$ ion.

What is Given?
The chemical formula of the ion is $[Cu(NH_3)_4]^{2+}$. The average atomic mass can be obtained from any Periodic Table.

Plan Your Strategy
The molar mass of the ion is the sum of all the molar masses of its component elements. Find the atomic mass of 1 mol of Cu, 4 mol of N, and 12 mol of H from the Periodic Table and add them together.

Act on Your Strategy

$$\begin{aligned}\text{Molar mass of } [Cu(NH_3)_4]^{2+} &= 63.55 \text{ g/mol} + 4(14.01 \text{ g/mol}) + 12(1.01 \text{ g/mol})\\ &= 63.55 \text{ g/mol} + 56.04 \text{ g/mol} + 12.12 \text{ g/mol}\\ &= 131.71 \text{ g/mol.}\end{aligned}$$

Check Your Solution
Using round numbers for a quick check, you get $63 + 56 + 12 = 131$. This estimate is close to the answer of 131.71 g/mol.

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27. Problem
Calculate the mass of each molar quantity.
(a) 3.90 mol of carbon, C
(b) 2.50 mol of ozone, O_3
(c) 1.75×10^7 mol of propanol, C_3H_8O
(d) 1.45×10^{-5} mol of ammonium dichromate, $(NH_4)_2Cr_2O_7$

What is Required?
You need to find the mass of the substances listed.

What is Given?
The number of moles of the substances are given as well as the chemical formulas. The molar mass of the elements that make up the substances can be found in a Periodic Table.

Plan Your Strategy
In order to convert moles to grams, you need to determine the molar mass (M) of the substance. The molar mass of a compound is the sum of all the molar masses of its component elements. Find the atomic mass (which is numerically equivalent to the molar mass) of the elements from the Periodic Table and add them together. Then multiply the molar mass by the given number of moles to obtain the mass (m) of the substance.

Act on Your Strategy
(a) For C: $M = 12.01$ g/mol

$$m = 3.90 \cancel{\text{mol}} \times 12.01 \text{ g/}\cancel{\text{mol}} = 46.8 \text{ g}$$

(b) For O_3: $M = 3(16.00\text{ g/mol}) = 48.00\text{ g/mol}$
$m = 2.50\cancel{\text{mol}} \times 48.00\text{ g/}\cancel{\text{mol}} = 120\text{ g}$

(c) For C_3H_8O: $M = 3(12.01\text{ g/mol}) + 8(1.01\text{ g/mol}) + 16.00\text{ g/mol} = 60.11\text{ g/mol}$
$m = 1.75 \times 10^7\cancel{\text{mol}} \times 60.11\text{ g/}\cancel{\text{mol}} = 1.05 \times 10^9\text{ g}$

(d) For $(NH_4)_2Cr_2O_7$: $M = 2[14.01\text{ g/mol} + 4(1.01\text{ g/mol})] + 2(52.00\text{ g/mol}) + 7(16.00\text{ g/mol}) = 252.10\text{ g/mol}$
$m = 1.45 \times 10^{-5}\cancel{\text{mol}} \times 252.10\text{ g/}\cancel{\text{mol}} = 3.66 \times 10^{-3}\text{ g}$

Check Your Solution
Work backwards. Divide the calculated mass by the number of moles given. It should give the molar mass of the substance. For example, in (a), $\frac{46.8\text{ g}}{3.90\text{ mol}} = 12\text{ g/mol}$. This matches the molar mass of carbon.

28. Problem
For each group, which sample has the largest mass?
(a) 5.00 mol of C, 150 mol of Cl_2, 0.50 mol of $C_6H_{12}O_6$
(b) 7.31 mol of O_2, 5.64 mol of CH_3OH, 12.1 mol of H_2O

What is Required?
You need to find the substance with the largest mass among the substances listed.

What is Given?
The number of moles of the substances are given as well as the chemical formulas. The molar mass of the elements that make up the substances can be found in a Periodic Table.

Plan Your Strategy
In order to convert moles to grams, you need to determine the molar mass (M) of the substance. The molar mass of a compound is the sum of all the molar masses of its component elements. Find the atomic mass (which is numerically equivalent to the molar mass) of the elements from the Periodic Table and add them together. Then multiply the molar mass by the given number of moles to obtain the mass (m) of the substance.

Act on Your Strategy
(a) For C: $M = 12.01\text{ g/mol}$; $m = 5.00\cancel{\text{mol}} \times 12.01\text{ g/}\cancel{\text{mol}} = 60.05\text{ g}$
For Cl_2: $M = 70.90\text{ g/mol}$; $m = 1.50\cancel{\text{mol}} \times 70.90\text{ g/}\cancel{\text{mol}} = 106.35\text{ g}$
For $C_6H_{12}O_6$: $M = 180.18\text{ g/mol}$; $m = 0.50\cancel{\text{mol}} \times 180.18\text{ g/}\cancel{\text{mol}} = 90.09\text{ g}$
Therefore, the sample of Cl_2 has the largest mass.
(b) For O_2: $M = 32.00\text{ g/mol}$; $m = 7.31\cancel{\text{mol}} \times 32.00\text{ g/}\cancel{\text{mol}} = 234\text{ g}$
For CH_3OH: $M = 32.05\text{ g/mol}$; $m = 5.64\cancel{\text{mol}} \times 32.05\text{ g/}\cancel{\text{mol}} = 181\text{ g}$
For H_2O: $M = 18.02\text{ g/mol}$; $m = 12.1\cancel{\text{mol}} \times 18.02\text{ g/}\cancel{\text{mol}} = 218\text{ g}$
Therefore, the sample of O_2 has the largest mass.

Check Your Solution
Work backward to check the mass. Divide the calculated mass by the number of moles given. It should give the molar mass of the substance. For example, for C in (a), $\frac{60.05\text{ g}}{5.00\text{ mol}} = 12.01\text{ g/mol}$. This matches the molar mass of carbon.

29. Problem
A litre, 1000 mL, of water contains 55.6 mol. What is the mass of a litre of water.

What is Required?
You need to find the mass of 1000 mL of water.

What is Given?
The number of moles of water is given as well as the volume. The molar mass of the elements that make up water can be found in a Periodic Table.

Plan Your Strategy
In order to convert moles to grams, you need to determine the molar mass (M) of water. Find the atomic mass (which is numerically equivalent to the molar mass) for 2 mol H and 1 mol O from the Periodic Table and add them together. Then multiply the molar mass by the given number of moles to obtain the mass (m).

Act on Your Strategy
Molar mass = 2(1.01 g.mol) + 16.00 g/mol = 18.02 g/mol
$m = 55.6\ \cancel{\text{mol}} \times 18.02\ \text{g}/\cancel{\text{mol}} = 1001.9\ \text{g}$

Check Your Solution
Since the density of water is 1.0 g/mL, the mass of 1000 mL is 1000 g or 1 kg. This agrees with the result calculated by the mole method.

30. Problem
To carry out a particular reaction, a chemical engineer needs 255 mol of styrenem C_8H_8. How many kilograms of styrene does the engineer need?

What is Required?
You need to find the mass of styrene needed.

What is Given?
The number of moles of the styrene is given as well as the chemical formula. The molar mass of the elements that make up styrene can be found in a Periodic Table.

Plan Your Strategy
In order to convert moles to grams, you need to determine the molar mass (M) of the styrene. The molar mass of a compound is the sum of all the molar masses of its component elements. Find the atomic mass (which is numerically equivalent to the molar mass) of the elements from the Periodic Table and add them together. Then multiply the molar mass by the given number of moles to obtain the mass (m) of the substance. Divide the mass in gram by 10^3 for a kilogram conversion.

Act on Your Strategy
$M = 8(12.01\ \text{g/mol}) + 8(1.01\ \text{g/mol}) = 104.16\ \text{g/mol}$
$m = 255\ \cancel{\text{mol}} \times 104.16\ \text{g}/\cancel{\text{mol}} = 26\ 561\ \text{g}$
Mass in kilogram = $26\ 561\ \text{g}/10^3 = 26.56\ \text{kg}$
Therefore, the engineer needs 26.6 kg of styrene for the reaction.

Check Your Solution
Work backwards. Divide the calculated mass by the number of moles given. It should give the molar mass of the substance. $\frac{26\ 561\ \text{g}}{255\ \text{mol}} = 104.06\ \text{g/mol}$, which agrees with the calculations.

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31. Problem
Calculate the number of moles in each sample.
(a) 103 g of Mo
(b) 1.32×10^4 g of Pd
(c) 0.736 kg of Cr
(d) 56.3 mg of Ge

What is Required?
You need to find the number of moles in the given mass of elements listed.

What is Given?
The masses of the elements are given. You have access to a Periodic Table.

Plan Your Strategy
In each case, determine the molar mass (M) from a Periodic Table and divide it into the given sample mass (m) to obtain the number of moles (n). Use $n = \frac{m}{M}$. In the case of (c) and (d), first convert the given amount to grams before proceeding with the calculation.

Act On Your Strategy
(a) For Mo: $M = 95.94$ g/mol; $m = 103$ g;
$$n = \frac{103 \cancel{g}}{95.94 \cancel{g}/\text{mol}} = 1.07 \text{ mol}$$
(b) For Pd: $M = 106.42$ g/mol; $m = 1.32 \times 10^4$ g;
$$n = \frac{1.32 \times 10^4 \cancel{g}}{106.42 \cancel{g}/\text{mol}} = 124 \text{ mol}$$
(c) For Cr: $M = 52.00$ g/mol; $m = 0.736 \text{ kg} \times 10^3 = 736$ g;
$$n = \frac{736 \cancel{g}}{52.00 \cancel{g}/\text{mol}} = 14.2 \text{ mol}$$
(d) For Ge: $M = 72.61$ g/mol; $m = 56.3 \text{ mg}/10^3 = 0.0563$ g;
$$n = \frac{0.0563 \cancel{g}}{72.61 \cancel{g}/\text{mol}} = 7.75 \times 10^{-4} \text{ mol}$$

Check Your Solution
In each case, work backward to reestablish the given mass of the element. For example, in (a), 1.07 mol × 95.94 g/mol = 102.66 g. This value matches the amount given in the question.

32. Problem
How many moles of compound are in each sample?
(a) 39.2 g of silicon dioxide, SiO_2
(b) 7.34 g of nitrous acid, HNO_2
(c) 1.55×10^5 kg of carbon tetrafluoride, CF_4
(d) 8.11×10^{-3} mg of 1-iodo-2,3,-dimethylbenzene, C_8H_9I

What is Required?
You need to find the number of moles in the given mass of substances listed.

What is Given?
The masses of the substances are given as well as the chemical formulas. You have access to a Periodic Table.

Plan Your Strategy
In each case, determine the molar mass (M) from a Periodic Table and divide it into the given sample mass (m) to obtain the number of moles (n). Use $n = \frac{m}{M}$. In the case of (c) and (d), first convert the given amount to grams before proceeding with the calculation.

Act On Your Strategy
(a) For SiO_2: $M = 60.09$ g/mol; $m = 39.2$ g;
$$n = \frac{39.2 \cancel{g}}{60.09 \cancel{g}/\text{mol}} = 0.652 \text{ mol}$$
(b) For HNO_2: $M = 47.02$ g/mol; $m = 7.34$ g;
$$n = \frac{7.34 \cancel{g}}{47.02 \cancel{g}/\text{mol}} = 0.156 \text{ mol}$$
(c) For CF_4: $M = 88.01$ g/mol; $m = (1.550 \times 10^5 \text{ kg}) \times 10^3 = 1.550 \times 10^8$ g;
$$n = \frac{1.550 \times 10^8 \cancel{g}}{88.01 \cancel{g}/\text{mol}} = 17.6 \times 10^6 \text{ mol}$$

(d) For C_8H_9I: $M = 232.07$ g/mol; $m = 8.11 \times 10^{-3}$ mg/$10^3 = 8.11 \times 10^{-6}$ g;

$n = \frac{8.11 \times 10^{-6} \cancel{g}}{232.07 \cancel{g}/mol} = 3.49 \times 10^{-8}$ mol

Check Your Solution

In each case, work backward to reestablish the given mass of the element. For example, in (a), 0.652 mol × 60.09 g/mol = 39.18 g. This value matches the amount given in the question.

33. Problem

Sodium chloride, NaCl, can be used to melt snow. How many moles of sodium chloride are in a 10 kg bag?

What is Required?

You need to find the number of moles of NaCl in the bag.

What is Given?

The mass of the salt is 10 kg. You have access to a Periodic Table.

Plan Your Strategy

Determine the molar mass (*M*) from a Periodic Table and divide it into the given sample mass (*m*) to obtain the number of moles (*n*). Use $n = \frac{m}{M}$. Remember to convert the given amount to grams before proceeding with the calculation.

Act On Your Strategy

$M = 58.44$ g/mol; $m = 10$ kg $\times 10^3 = 10\ 000$ g;

$n = \frac{10,000 \cancel{g}}{58.44 \cancel{g}/mol} = 1.7 \times 10^2$ mol.

There is 1.7×10^2 mol of salt in 10 kg.

Check Your Solution

Work backward to reestablish the given mass of the salt.

$(1.7 \times 10^2$ mol$) \times 58.44$ g/mol $= 9934.8$ g. This value closely matches the amount given in the question (variations depend on the number of significant units used in calculations).

34. Problem

Octane, C_8H_{18}, is a principal ingredient of gasoline. Calculate the number of moles in a 20.0 kg sample of octane.

What is Required?

You need to find the number of moles of octane in the sample of gasoline.

What is Given?

The mass of the octane is 20.0 kg and its chemical formula is given. You have access to a Periodic Table.

Plan Your Strategy

Determine the molar mass (*M*) of octane using the atomic mass units from a Periodic Table. Divide *M* into the given sample mass (*m*) to obtain the number of moles (*n*). Use $n = \frac{m}{M}$. Remember to convert the given amount to grams before proceeding with the calculation.

Act On Your Strategy

$M = 114.26$ g/mol; $m = 20.0$ kg $\times 10^3 = 20\ 000$ g;

$n = \frac{20,000 \cancel{g}}{114.26 \cancel{g}/mol} = 175$ mol.

Check Your Solution

Work backward to reestablish the given mass of octane.

175 mol × 114.26 g/mol = 19 995.5 g. This value closely matches the 20.0 kg

amount given in the question (variations depend on the number of significant units used in calculations).

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35. Problem

Determine the mass of each sample.

(a) 6.02×10^{24} formula units of $ZnCl_2$

(b) 7.38×10^{21} formula units of $Pb_3(PO_4)_2$

(c) 9.11×10^{23} molecules of $C_{15}H_{21}N_3O_{15}$

(d) 1.20×10^{29} molecules of N_2O_5

What is Required?

You need to find the masses of the substances listed.

What is Given?

You are given the number of formula units or molecules of the substances. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

Convert the number of formula units or the number of molecules into moles (n) by dividing by Avogadro's number. Then multiply the number of moles by the molar mass (M) to get the mass in grams (m) of the substance.

Act on Your Strategy

(a) $n = \dfrac{(6.02 \times 10^{24}\ \cancel{\text{formula units}})}{(6.02 \times 10^{23})\ \cancel{\text{formula units}}/\text{mol}} = 10\text{ mol}$

M of $ZnCl_2 = 136.29$ g/mol;

Therefore, $m = 10\ \cancel{\text{mol}} \times 136.29\text{ g}/\cancel{\text{mol}} = 1362.9\text{ g}$ or 1.363×10^3 g

(b) $n = \dfrac{7.38 \times 10^{21}\ \cancel{\text{formula units}}}{(6.02 \times 10^{23})\ \cancel{\text{formula units}}/\text{mol}} = 0.0122\text{ mol}$

M of $Pb_3(PO_4)_2 = 811.54$ g;

Therefore, $m = 0.0122\ \cancel{\text{mol}} \times 811.54\text{ g}/\cancel{\text{mol}} = 9.95$ g

(c) $n = \dfrac{(9.11 \times 10^{23}\ \cancel{\text{molecules}})}{(6.02 \times 10^{23})\ \cancel{\text{molecules}}/\text{mol}} = 1.51\text{ mol}$

M of $C_{15}H_{21}N_3O_{15} = 483.39$ g/mol;

Therefore, $m = 1.51\ \cancel{\text{mol}} \times 483.39\text{ g}/\cancel{\text{mol}} = 732$ g

(d) $n = \dfrac{(1.20 \times 10^{29}\ \cancel{\text{molecules}})}{(6.02 \times 10^{23})\ \cancel{\text{molecules}}/\text{mol}} = 1.99 \times 10^5\text{ mol}$

M of $N_2O_5 = 108.02$ g/mol;

Therefore, $m = 1.99 \times 10^5\ \cancel{\text{mol}} \times 108.02\text{ g}/\cancel{\text{mol}} = 2.15 \times 10^7$ g or 2.15×10^4 kg.

Check Your Solution

Work backwards. Dividing the mass by the number of moles should give the molar mass of the substance. For example, in (a), $\dfrac{1362.9\text{ g}}{10\text{ mol}} = 136.29$ g/mol. The results match.

36. Problem

What is the mass of lithium in 254 formula units of lithium chloride, LiCl?

What is Required?

You need to find the mass of lithium in the LiCl.

What is Given?

You are given the number of formula units of lithium chloride as well as the chemical formula. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

From the chemical formula, you know there is 1 unit of lithium for every formula unit of lithium chloride, so you can assume that there are 254 units of lithium in the sample. Convert the number of units of the lithium into moles (n) by dividing by Avogadro's number. Then multiply the number of moles by the molar mass (M) of lithium only to get the mass in grams (m) of the lithium.

Act on Your Strategy

$$n = \frac{254 \cancel{\text{formula units}}}{(6.02 \times 10^{23}) \cancel{\text{formula units}}/\text{mol}} = 4.22 \times 10^{-22} \text{ mol}$$

M of Li $= 6.94$ g/mol;

Therefore, m $= (4.22 \times 10^{-22}) \cancel{\text{mol}} \times 6.94 \text{ g}/\cancel{\text{mol}} = 2.93 \times 10^{-21}$ g.

Check Your Solution

Work backwards. Dividing the mass by the number of moles should give the molar mass of lithium. $\frac{2.93 \times 10^{-21}\text{ g}}{4.22 \times 10^{-22}\text{ mol}} = 6.94$ g/mol. The results match.

37. Problem

Express the mass of a single atom of titanium, Ti, in grams.

What is Required?

You have to find the mass of a single Ti atom.

What is Given?

The chemical symbols and chemical names are given. You have access to a periodic table, and Avogadro's number, $N_A = 6.02 \times 10^{23}$ atoms/mol.

Plan Your Strategy

According to the Avogadro constant, there are 6.02×10^{23} atoms of Ti in every mole of titanium. Therefore, the number of moles (n) of 1 atom of Ti is 1 mol divided by 6.02×10^{23} atoms. Multiply this value by the molar mass (M) of Ti to obtain the mass (m) of 1 atom of Ti.

Act on Your Strategy

Number of moles of 1 atom $= 1$ mol/6.02×10^{23} atoms $= 1.66 \times 10^{-24}$ mol

M of Ti $= 47.87$ g/m;

Therefore, $m = 1.66 \times 10^{-24} \cancel{\text{mol}} \times 47.87 \text{ g}/\cancel{\text{mol}} = 7.95 \times 10^{-23}$ g.

Check Your Solution

Work backwards. Dividing the mass by the number of moles should give the molar mass of titanium. $\frac{7.95 \times 10^{-23}\text{ g}}{1.66 \times 10^{-24}\text{ mol}} = 47.90$ g/mol. The results match.

38. Problem

Vitamin B_2, $C_{17}H_{20}N_4O_6$, is also called riboflavin. What is the mass, in grams, of a single molecule of riboflavin?

What is Required?

You have to find the mass of a single riboflavin molecule.

What is Given?

The chemical formula of the vitamin is given. You have access to a periodic table, and Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy

According to the Avogadro constant, there are 6.02×10^{23} molecules of riboflavin in every mole of the vitamin. Therefore, the number of moles (n) of 1 molecule of riboflavin is 1 mol divided by 6.02×10^{23} molecules. Multiply this value by the molar mass (M) of $C_{17}H_{20}N_4O_6$ to obtain the mass (m) of a single molecule.

Act on Your Strategy

Number of moles of 1 molecule $= \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}} = 1.66 \times 10^{-24}$ mol

M of $C_{17}H_{20}N_4O_6 = 376.41$ g/m;

Therefore, $m = 1.66 \times 10^{-24} \cancel{\text{mol}} \times 376.41 \text{ g}/\cancel{\text{mol}} = 6.25 \times 10^{-22}$ g.

Check Your Solution

Work backwards. Dividing the mass by the number of moles should give the molar mass of titanium. $\frac{6.25 \times 10^{-22} \text{ g}}{1.66 \times 10^{-24} \text{ mol}} = 376.51$ g/mol. The results match.

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39. Problem

Determine the number of molecules or formula units in each sample.

(a) 10.0 g of water, H_2O

(b) 52.4 g of methanol, CH_3OH

(c) 23.5 g of disulfur dichloride, S_2Cl_2

(d) 0.337 g of lead(II) phosphate, $Pb_3(PO_4)_2$

What is Required?

You need to find the number of formula units or molecules of the substances listed.

What is Given?

You are given the masses of the substances as well as their molecular formulas. You know Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

First convert the mass (m) to moles (n), using the molar mass (M) of the substance and the formula $n = m/M$. Multiplying this value by Avogadro's number will yield the number of formula units (or the number of molecules).

Act on Your Strategy

(a) $n = \frac{10.0 \cancel{\text{g}}}{18.02 \cancel{\text{g}}/\text{mol}} = 0.555$ mol

Therefore, the number of molecules of water

$= 0.555 \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}} = 3.34 \times 10^{23}$ molecules.

(b) $n = \frac{52.4 \cancel{\text{g}}}{32.05 \cancel{\text{g}}/\text{mol}} = 1.63$ mol

Therefore, the number of molecules of CH_3OH.

$= 1.63 \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}} = 9.84 \times 10^{23}$ molecules.

(c) $n = \frac{23.5 \cancel{\text{g}}}{135.04 \cancel{\text{g}}/\text{mol}} = 0.174$ mol

Therefore, the number of molecules of S_2Cl_2

$= 0.174 \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}} = 1.05 \times 10^{23}$ molecules.

(d) $n = \frac{0.337 \cancel{\text{g}}}{811.54 \cancel{\text{g}}/\text{mol}} = 4.15 \times 10^{-4}$ mol

Therefore, the number of formula units of $Pb_3(PO_4)_2$

$= 4.15 \times 10^{-4} \cancel{\text{mol}} \times (6.02 \times 10^{23})$ formula units/$\cancel{\text{mol}}$

$= 2.50 \times 10^{20}$ formula units.

Check Your Solution

In each case, work backwards to reestablish the mass given in the question. For example, in (a), each mole of water has a mass of 18.02 g.

Therefore, 3.34×10^{23} molecules of water has a mass of:

$3.34 \times 10^{23} \cancel{\text{molecules}} \times \left(\frac{1 \cancel{\text{mol}}}{6.02 \times 10^{23} \cancel{\text{molecules}}}\right) \times \left(\frac{18.02 \text{ g}}{1 \cancel{\text{mol}}}\right) = 9.99$ g

The answer is close to the 10.0 g given in the question (a). Your answer is reasonable.

40. Problem

How many atoms of hydrogen are in 5.3×10^4 molecules of sodium glutamate, $NaC_5H_8NO_4$?

What is Required?

You need to find the number of H atoms in the given number of molecules.

What is Given?

You are given the number of molecules of sodium glutamate as well as its molecular formula.

Plan Your Strategy

From the formula, you can tell that there are 8 hydrogen atoms for every formula unit of sodium glutamate. The total number of hydrogen atoms is therefore 8 multiplied by the given number of molecules.

Act on Your Strategy

$$\text{Total number of H atoms} = 8\ \text{atoms}/\cancel{\text{molecule}} \times (5.3 \times 10^4)\ \cancel{\text{molecules}}$$
$$= 4.2 \times 10^5\ \text{atoms}$$

Check Your Solution

Using round numbers for a quick check, $8 \times 50\ 000 = 400\ 000$ or 4×10^5. This estimate is close to the answer.

41. Problem

How many molecules are in a 64.3 mg sample of tetraphosphorus decoxide, P_4O_{10}?

What is Required?

You need to find the number of molecules in the P_4O_{10} sample.

What is Given?

You are given the mass of the tetraphosphorus decoxide as well as its molecular formula. You know Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy

First convert the mass (m) to moles (n), using the molar mass (M) of the tetraphosphorus decoxide and the formula $n = m/M$. Multiplying this value by Avogadro's number will yield the number of molecules. Remember to convert the mass to grams before proceeding with the calculation.

Act on Your Strategy

$m = \frac{64.3\ \text{mg}}{10^3} = 6.43 \times 10^{-2}\ \text{g}$; $M = 283.88$ g/mol;

Therefore, $n = \frac{6.43 \times 10^{-2}\ \cancel{\text{g}}}{283.88\ \cancel{\text{g}}/\text{mol}} = 2.26 \times 10^{-4}$ mol

Hence, the number of molecules of P_4O_{10}

$= 2.26 \times 10^{-4}\ \cancel{\text{mol}} \times (6.02 \times 10^{23})\ \text{molecules}/\cancel{\text{mol}} = 1.36 \times 10^{20}$ molecules.

Check Your Solution

Work backwards to reestablish the mass given in the question. Therefore,

1.36×10^{20} molecules of P_4O_{10} has a mass of:

$1.36 \times 10^{20}\ \cancel{\text{molecules}} \times \left(\frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}}\right) \times \left(\frac{283.88\ \text{g}}{1\ \cancel{\text{mol}}}\right)$

$= 0.0641$ g or 64.1 mg.

The answer is close to the 64.3 mg given in the question.

42. Problem

(a) How many formula units are in a 4.35×10^{-2} g sample of potassium chlorate, $KClO_3$?

(b) How many ions (chlorate and potassium) are in this sample?

What is Required?

(a) You need to find the number of formula units in the given sample.

(b) You need to find the number of chlorate and potassium ions in this sample.

What is Given?

(a) You are given the mass of the potassium chlorate.

(b) You are given its molecular formula.

You know Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

(a) First convert the mass (m) to moles (n), using the molar mass (M) of the potassium chlorate and the formula $n = \frac{m}{M}$. Multiplying this value by Avogadro's number will yield the number of formula units.

(b) From the chemical formula of potassium chlorate, you know there is 1 ion of potassium and 1 ion of chlorate for every formula unit of potassium chlorate. The total number of ions will therefore be the number of formula units calculated in (a) multiplied by 2.

Act on Your Strategy

(a) $m = 4.35 \times 10^{-2}$ g; $M = 122.55$ g/mol;

Therefore, $n = \frac{4.35 \times 10^{-2}\ \cancel{\text{g}}}{122.55\ \cancel{\text{g}}/\text{mol}} = 3.55 \times 10^{-4}$ mol

Hence, the number of formula units of $KClO_3$

$= 3.55 \times 10^{-4}\ \cancel{\text{mol}} \times (6.02 \times 10^{23})$ molecules/$\cancel{\text{mol}}$

$= 2.14 \times 10^{20}$ formula units

(b) Each formula unit of $KClO_3$ consists of two ions, so there are

2 ions/~~formula unit~~ $\times\ (2.14 \times 10^{20})$ ~~formula units~~

$= 4.28 \times 10^{20}$ ions in the sample.

Check Your Solution

(a) Work backwards to reestablish the mass given in the question. Therefore, 2.14×10^{20} formula units of $KClO_3$ has a mass of:

$2.14 \times 10^{20}\ \cancel{\text{formula units}} \times \left(\frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}}\right) \times \left(\frac{122.55\ \text{g}}{1\ \cancel{\text{mol}}}\right)$

$= 4.36 \times 10^{-2}$ g. The answer is close to the 4.35×10^{-2} g given in the question.

(b) Using round numbers for a quick check, $2 \times (2 \times 10^{20}) = 4 \times 10^{20}$. This estimate is close to the answer.