1. A clear unlabelled beaker with a small crack containing what looks like water is sitting on a lab bench beside a lit Bunsen burner. What are the hazards that are **most likely** present in this situation?

   A. The unknown solution may be hazardous, beaker may break if moved, and the glass may be hot
   B. The water in glass may be hot, and it is possible the glass could break if dropped on floor
   C. The solution may be dangerous and may leak out of the crack if the solution is heated
   D. The solution is not labelled and may be dangerous, and the heat from the burner may break glass

2. If a small fire starts in the classroom you should immediately

   A. use the fire extinguisher and inform the teacher and class of the fire
   B. tell the teacher and ask for permission to use the fire extinguisher
   C. run from the room, pull the fire alarm and then go tell the teacher
   D. yell out a warning to the class and then run out to pull fire alarm

3. What should you do if your lab partner catches her hair and shirt on fire in the Bunsen burner?

   A. Have her run to emergency shower to douse the flames.
   B. Have her stop drop and roll and then help her smother the flames.
   C. Have her move quickly to the fire blanket and smother the flames.
   D. Have her shove her head under the tap and splash water on the shirt.

Measurement and Communication

1.2.1 Use SI units and their values in Chemistry

1.2.2 Demonstrate skills in measuring mass, volume (liquid), and temperature

SI, International System, of metric units is the standard form of measurement in science. The common units used in Chemistry 11 are:

UNIT	SYMBOL	MEASURES
Metre	m	Length
Kilogram	kg	Mass
Second	s	Time
Mole	mol	Amount of a substance
Derived Measures		
Litre	L	Volume

SI Prefixes		
NAME	**SYMBOL**	**FACTOR**
Micro	μ	10^{-6}
Milli	m	10^{-3}
Centi	c	10^{-2}
Kilo	k	10^{3}
Mega	M	10^{6}

Related Questions:

4. The metric base unit for measuring distance is

   A. acre
   B. celsius
   C. kilometre
   D. metre

5. Which of the following is **not** an SI base unit?

   A. Metre
   B. Second
   C. Centigrade
   D. Millilitre

6. How many millimetres are in 2.5 kilometres?

   A. 2.5×10^{-6} mm
   B. 2.5×10^{-3} mm
   C. 2.5×10^{3} mm
   D. 2.5×10^{6} mm

7. To accurately convert 7.62 mm to m in one step, a student should

A. multiply 7.62 mm by 1 000
B. divide 7.62 mm by 1 000
C. move the decimal in 7.62 two places to the left
D. move the decimal in 7.62 two places to the right

8. Convert 1.750×10^5 L to mL

A. 1.75×10^9 mL
B. 175,000 mL
C. 1.750×10^2 mL
D. 1.750×10^8 mL

9. Convert 13.5 kg into micrograms:

A. 1.35×10^{10} μg
B. 1.35×10^7 μg
C. 1.35×10^{-8} μg
D. 1.35×10^4 μg

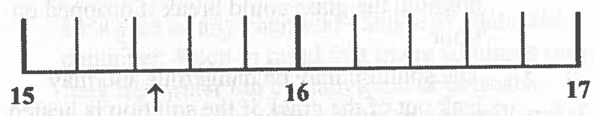
1.2.3 Describe the imprecise nature of all measurements

1.2.4 Determine the number of significant figures in a measured quantity and relate to the uncertainty

1.2.5 Round off calculated results to the appropriate number of significant figures

The numbers we work with in chemistry, or any science, are only as good as the tools we use to acquire them. The calculations we do with these values must reflect the accuracy of our instruments. Values which we know are very clearly measured are known as certain digits. The certain digits are followed by an uncertain digit which is an estimate and is just outside the ability of the instrument to measure precisely. The combination of the certain digits and the one uncertain digit are known as the significant figures.

Refer to the example ruler below. The arrow is indicating a point which is clearly between 15 and 16 so we are certain the value is at least 15. Each small division between 15 and 16 has a value of 0.2 so we can further state that the arrow is indicating a point between 15.4 and 15.6 so we are certain the value is at least 15.4. We could further guess the indicated point is at 15.46 but the last digit, the 6, is somewhat uncertain as it is an estimate. The 6 has some significance but not as much as the 15.4 and is thus referred to as the uncertain digit. This measurement, 15.46, has four significant figures; three certain digits and one uncertain digit.



Significant digits and when zero is significant

* all nonzero digits are significant
* zeroes are significant if they appear between nonzero digits. E.g. 101
* zeroes are significant if they appear to the right of a nonzero digit and there is a decimal place present. E.g. 20.0

Examples

* **15** has two significant figures
* **10** has one significant figures as the zero does not count
* **101** has three significant figures as the zero counts because it is between two nonzero digits
* **10.0** has three significant figures as the zeroes are to the right of the nonzero digit and there is a decimal place
* **0.050** has two significant figures; the first two zeroes are not significant and are merely place values; the last zero is significant as it is to the right of a nonzero digit and there is a decimal place
* **1.30×10^{-4}** has three significant figures; all numbers are significant in scientific notation

Operations with significant figures

* Any number that represents a numerical count or is an exact definition has an unlimited number of significant figures. Is considered to be an absolute value.
* Superfluous digits, those beyond the significant figures, are rounded to the nearest significant figure. If it is 5 or more it

is rounded up and if it is less than five the superfluous digit is dropped.

* In multiplication and division the answer of the operation can have no more significant figures than the factor having the fewest number of significant figures.

* The root or power of a number will have as many significant figures as the number.

* When adding or subtracting arrange the numbers in columns and round any numbers to the right of the least certain digit present.

Examples

* $45.0 \div 9.0 = 5.0$ The answer, 5.0, can have two significant figures as 9.0 has two significant figures and the answer can only be as accurate as the least accurate value used in any multiplication or division operation

* $46.4 + 4 + 100.55 = 151$ The 4 is the least certain of the numbers given as it is only giving a value to the ones position while the 46.4 is giving values to the tenths position and 100.55 is giving values to the hundredths position. All digits must then be rounded to the ones position to reflect the accuracy of the 4. The answer can be no more accurate the least accurate value used.

Related Questions:

10. Which of the following does not have three significant figures?

A. 1 340
B. 0.078
C. 1.55×10^{-2}
D. 0.0105

1.2.6 Correctly determine the units of a derived quantity

Derived Units

These are units that are produced by combining two or more other units. A common use of this in everyday life is velocity; Km/hr. Velocity is the combination of distance and time.

Velocity = distance per time

v = m/s as a base unit and is converted to Km/hr when used as a speed in our society. Velocity's unit is m/s or Km/hr which are derived units combining the units of distance and time.

Density is a common derived unit used in chemistry and is a function of mass and volume.

Density = mass/volume
$d = g/mL$ or g/cm^3

1.2.7 State the acceptability of the numerical results of a lab experiment with regard to the uncertainty of the results

1.2.8 Communicate results and data in clear and understandable forms

Experimental uncertainty refers to the degree to which any measured value might be in error. The uncertainty for any measurement is indicated after the number.

Example

35.6 ± 0.1 mm this tells us that the value lies between 35.5mm and 35.7mm

Matter and Its Changes

1.3.1 Define matter

1.3.2 Describe chemistry as the science concerned with the properties, composition and behaviour of matter

1.3.3 Describe and give examples of a variety of forms and properties that matter can exhibit

1.3.4 Distinguish between observation and interpretation

1.3.5 Describe the types of changes that may be observed when matter is heated, combined, or separated

1.3.6 Differentiate between physical and chemical changes

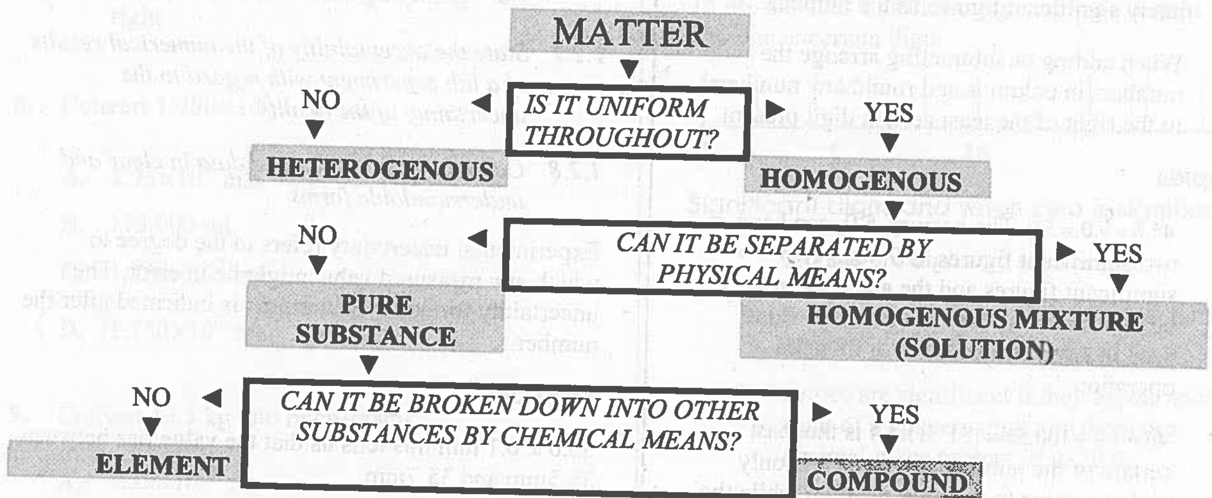
1.3.7 Classify a substance as solid, liquid, or gas, and describe its different properties

1.3.8 Define boiling point, freezing point, and melting point

1.3.9 Describe the simple molecular motions and arrangements for solids, liquids, and gases

1.3.10 Relate the heat changes that occur during phase changes to changes in molecular motions and arrangements

Matter can be defined as anything that has mass and occupies space. There are more issues with this simple definition than one would first think. For example why does matter exist? Mass can be defined as an objects resistance to a change in its motion but this explains little as to what or why matter really is. Space then also becomes difficult to describe; its definition is fairly circular – space is that which is occupied by matter yet space is used in the definition of matter.



a) Elements

These are pure substances which cannot be broken down by further chemical means. An atom can be destroyed but not by chemical means; it requires a nuclear reaction. An element is composed of only one kind of atom.

b) Compounds

These are pure substances composed of two or more atoms of different elements chemically bonded together. A compound can be destroyed by chemical means. It can be broken down into the elements it is composed of or it can be combined with other elements or compounds to produce new substances.

c) Mixtures

These pure substances are composed of two or more substances, but each keeps its original properties.

d) Atom

It is the smallest object that retains properties of an element. It is composed of electrons and a nucleus (containing protons and neutrons).

e) Molecule

These pure substances are simply two or more atoms chemically bonded together. The atoms do not have to be different atoms and a molecule can be broken down by chemical means into the atoms of which it is made.

Chemistry is the science that is concerned with the properties, composition and behaviour of matter. All substances have a unique set of properties which will characterize their behaviour. These properties can also be used to identify the substances. There are two types of properties used in chemistry:

* Physical properties: these are properties that can be established without changing the actual substance. Some examples of physical properties are density, colour, electrical conductivity, thermal conductivity, melting and boiling points and hardness.
* Chemical properties: these properties describe the substance's ability to chemically react and produce new substances with new properties. Some examples of chemical properties would be iron's ability to react to oxidize in the presence of oxygen and water to produce Iron III oxide (rust) and hydrogen's ability to violently burn with oxygen to produce water.

Physical Change involves changes that are observable but do not change the underlying structure of the substance. Such change includes freezing, evaporating, or filtering.

Chemical Change occurs when chemical reactions create new chemical substances. Evidence that chemical change has occurred may be energy (heat) release, formation of a gas or precipitate, a colour change, or an odour change.

All substances will exist in one of the various states (phases) of matter. There are several phases of matter but only three which are relevant at this time – solid, liquid and gas.

* Solid: definite shape, definite volume
* Liquid: indefinite shape, definite volume
* Gas: indefinite shape, indefinite volume

Indefinite shape means the substance does not take on the shape of the container. Indefinite volume means the substance would expand to fill the entire compound.

These are the common states of matter used in chemistry. There are other states of matter such as plasma and Bose-Einstein condensate, but they are commonly thought of as in the realm of physics. Plasma exists at extremely high temperatures, several million degrees Celsius, and exists as nuclei and free electrons. Bose-Einstein condensate exists at extremely low temperatures near absolute zero where many atoms of an element are thought to merge into one super atom of the element.

No matter what state matter is in it will be in motion unless the temperature is absolute zero (-273°C). Anything in motion has kinetic energy and there are three basic ways, thus energies, in which molecules can be moving.

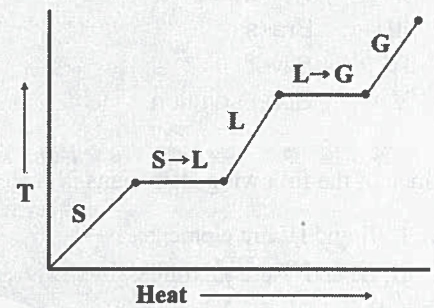
* Translational energy – the molecule is moving in a straight line
* Rotational energy – the molecule is rotating about one of its axes
* Vibrational energy – the molecule is vibrating and the bond lengths and angles between atoms are changing

Energy and phase changes

Matter in the solid state has the least amount of energy and the molecules are not moving about freely but are vibrating. As the kinetic energy of the matter increases, usually due to a temperature increase, the molecules begin to oscillate back and forth. The energy gained is mostly translational and vibrational energy. They will impact their neighbours more frequently and with greater force as the kinetic energy increases. At a given point for any type of matter the molecules will have enough energy to break the bonds which hold them as a solid and move apart from one another. This is melting and the matter is in the liquid state with the molecules moving about freely.

The molecules in the liquid state will also move about with greater speed as energy is continually added to the system. They will move about with ever increasing speed until they have enough energy to break away from the other molecules in the liquid and enter the gas phase – boiling.

Substances will go through phase changes as temperature changes. The heating curve below shows the typical behaviour of a pure substance as it is heated.



On the above graph the temperature of the solid, S, increases until it hits the melting point. The melting temperature and the freezing temperature of a pure substance are the same, depending upon whether or not the temperature is increasing or decreasing. At the melting point, the graph plateaus as all the energy entering the system is being used to change phase from solid to liquid. Solid and liquid will coexist at this temperature. Once the solid has changed phase into liquid the temperature will again increase until the liquid hits the boiling point of the substance. At this temperature the substance begins to change phase into a gas and the temperature plateaus. All the heat entering the system is being used to for the state change. Liquid and gas will coexist at this temperature. Once the liquid has changed phase into a gas the temperature will again begin to rise.

Related Questions:

Use the following information to answer the next question.

I	Air
II	Methane
III	Water
IV	Silver
V	Carbon dioxide

11. Which of the following statements is correct?

A. I, III and IV are elements.
B. II, III and V are compounds.
C. III and IV are elements.
D. III and IV are mixtures.

Use the following list to answer the next question.

I	Zinc
II	Hydrogen
III	Brass
IV	Silver
V	Soap solution

12. Which of the following statements is correct?

A. I, III and IV are elements.
B. II, III and V are mixtures.
C. I, II and IV are elements.
D. III and IV are compounds.

13. What is the effect on the particles in a system if heat is added to the system?

A. They begin to move faster and the temperature decreases.
B. They begin to move faster and the temperature increases.
C. They begin to mover slower and the temperature increases.
D. They begin to move slower and the temperature decreases.

14. As translation energy of a system is reduced, the volume of the system

A. decreases
B. increases
C. doubles
D. remains constant

Written Response

1. A pure gold nugget has a mass of 3.56 g and volume of 0.184 cm^3. What is the density of the gold nugget?

19.35

2. A 4.5g piece of iron is dropped into a graduated cylinder containing 10.2mL of water. Given that the density of iron is 7.9 g/cm^3, what is the new volume reading on the graduated cylinder?

3. Determine the number of significant figures for the following:

i) 120 2
ii) 1.050 4
iii) 0.345 3
iv) 1010 3
v) 6.02×10^{23} 3

4. Calculate the following showing the answer using correct significant figures.

i) $237050 \times 3.41 =$ 809000
ii) $435.00 \div 0.03100 =$ 14030
iii) $23.1 + 2.55 + 9.367 =$ 35.1

TABLE OF CORRELATIONS

	Atoms, Molecules, and Ions		
	Outcomes		Questions
	It is expected that students will:		
Classification	2.1.1	describe a substance as having a set of unique and identifiable properties	
	2.1.2	classify a given material as an element, compound, or mixture, using the properties of the material	
	2.1.3	describe several ways that substances may be separated from one another	2
	2.1.4	relate the observable properties and characteristics of elements, compounds, and mixtures to the concept of atoms and molecules	
	2.1.5	define atom, molecule, and ion	1
Nomenclature	2.2.1	write chemical symbols for elements and formulae for ions from appropriate charts	3, 13
	2.2.2	name the ionic compound from a formula, and write the formula given a name	4, 5
	2.2.3	name the covalent compound from a formula using the prefix naming system, and write the formula given a name	6, 7, 8, 9, 10, 11, WR1
	2.2.4	predict the formulae of covalent compounds given the formula of another compound containing elements in the same family (families)	12
	2.2.5	write the names and formulae for some common acids	14

Atoms, Molecules, and Ions

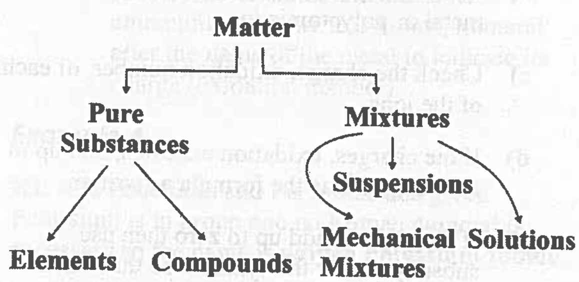
CLASSIFICATION

2.1.1 Describe a substance as having a set of unique and identifiable properties

Any substance has a set of properties which are collectively unique to it and are uniform throughout it. These properties can be used to identify the substance. Examples of these properties are density, boiling temperature, melting temperature, hardness, malleability, lustre and ductility. Not all properties would be used for all substances however determining several of them would allow you to classify the substance in question.

2.1.2 Classify a given material as an element or mixture using the properties of the material

The following classification of matter is based upon its initial appearance and then by whether it can be separated by physical or chemical means:



2.1.3 Describe several ways that substances can be separated from one another

a) Hand Separation

This is the actually act of picking apart a mixture using your hands or perhaps a sieve or magnet. This method is only possible with mechanical mixtures.

b) Filtration

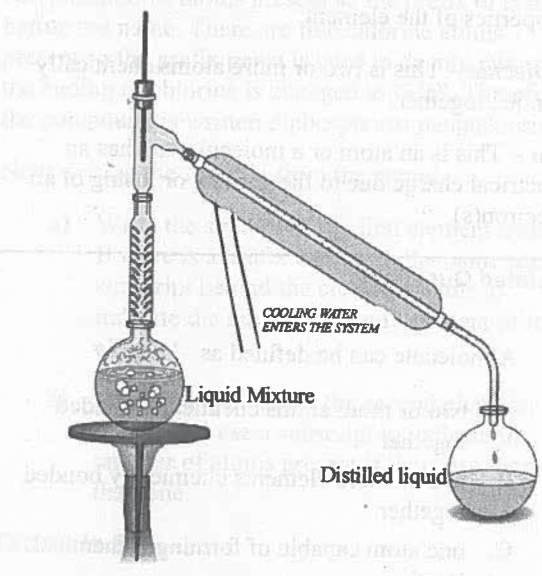
This method is good for separating a solid from a liquid as long as the solid is not in solution with the liquid. It is in essence a mechanical mixture.

c) Evaporation

This method is again good for separating a solid from a liquid. It allows you to retain the solid after the liquid has evaporated away. It can be done with mechanical mixtures of solid and liquid and with solids that are in solution with the liquid. Evaporation will not allow you to retain the liquid as it is lost to the atmosphere.

d) Distillation

This method is primarily for mixtures of liquids where the liquids are in solution with one another. The liquids will have different boiling temperatures as they are different substances and it is this property which is used to separate them. The liquid that has the lower boiling temperature will boil first and enter the gas phase. The gas is condensed when it hits the cooler temperature of the condenser and is collected separately from the remaining liquid in the original flask.



e) Gravity Separation

This method is used to separate solids in a liquid, a mechanical mixture, using the solid's different densities. A common example of this is the centrifuge which spins at extremely high speeds causing solids in the liquid to collect on the bottom of a tube with the most dense solids being on the outside and the less dense solids on top.

f) Chromatography

This method is used to separate coloured solids which are in solution with a liquid. Absorbent material, such as paper or talc – a silica based product, is placed in the solvent and the solvent with the dissolved solids is absorbed up the material. The solids are separated by their differing abilities to dissolve in the solvent versus the absorbent material. Those that dissolve better in the absorbent material than the solvent do not travel as far up the material as those that dissolve better in the solvent than the absorbent material.

When dried, the separated solids can be separated from the absorbent material and retained as a pure substance.

2.1.4 *Relate the observable properties and characteristics of elements, compounds, and mixtures to the concept of atoms and molecules*

2.1.5 *Define atom, molecule and ion*

Atom – This is the smallest amount or quantity of an element you can have that will still retain the properties of the element.

Molecule – This is two or more atoms chemically bonded together.

Ion – This is an atom or a molecule that has an electrical charge due to the gaining or losing of an electron(s).

Related Questions:

1. A molecule can be defined as
   - **A.** two or more atoms chemically bonded together
   - **B.** two or more elements chemically bonded together
   - **C.** one atom capable of forming a chemical bond
   - **D.** the smallest possible portion of a mole

2. A reasonable method of separating a suspended solid from a liquid is
   - **A.** chromatography
   - **B.** distillation
   - **C.** filtration
   - **D.** hand separation

NOMENCLATURE

2.2.1 *Write chemical symbols for elements and formulae for ions from appropriate charts*

2.2.2 *Name the ionic compound from a formula and write the formula given a name*

2.2.3 *Name the covalent compound from a formula using the prefix naming system, and write the formula given a name*

2.2.4 *Predict the formulae of covalent compounds given the formula of another compound containing elements in the same family (families)*

2.2.5 *Write the names and formulae for some common acids*

Nomenclature is the terminology of chemical compounds. It is necessary to understand the language of chemistry in order to work with it.

Ionic Compound basically for naming purposes a metal ion and a non-metal ion. First is writing the formula from the name:

a) Write the symbol for the metal ion (or ammonium ion).

b) Write the symbol or formula for the non-metal or polyatomic ion.

c) Check the charge, oxidation number, of each of the ions.

d) If the charges, oxidation numbers, add up to zero then that is the formula as written.

e) If they do not add up to zero then use subscripts after the symbols so that when you multiply them by the charges the sum is zero.

Example 1

Sodium	Chloride
↓	↓
Na^+	Cl^-

Sodium is +1 and Chlorine is –1 which has a sum of zero so the formula is written as **NaCl**.

Example 2

Copper I	Sulphate	(note Roman numeral indicates the charge on the copper)
↓	↓	
Cu^+	SO_4^{-2}	

Copper is +1 and Sulphate is – 2 so a subscript of 2 after copper would give us + 2 and then the sum would be zero. The formula is written Cu_2SO_4.

Example 3

Iron III	Carbonate	(note Roman numeral indicates charge on the Iron)
↓	↓	
Fe^{+3}	CO_3^{-2}	

Iron is + 3 and Carbonate is – 2 so a subscript of 2 after Iron and a subscript of 3 after Carbonate would give a + 6 and a –6 with a sum of zero. The formula is written $Fe_2(CO_3)_3$.

Secondly is writing the name from the formula for Ionic Compounds:

a) Write the name of the positive ion (metal or ammonium).

b) Write the name of the negative ion.

c) If the positive is **not** in group 1 or 2, is not silver, zinc or aluminum and is not ammonium then write a Roman numeral after the name of the metal to indicate its charge (oxidation number).

Example 4

KI: K is Potassium and I is Iodide and given Potassium is in group one no Roman numeral is necessary so the name is written **Potassium Iodide**.

Example 5

YF_2 : Y is Yttrium and F is Fluoride. The fluoride has a charge of –1 and a subscript of 2 therefore a total charge of –2. We know the sum of the charges is zero which means the charge of the Yttrium must be + 2. Yttrium requires a Roman numeral to indicate charge so the formula is written Yttrium II Fluoride.

Covalent compound – for the purposes of naming, covalent compounds are composed of two non-metals chemically bonded together. The first step is to write the name from the formula:

a) Write the name of the first element.

b) If there is more than one atom of the first element present then use a prefix to indicate the number present.

c) Write the name of the second element changing its ending to "ide".

d) Use a prefix to indicate the number of atoms of the second element present.

Example 6

NO_2 : N is Nitrogen and O is Oxygen. There is one nitrogen present so we do not use a prefix with its name. There are two oxygens present so the prefix di is used to indicate the number and we change oxygen's ending to "ide". Therefore the compound is written **nitrogen dioxide**.

Example 7

P_2Cl_5 : P is Phosphorus and Cl is Chlorine. There are two phosphorus atoms present so the prefix di is used before the name. There are five chlorine atoms present so the prefix penta is used to denote this and the ending of chlorine is changed to "ide". Therefore the compound is written diphosphorus pentachloride.

Now writing the formula from the name:

a) Write the symbol of the first element named. If there is a prefix used with the name use a subscript behind the elements name to indicate the number of atoms present of the element.

b) Write the symbol for the second element named and use a subscript to indicate the number of atoms present if there are more than one.

Example 8

Sulphur dioxide: Sulphur's symbol is S and there is no prefix meaning there is one sulphur atom in the compound. Oxide means oxygen and the prefix di indicates there are two atoms of oxygen present in the compound. Therefore the compound is written SO_2.

Acids

These are the substances that have following properties

i) Sour taste.
ii) Corrosive in nature.
iii) Turns blue litmus to red.
iv) The pH value less than 7.

The acids containing hydrogen and non-metallic element are written with a prefix "hydro" and suffix "ic". So, HI is called **hydroiodic** acid.

If the acid contains oxygen, then the prefix comes from the anion. So, H_2CO_3 is named as carbonic acid where carbonic comes from carbonate.

ACID	NAME
HCl	Hydrochloric Acid
H_3PO_4	Phosphoric Acid
HNO_3	Nitric Acid
H_2SO_4	Sulfuric Acid

Related Questions:

3. Which of the following symbols represents an ion?

   A. +2
   B. Na
   C. e^-
   D. F^-

4. What is the chemical formula for Calcium Fluoride?

   A. CaF
   B. Ca_2F
   C. CaF_2
   D. Ca_2F_2

5. The chemical formula for Yttrium II Permanganate is

   A. $Y_4(MnO)_2$
   B. $Y_2(MnO_4)$
   C. $Yt(MnO_4)_2$
   D. $Y(MnO_4)_2$

6. The chemical formula for Nitrogen dioxide is

   A. NO
   B. N_2O
   C. NO_2
   D. N_2O_2

7. What is the chemical formula for Manganese VI Phosphate?

   A. $Mn(PO)_2$
   B. $Mn(PO_4)_2$
   C. $Mn_3(PO_4)_6$
   D. $Mn(PO_4)_6$

8. What is the correct chemical name for P_2O_5 ?

   A. pentaPhosphorus diOxide
   B. diPhosphorus pentOxide
   C. Phosphorus II Oxide
   D. Phosphorus V Oxide

CHALLENGER QUESTION

9. The correct chemical name for $Na_2SO_4 \cdot 10H_2O$ is

   A. Sodium Sulphate aqueous
   B. Sodium Sulphate hydrate
   C. Sodium Sulphate decahydrate
   D. Sodium decahydrate Sulphate

10. The proper name for CaF_2 is

    A. Calcium II fluoride
    B. Calcium fluoride
    C. Calcium fluorine
    D. Cobalt fluoride

CHALLENGER QUESTION

11. The proper name for V_2O_3 is

    A. Vanadium II oxide
    B. Vanadium III oxide
    C. di Vanadium trioxide
    D. Vanadium II oxygen

12. The chemical formula for Phosphorus trichloride is

A. PCl_3
B. P_3Cl
C. P_3Cl_3
D. Cl_3P

CHALLENGER QUESTION

13. What is the charge on the metal in the compound $Pb(CO_3)_2$?

A. – 4
B. + 2
C. + 4
D. 2

14. What is the correct chemical name for HBr?

A. Hydrochloric Acid
B. Hydrobromic Acid
C. Carbonic Acid
D. Acetic Acid

Written Response

1. a) Classify the following compounds as ionic or covalent

i) CaF_2 ionic
ii) CO covalent
iii) CuCl ionic
iv) $MO_3(PO_4)_2$ ionic
v) SiF_4 covalent
vi) P_2O_5 covalent

b) Name the above compounds

Calcium flouride
Carbon monoxide
Copper I chloride

silicon tetrafluoride
di phosphorus penta oxide

TABLE OF CORRELATIONS

	Mole Concept		
	Outcomes		**Questions**
	It is expected that students will:		
Introduction	3.1.1	explain the relative nature of atomic mass	
	3.1.2	identify the unit for counting atoms, molecules, or ions as the mole	
	3.1.3	define the mole	1
	3.1.4	determine the molar mass of an element or compound	2, 7
	3.1.5	perform calculations relating the number of particles, moles, and mass	3, 4, 5, 6, 8, 9
Molar Volume of Gases	3.2.1	state Avogadro's hypothesis	10, 13
	3.2.2	determine experimentally the molar volume of a gas at room temperature and pressure	
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Mole Concept

INTRODUCTION

3.1.1 Explain the relative nature of atomic mass

3.1.2 Identify the unit for counting atoms, molecules or ions as the mole

3.1.3 Define the mole

3.1.4 Determine the molar mass of an element or compound

3.1.5 Perform calculations relating the number of particles, moles and mass

The mole is the standard unit in chemistry for communication of how much of a substance is present. Basically, a mole is the number of particles in 12 grams of carbon–12. The atomic masses of all the elements on the periodic table are relative to the mass of carbon–12 so one mole of any element will contain the same number of particles as one mole of carbon-12. Atomic masses are measured in Atomic Mass Units. An atomic mass unit (amu) is defined as having 1/12 the mass of a carbon-12 atom.

The number of particles, atoms, in 12 grams of carbon-12 and thus in one mole of any element is 6.02×10^{23}. In fact, the number of particles in one mole of any substance is 6.02×10^{23}. The substance does not need to be an element and the particles do not need to be atoms. One mole of sodium chloride contains 6.02×10^{23} molecules of NaCl. One mole of anything contains 6.02×10^{23} units of the substance in question.

This is a very important concept in chemistry as it is used to relate relative quantities of any substance to any other substance. The value 6.02×10^{23} is given a name – Avogadro's number in honour of Amedeo Avogadro. Given that the atomic masses of the elements are relative to one another and one mole of carbon-12 is equal to 12 grams we can say that one mole of any element is equal to its atomic mass is grams. Then we could also say that the molar mass of any substance is equal to its molecular weight in grams.

Example 1

Determine the molar mass of Zr.

The atomic mass of Zr, Zirconium, is 91.2 amu therefore the molar mass of Zr is 91.2 g/mol.

Example 2

Calculate the molar mass of $Zn(NO_3)_2$.

This compound has

» one atom of Zinc which has a mass of 65.4 amu

» two atoms of Nitrogen each having a mass of 14.0 amu

» six atoms of Oxygen each having a mass of 16.0 amu

Recall that a mole of each element has a mass equal to its atomic mass in grams. To find the molar mass determine the total mass of all the atoms of each element in the compound:

$$(1\,\text{Zn}\times65.4\text{ g})+(2\,\text{N}\times14.0\text{ g})+(6\,\text{O}\times16.0\text{ g})=189.4\text{ g/mol}$$

Example 3

Calculate the molar mass of $Pt_2O_3 \bullet 3H_2O$.

$$(2\,\text{Pt}\times195.1\text{ g})+(3\,\text{O}\times16.0\text{ g})+(6\,\text{H}\times1.0\text{ g})+(3\,\text{O}\times16.0\text{g})$$
$$=492.2\text{ g/mol}$$

Calculations involving particles, mole and mass

a) Moles to grams

Multiply moles by number of grams in a mole of the substance.

Example

Calculate the mass of 0.550 moles of NaOH.

$$0.550\text{ mol}\times\frac{40\text{ g}}{1\text{ mol}}=22.0\text{ g NaOH}$$

b) Grams to moles

Divide mass by number of grams in a mole.

Example

Calculate the number of moles in 34.5g HCl.

$$34.5\text{ g HCl}\times\frac{1\text{ mol}}{36.5\text{ g}}=0.945\text{ mol HCl}$$

c) Moles to particles

Multiply moles by the number of particles in a mole.

Example

Calculate the number of molecules in 1.3×10^{12} moles of H_2O.

$$1.3\times10^{12}\text{ mol } H_2O\times\frac{(6.02\times10^{23}\text{ molecules})}{1\text{ mol}}$$

$$=7.8\times10^{35}\text{ molecules } H_2O$$

d) Particles to moles

Divide particles by the number of particles in a mole.

Example

Calculate the number of moles in 3.550×10^{19} molecules of $CaCl_2$.

$$3.550\times10^{19}\text{ molecules}\times\frac{1\text{ mol}}{6.02\times10^{23}\text{ molecules}}$$

$$=5.897\times10^{-5}\text{ mol } CaCl_2$$

e) Mass to particles: these sort of calculations require a conversion first to moles and then to the second unit.

Divide mass by molar mass then multiply be number of particles in a mole.

Example

Calculate the number of atoms in 23.5 g Ag.

$$23.5\text{ g Ag}\times\frac{1\text{ mol}}{107.9\text{ g}}\times\frac{6.02\times10^{23}\text{ atoms}}{1\text{ mol}}$$

$$=1.31\times10^{23}\text{ atoms Ag}$$

f) Particles to mass:

Divide molecules present by number of molecules in a mole then multiply by molar mass of $AgNO_3$.

Example

Calculate the mass of 2.54×10^{15} molecules of $AgNO_3$.

$$2.54\times10^{15}\text{ molecules}\times\frac{1\text{ mol}}{6.02\times10^{23}\text{ molecules}}\times\frac{169.9\text{ g}}{1\text{ mol}}$$

$$=7.17\times10^{-7}\text{ g } AgNO_3$$

Related Questions:

1. A mole can be defined as
   - **A.** the number of carbon atoms in 12g of carbon-12
   - **B.** the number of carbon atoms in 12g of carbon-13
   - **C.** the number of carbon atoms in 12g of carbon
   - **D.** the number of atoms in 12g of the average mass of carbon

2. What is the mass of 2.00 moles of Sodium?
   - **A.** 11.5 g
   - **B.** 23.0 g
   - **C.** 46.0 g
   - **D.** 46 g

   $2\text{ mol} \times \frac{22.99}{1\text{ mol}} =$

3. How many moles are in 32.0g of CO_2?
   - **A.** 1400 mol
   - **B.** 1.38 mol
   - **C.** 0.73 mol
   - **D.** 0.727 mol

   $32.0\text{g} \times \frac{1\text{ mol}}{44\text{g}}$

4. How many molecules are present in 1.50 moles of NaCl?
   - **A.** 2.49×10^{-24}
   - **B.** 9×10^{23}
   - **C.** 9.03×10^{23}
   - **D.** 9.30×10^{23}

   $1.5\text{ mol} \times \frac{6.02\times10^{23}}{1\text{ mol}} =$

5. How many moles of Gold are in 4.55×10^{30} atoms of Gold?
   - **A.** 1.32×10^{-7} mol
   - **B.** 7.56×10^{6} mol
   - **C.** 7.65×10^{6} mol
   - **D.** 2.74×10^{54} mol

   $4.55\times10^{30}\text{ atoms} \times \frac{1\text{ mol}}{6.02\times10^{23}}$

6. How many molecules are in 10.0g of KCl?

$10g \times \frac{1 mol}{74.55g} \times \frac{6.02}{mol}$

A. 8.07×10^{22} molecules

B. 8.07×10^{-22} molecules

C. 4.49×10^{26} molecules

D. 2.23×10^{-25} molecules

CHALLENGER QUESTION

7. What is the mass of 1.00×10^{-6} mol $CaCl_2$?

$\times \frac{}{1 mol}$

A. 9.00×10^{-9} g $CaCl_2$

B. 1.111×10^{-4} g $CaCl_2$

C. 1.11×10^{-5} g $CaCl_2$

D. 1.11×10^{-4} g $CaCl_2$

CHALLENGER QUESTION

8. How many moles of PCl_3 is present in 1.55×10^{45} molecules of PCl_3?

A. 2.57×10^{21} moles of PCl_3

B. 9.33×10^{33} moles of PCl_3

C. 2.57×10^{11} moles of PCl_3

D. 5.27×10^{21} moles of PCl_3

9. What is the mass of 584 atoms of Au?

A. 4.92×10^{-24} g Au

B. 6.93×10^{28} g Au

C. 1.91×10^{-19} g Au

D. 1.78×10^{24} g Au

MOLAR VOLUME OF GASES

3.2.1 *State Avogadro's hypothesis*

3.2.2 *Determine experimentally the molar volume of a gas at room temperature and pressure*

3.2.3 *State the molar volume of a gas at STP*

3.2.4 *Calculate moles or mass of a gas from a given volume at STP or visa versa*

Avogadro's hypothesis states that:

Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles.

Now we can state that one mole of any gas at STP occupies 22.4 L. STP is standard temperature and pressure which is equal to 0°C and 101.3 kPa.

Example

The molar of $H_{2(g)}$ = 22.4 L at STP.

The molar volume of $CO_{2(g)}$ = 22.4 L at STP

Thus the molar volume of $H_{2(g)}$ = the molar volume of $CO_{2(g)}$ = 22.4 L at STP

Calculations

i) Moles to volume: Multiply moles present by the number of litres in a mole at STP.

Example

Calculate the volume of 3.00 mol $O_{2(g)}$.

$$3.00\,\text{mol } O_{2(g)} \times \frac{22.4\,\text{L}}{1\,\text{mol}} = 67.2\,\text{L } O_{2(g)}$$

ii) Volume to moles: Divide the volume by the number of litres in a mole.

Example

Calculate the number of moles in 9.75 L $F_{2(g)}$.

$$9.75\,\text{L } F_{2(g)} \times \frac{1\,\text{mol}}{22.4\,\text{L}} = 0.435\,\text{mol}$$

iii) Volume to mass: In this conversion it is necessary to first determine number of moles and then the mass. We cannot convert directly from volume to mass but can convert any measurable unit to moles and moles to any measurable unit.

Example

Calculate the mass of 15.9 L $Cl_{2(g)}$.

$$15.9\,L\ Cl_{2(g)} \times \frac{1\,mol}{22.4\,L} \times \frac{71.0\,g}{1\,mol}$$

$$= 50.4\,g\ Cl_{2(g)}$$

iv) Mass to volume: As with the last example we cannot convert directly to volume from mass but first must determine moles.

Example

Determine the volume occupied by 100.0g $NO_{(g)}$.

$$100.0\,g\ NO_{(g)} \times \frac{1\,mol}{30.0\,g} \times \frac{22.4\,L}{1\,mol}$$

$$= 74.67\,L\ NO_{(g)}$$

Related Questions:

10. Avogadro's hypothesis is

A. equal volumes of different gases, at the same temperature and pressure, contain the same number of particles
B. equal volumes of different gases, at different temperature and pressure, contain the same number of particles
C. proportionate volumes of different gases, at the same temperature and pressure, contain equal numbers of particles
D. equal volumes of varying gases, at similar temperatures and pressures, contain the same number of particles

11. At STP, the molar volume of any gas is:

A. 22.4 L
B. 24.4 L
C. 24.2 L
D. 22.2 L

12. What volume does 10.0 moles of $F_{2(g)}$ occupy

A. 224 L
B. 0.446 L
C. 244 L
D. 0.410 L

PERCENT COMPOSITION

3.3.1 *Compare and contrast molecular and empirical formulae*

3.3.2 *Determine the percent composition by mass from the formula of a compound*

3.3.3 *Determine the empirical formula for the compound from the percent composition by mass*

3.3.4 *Determine the molecular formula from the molecular mass and the empirical formula*

Molecular formula

The molecular formula is the true chemical formula of any given substance. The subscripts in the formula tell the actual number of atoms of each element in the compound.

Empirical formula

This is the chemical formula of a substance expressed in the lowest possible whole number ratio. The subscripts do not necessarily indicate the actual number of atoms of each element in the compound. It is possible that the empirical formula and the molecular formula of a substance are the same. For example the compound $C_3H_6O_3$ would be expressed as CH_2O.

Percent composition

This is the percent by mass of each element in the compound.

Determining percent composition

Steps

1 Determine the molar mass of the compound from its formula.

2 Determine the sum of the masses of each element by multiplying the molar mass of the element by the subscript.

3 Divide the mass attributed to each element by the molar mass of the compound and then multiply this value by 100 to get the percent.

Example

The example will be HNO_3, Nitric acid.

The molar mass of HNO_3 = 63.0g/mol. There is one Hydrogen atom in the compound with a molar mass of H = 1.0g/mol.
There is one atom of Nitrogen in the compound with a molar mass of N = 14.0g/mol. Finally there are three atoms of Oxygen in the compound having a combined molar mass of $3 \times 16.0\,g/mol = 48.0\,g$.

To find the percent composition by mass you divide the molar mass total of each element by the molar mass of the entire compound and multiply by 100.

% composition by mass of H –	$1.0\,g/63.0g \times 100 = 1.59\%$
% composition by mass of N –	$14.0\,g/63.0g \times 100 = 22.2\%$
% composition by mass of O –	$48.0\,g/63.0g \times 100 = 76.2\%$

Determination of the empirical formula (from percent composition)

Steps

1 Change the percent composition by mass into a mass for each element.

2 Determine the moles of each element from the mass.

3 Determine the ratio of the elements to one another by dividing the number of moles of each by the one with the smallest value.

4 Ensure that the ratio numbers are whole – these represent the subscripts for each element in the empirical formula.

Example

A compound is known to be composed of carbon and hydrogen. The percent composition by mass of the compound is C 80.0% and H 20.0%. Calculate its empirical formula.

Step 1

Assume there is 100g of the substance so then there would be 80.0 g C and 20.0 g H.

Step 2

$$80.0\,g\,C \times \frac{1\,mol}{12.0\,g} = 6.67\,mol\,C$$

$$20.0\,g\,H \times \frac{1\,mol}{1.0\,g} = 20.0\,mol\,H$$

Step 3

$$\frac{6.67}{6.67} = 1\,C$$

$$\frac{20}{6.67} = 3\,H$$

Step 4

So there is **1C:3 H** ratio therefore the Empirical Formula is CH_3CH_3

Example

A compound is known to be 73.72% C, 5.12% H, 4.78% N and 16.38% O. Determine the empirical formula for the compound.

Step 1

Assume there is 100g of the substance so then there would be 73.72g C, 5.12g H, 4.78g N and 16.38g O.

Step 2

$$73.12\,g\,C \times \frac{1\,mol}{12.01\,g} = 6.14\,mol\,C$$

$$5.12\,g\,H \times \frac{1\,mol}{1.01\,g} = 5.11\,mol\,H$$

$$4.78\,g\,N \times \frac{1\,mol}{14.01\,g} = 0.341\,mol\,N$$

$$16.38\,g\,O \times \frac{1\,mol}{16.00\,g} = 1.02\,mol\,O$$

Step 3

0.341 = 18

0.341 = 14.98 = 15

0.341 = 1

0.341 = 2.99 = 3

Step 4

So there is a 18 C:15 H: 1 N:3 O ratio therefore the Empirical formula is $C_{18}H_{15}NO_3$.

Example

A substance is known to be composed of 69.9% Fe and 30.1% O. Calculate the empirical formula of the compound.

Step 1

Assume there is 100g of the compound so there would be 69.9g Fe and 30.1g O.

Step 2

$$69.9\,g\ Fe \times \frac{1\,mol}{55.8\,g} = 1.25\,mol\ Fe$$

$$30.1\,g\ O \times \frac{1\,mol}{16.0\,g} = 1.88\,mol\ O$$

Step 3

$1.25 = 1$

$1.25 = 1.5$

This is not a whole number and so multiply both ratio values by a number that will make this value whole = 2.

So 1 $Fe \times 2.15 \times 2$ gives us 2 Fe: 3 O and the empirical formula will be Fe_2O_3

Expansion: Combustion Analysis

Combustion analysis is predominantly for hydrocarbons (organic compounds). The procedure requires you to burn the hydrocarbon and trap the gases produced by the burning. The information this yields will allow you to calculate the empirical formula of the substance.

Example

A 0.500 g sample of a hydrocarbon undergoes combustion in excess pure oxygen to produce 1.69 g of CO_2 and 0.346 g of H_2O.

Step 1

Determine the mass of C you have in the CO_2 and the mass of H you have in the H_2O.

$$C \rightarrow 1.69\,g \times \frac{12.011\,g\ C}{44.0098\,g\ CO_2} = 0.4612\,g\ C$$

$$H \rightarrow 0.346\,g \times \frac{2.0158\,g\ H}{18.0152\,g\ H_2O} = 0.03872\,g\ H$$

Step 2-4

Step 2-4 is the same as the standard empirical formula procedure.

$$0.4612\,g\ C \times \frac{1\,mol}{12.011\,g\ C} = 0.03840\,mol$$

$$C/0.03840 = 1$$

$$0.03872\,g\ H \times \frac{1\,mol}{1.0079\,g\ H} = 0.03842\,mol$$

$$H/0.03840 = 1$$

1C:1H therefore empirical formula is CH.

Determination of molecular formula (from empirical formula and molar mass)

Now recall that the molecular formula is the true chemical formula of the compound and not the lowest ratio like the empirical formula. In order to determine the molecular formula you require two pieces of information, the empirical formula and the molar mass of the compound.

To determine the molecular formula use the following procedure

(mass of empirical formula)$_n$ = molar mass of compound

solve for n by $n = \dfrac{\text{molar mass of compound}}{\text{mass of empirical formula}}$

Then multiply n in against the subscripts in the empirical formula to get the molecular formula.

(empirical formula)$_n$ = molecular formula

Example

The empirical formula of a substance is CH and its molar mass is 78.0g. Determine the molecular formula of the compound.

$(13.0\ g)_n = 78.0g$ therefore $n = 78.0\ g/13.0\ g = 6$

$(CH)_6 = C_6H_6$

Example

The empirical formula of a compound (glucose) is CH_2O and its molar mass is 180.0g. Determine the molecular formula of the compound glucose.

$(30.0\ g)_n = 180.0\ g$ therefore $n = 180.0\ g/30.0\ g = 6$

$(CH_2O)_6 = C_6H_{12}O_6$

Related Questions:

13. What is the mass of 50.0L $O_{2(g)}$

A. 26.6 g
B. 35.7 g
C. 65.5 g
D. 71.4 g

14. What is the percent composition of chlorine in PCl_5

A. 14%
B. 17%
C. 60%
D. 85%

CHALLENGER QUESTION

15. What is the percent composition of $KMnO_4$

A. 34.7% K, 24.7% Mn, 40.5% O
B. 24.7% K, 34.7% Mn, 40.5% O
C. 14.7% K, 44.7% Mn, 40.5% O
D. 42.7% K, 34.7% Mn, 22.6% O

16. Jacqueline analyzed a compound and found it to be 52.7% Potassium and 47.3% Chlorine. She determined the empirical formula to be

A. KCl
B. KCl_2
C. K_2Cl
D. K_2Cl_2

17. Sonia analyzed a compound and found that it contained 53.7% iron and 46.3% sulphur. She determined the empirical formula to be

A. FeS
B. Fe_2S_3
C. FeS_2
D. Fe_3S_2

18. The empirical formula of a compound is NO_2. Given its molar mass is 92.0 g, its molecular formula would be

A. N_2O_4
B. N_4O_2
C. N_4O_6
D. NO_2

19. Given the empirical formula of a compound is CH_2 and its molar mass is 70.0 g its molecular formula would be

A. C_3H_6
B. C_4H_8
C. C_4H_{10}
D. C_5H_{10}

20. A compound was analyzed and found to be 92.3% C and 7.7% H with a molar mass of 78.0g. Its molecular formula would be

A. C_6H_6
B. C_3H_5
C. CH_3
D. CH

21. What is the percent composition of the species in bold in $Al_2(SO_3)_3 \cdot 16\mathbf{H_2O}$?

A. 3.09%
B. 5.83%
C. 49.5%
D. 64.5%

22. Determine the empirical formula given for a compound that is 13.0% Mg and 87.0% Br.

A. MgBr
B. Mg_2Br
C. $MgBr_2$
D. Mg_2Br_2

CHALLENGER QUESTION

23. Determine the empirical formula for a compound given that it is 25.3% Cu, 12.9% S, 25.7% O and 36.1 % H_2O

A. $Cu(SO_4)_2 \cdot 5H_2O$

B. $CuSO_4 \cdot 5H_2O$

C. $Cu(SO_4)_2 \cdot H_2O$

D. $CuSO_4 \cdot H_2O$

MOLARITY

3.4.1 Describe molarity as (mol/L or M) as a measure of molar concentration

3.4.2 Prepare a standard solution

3.4.3 perform calculations relating mass (or moles) of solute, volume of solution and molarity

3.4.4 Calculate the resulting concentration when a given volume of a standard solution is diluted with water to a given volume

Molarity is describing the concentration of a solute in solution. It is a measure of the amount of moles of solute in solution per litre of solvent. It is applied primarily with liquid solutions and with gases. The units of molarity are mol/L or simply M. Square brackets can also be used to indicate concentration, [ ].

$$\text{Molarity (M)} \frac{\text{Moles of solute}}{\text{Litres of solvent}}$$

Sample molarity calculations

Example

Determine the molarity of a solution produced by dissolving 1.55×10^{-3} mol KCl in 2.00L water.

$$M = \frac{\text{mol}}{\text{L}} = \frac{1.55\times10^{-3}\ \text{mol}}{2.00\ \text{L}} = 7.75\times10^{-4}\ \text{M KCl mol}$$

$$= 1.55\times10^{-3}\ \text{mol}$$

Example

13.5g of NaCl is put into solution with 2.50L of water. Determine the [NaCl] or concentration NaCl or molarity NaCl.

First the mass NaCl must be converted into moles of NaCl. Secondly the moles of NaCl will be divided by the volume of water to determine the molarity of the NaCl.

$$13.0\ \text{g NaCl} \times \frac{1\ \text{mol}}{58.5\ \text{g}} = 0.222\ \text{mol NaCl then}$$

$$M = \frac{\text{mol}}{\text{L}} = \frac{0.222\ \text{mol}}{2.50\ \text{L}} = 0.0889\ \text{M NaCl}$$

This procedure can be integrated into one formula

$$MV = \frac{\text{grams}}{\text{molar mass}}$$ where M is molarity and V is Volume

So to do the above example: $M(2.50\ \text{L}) = \frac{13.0\ \text{g}}{58.5\ \text{g}}$

To solve for molarity: M = (13.0 g/58.5 g)/2.50L = 0.0889 M NaCl

Example

How many grams of $KMnO_4$ would be required to produce 0.400 L of a 0.200M $KMnO_4$ solution?

Use the formula $MV = \frac{\text{grams}}{\text{molar mass}}$

$$(0.200\ \text{M})(0.400\ \text{L}) = \frac{\text{grams } KMnO_4}{158.0\ \text{g}} \rightarrow$$

$$\text{grams } KMnO_4 = (0.200\ \text{M})(0.400\ \text{L})(158.0\ \text{g})$$

$$\text{grams } KMnO_4 = 12.6\ \text{g } KMnO_4$$

Dilution calculations with molarity

The dilution formula is $M_iV_i = M_fV_f$ where M is molarity, V is volume, i is initial and f is final.

Example

Determine the new concentration of KI if 15.0mL of water is added to 50.0mL of a 0.250M KI solution.

$M_i = 0.250\ \text{M}$

$V_i = 0.0500\ \text{L}$

$M_f = ?$

$V_f = 0.0500\ \text{L} + 0.0150\ \text{L} = 0.0650\ \text{L}$

$$(0.250\ \text{M})(0.0500\ \text{L}) = M_f(0.0650\ \text{L})$$

$$M_f = \frac{(0.250\ \text{M})(0.0500\ \text{L})}{(0.0650\ \text{L})} = 0.192\ \text{M KI}$$

V_f – the final volume was determined by combining the initial volume of the solution with the water added.

Example

Determine the volume of water that is required to dilute 1.250 L of 0.0100 M HCl to 0.00340 M HCl.

$M_i = 0.0100\,M$
$V_i = 1.250\,L$
$M_f = 0.00340\,M$
$V_f = ?$

In this problem we must first determine the final volume of the solution after is has been diluted. To then determine the volume water added we would subtract the initial volume from the final volume.

$$(0.0100\,M)(1.25\,L) = (0.00340M)\,V_f$$

$$\text{Volume added } = V_f - V_i = 3.67\,L - 1.250\,L = 2.42\,L$$

Steps for preparing and diluting solutions

Since all of the solution concentration units are based on volume of solution, as opposed to volume of solvent, it is necessary to use special volumetric glassware and proper technique to prepare a solution of precise concentration. This likewise applies to diluting solutions.

What are the correct steps for preparing 100.0 mL of a 0.100 mol/L solution of NH_4Cl ?

a) Calculate the mass of NH_4Cl required: (this will be shown using unit analysis and formula methods)

Unit Analysis

$$0.100\,mol/L \times 0.1000\,L \times 53.50\,g/mol$$
$$= 0.535\text{ g}$$

Formula

$$n = c \times v$$
$$= 0.100\,mol/L \times 0.1000\,L$$
$$= 0.535\text{ g}$$
$$m = n \times M$$
$$= 0.0100\,mol \times 53.50\,g/mol$$
$$= 0.535\text{ g}$$

b) Using a balance, measure 0.535 g of NH_4Cl into a 100 mL beaker.

c) Add approximately 40 mL of water to the NH_4Cl in the beaker and stir to dissolve. If the solute will not completely dissolve, you may go as high as 60 mL of water, but try not to exceed this limit. Do **not** remove your stirring rod from the beaker.

d) Pour the NH_4Cl solution through a funnel into a 100.0 mL volumetric flask. Wash the beaker, stirring rod, and funnel into the volumetric flask using a distilled water rinse bottle.

e) Add additional distilled water to make the water's meniscus sit on the 100.0 mL mark on the volumetric flask.

f) Put a stop in the flask and invert to mix.

Sometimes it is necessary to prepare solutions by dilution. This will be necessary when the required mass of solute is too small to measure to the correct degree of precision on your balance. It is also necessary when you have a solution of precise concentration which is too high and must be diluted.

What are the correct steps for diluting a commercially available 0.500 mol/L $Fe(NO_3)_{3(aq)}$ solution to prepare 100 mL of 0.150 mol/L necessary for an experiment?

a) Calculate the volume of concentrated solution necessary for the dilution. Students generally prefer to use the formula $c_i \times v_i = c_f \times v_f$ for this calculation, where ci and vi are the initial concentration and volume and cf and vf are the final concentration and volume.

$$c_i \times v_i = c_f \times v_f$$
$$0.500\,mol/L \times v_i = 0.150\,mol/L \times 100\,mL$$
$$v_i = \frac{0.150\,mol/L \times 100\,mL}{0.500\,mol/L} = 30.0\,mL$$

b) Use a graduated pipette to obtain the 30.0 mL of concentrated solution by pouring approximately 50 mL into a 100 mL beaker. Rinse the pipette with a small volume of solution and discard this solution. Fill the pipet using a rubber bulb to beyond the 30 mL mark and dispense 30 mL of this solution into a clean 100.0 mL volumetric flask.

c) Add distilled water to the line on the volumetric flask.

d) Put a stop in the flask and invert to mix.

Note, when diluting a concentrated acid you need to put slightly less than the required volume of water into the flask before adding the concentrated acid. Once the acid is partially diluted, the remaining water can be added. Also it is a good idea, particularly with sulfuric acid, not to invert and mix until the solution is cooled down. (You will also notice that the solution volume will decrease as the temperature drops. Volumetric flasks are calibrated at 20°C.)

Related Questions:

24. Molarity is a measure of

A. moles per mass
B. molar volume
C. molar concentration
D. molar activity

25. If 0.15mol NaCl goes into solution with 2.00L water, the molarity of NaCl is

A. 0.500M NaCl
B. 0.30M NaCl
C. ~~0.050M NaCl~~ .075M NaCl
D. 0.0050M NaCl

26. If 20.5g of NaOH is dissolved in 0.950L water, the molarity of NaOH is

A. 0.539M NaOH
B. 0.0539M NaOH
C. 0.487M NaOH
D. 0.0487M NaOH

27. If 100.0mL water is added to 150.0mL of a 2.00M HCl solution, the new molarity of HCl is

A. 1.20M HCl
B. 0.120M HCl
C. 3.00M HCl
D. 0.0300M HCl

dilution formula
MiVi MfVf
Mi 2.00M
Vi 0.15L
Mf ?
Vf 0.15 x 0.01L
= 0.25L
(2.00M)(0.150L)
Mf (0.25L)
Mf (2.00M)(0.150L) / 0.2500L
= 1.20M HCl

28. The **best** way to prepare 100 mL of 0.0400 mol/L $HCl_{(aq)}$ from a standard 0.200 mol/L $HCl_{(aq)}$ solution is to

A. use a burette and transfer 20.0 mL of 0.200 mol/L $HCl_{(aq)}$ to a 100 mL Erlenmeyer flask and then top up the solution volume to the 100 mL line with distilled water. Finally, stopper the flask and invert it several times to mix the contents thoroughly.
B. use a volumetric pipette and transfer 20.0 mL of 0.200 mol/L $HCl_{(aq)}$ to a 100 mL volumetric flask and then top up the solution volume to the calibration mark with distilled water. Finally, put a stop in the flask and invert it several times to mix the contents thoroughly.
C. measure 20 mL of 0.200 mL/L $HCl_{(aq)}$ in a 100 mL volumetric cylinder and then top up the solution volume to 100 mL with distilled water.
D. dissolve 20 mL of 0.200 mol/L $HCl_{(aq)}$ in exactly 80 mL of distilled water in a beaker and then swirl the mixture to thoroughly mix the solution.

29. Which of the following techniques is the **best** way of preparing 100 mL of a 0.0500 mol/L sodium carbonate $Na_2CO_{3(aq)}$ standard solution?

A. Carefully pipette 10 mL of a 0.0500 mol/L sodium carbonate stock solution into a 100 mL volumetric flask and then add water to the 100 mL mark.
B. Carefully pour 0.500 mol/L sodium carbonate stock solution into a 100 mL graduated cylinder to the 10 mL mark then add water to the 100 mL mark.
C. Dissolve 0.530 g of solid sodium carbonate in a 200 mL beaker, add water to the 100 mL mark, then carefully transfer the solution to a 100 mL volumetric flask.
D. In a 50 mL beaker, dissolve 0.530 g of solid calcium carbonate in deionized water, carefully transfer the solution to a 100 mL volumetric flask, then add water to the 100 mL mark.

Written Response

1. What is the mass of $N_{2(g)}$ in a $30\,m^3$ room filled with the gas at STP?

2. Determine the percent by mass composition of all the elements in $Ag(NH_3)_2F$.

3. A compound has an empirical formula of HO_2 and a mass of 92.0 g. Determine its molecular formula?

4. A compound is known to be 40.0% C, 6.7% H and 53.5% O with a molar mass of 60.0 g. Determine the compound's molecular formula.

5. Calculate the molarity of a solution produced by dissolving 7.45×10^{-2} g K_2O in 340.0 mL of water.

6. What mass of NaCl is required to produce 200.0 mL of a 0.500 M NaCl solution?

7. Billy-Joe-Jim-Bob had 47.0 L of 0.250 M Na_2SO_4. He wants to dilute the solution to 0.125 M. Calculate the volume of water he must add to the solution in order to do he dilution.

TABLE OF CORRELATIONS

	Chemical Reactions	
	Outcomes	**Questions**
	It is expected that students will:	
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Stoichiometry	4.2.1 relate the coefficients in a balanced equation to the relative number of molecules or moles (the mole ratio) of reactants and products in the chemical reaction	
	4.2.2 perform calculations involving reactions using any of the following: • number of molecules • moles • mass • gas volume at STP • solution concentration and volume	12, 21, 22, 24, 25, 26, 27
	4.2.3 perform calculations involving limiting reagent	23, WR1

Chemical Reactions

INTRODUCTION

4.1.1 Define reactants and products

4.1.2 Observe and record changes that occur during a chemical reaction

4.1.3 Describe chemical reactions in terms of the rearrangement of the atoms as bonds are broken and new bonds are formed

4.1.4 Gather experimental data that lead to the law of conservation of mass

4.1.5 Apply the law of conservation of mass to a formula equation of a reaction to demonstrate that atoms are conserved in the reaction

4.1.6 Balance formula equations of several chemical equations

4.1.7 Use subscripts to represent solids, liquids, gases, and aqueous solutions

4.1.8 Classify, predict products and write balanced equations for the following types of chemical equations: synthesis, decomposition, single replacement, double replacement, combustion, acid-base neutralization

4.1.9 Define exothermic and endothermic reactions

4.1.10 Classify reactions as exothermic or endothermic based on experimental observations

4.1.11 Relate energy changes to bond breaking and formation

4.1.12 Writing equations for chemical reactions including the energy term

A chemical reaction has two important elements, the **reactants** and the **products**. The reactants are the substances which are actually undergoing a chemical change while the products are the substances produced by the chemical change. The coefficients in a balanced chemical reaction tell us in what ratio the reactants are used and the products are formed. Also involved in any chemical reaction is an energy change. The reaction will either absorb energy or release energy. If the reaction absorbs energy it is termed an endothermic reaction and if the reaction releases energy it is termed an exothermic reaction.

The chemical reaction involves the chemical bonds between the reactants breaking apart, this requires energy, and then reforming new bonds, this releases energy, to form new substances with new properties.

A chemical reaction is like a pile of lego pieces that are taken apart and rearranged in a new order. The products are composed of the same atoms the reactants were composed of which have been rearranged in a new order thus forming new substances with new properties.

Reactants → Products or Reactants yield Products

The **Law of Conservation of Mass** describes the above concept of the reactants forming the products. It simply states that the mass of the products will be equal to the mass of the reactants that formed them. This makes sense as the atoms of the reactants are the atoms that form the products. Now this law is not strictly true as there is a slight conversion of mass into energy so in fact a tiny amount of mass is lost in the chemical reaction. There is another law called the Law of conservation of mass-energy to account for this but it is beyond the scope of the course.

The coefficients tell us the numbers of molecules or atoms are involved in the reaction. The coefficients also tell us the numbers of moles are involved in the reaction. Consider the following:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

The coefficients tell us that one mole of Zinc is reacting with two moles of Hydrochloric acid to produce one mole of Zinc Chloride and one mole of Hydrogen.

The mole ratios in balanced chemical reaction equations provide quantitative information about substances

In a balanced chemical equation, the coefficients (the numbers in front of each chemical entity) give an indication of the relative amounts (numbers of moles) of each substance in the chemical reaction.

Consider the following example:

$$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$$

The coefficients are 2, 13, 8, and 10 respectively. The equation could be read 2 moles of $C_4H_{10(g)}$ + 13 moles of $O_{2(g)}$ produces 8 moles of $CO_{2(g)}$ 10 moles of $H_2O_{(g)}$

Likewise the mole ratios will be useful in calculating numbers of moles from a starting number. The mole ratio of $C_4H_{10(g)}$ to $CO_{2(g)}$ is $\frac{2}{8}$ or $\frac{1}{4}$.

In other words there are 4 times as many moles of $CO_{2(g)}$ as there are of $C_4H_{10(g)}$.

Given that the Law of Conservation is being followed and we know the atoms of the reactants form the products it is important to be able to balance a chemical reaction. The balanced equation will have equal numbers of each atom of each element on both sides of the reaction: reactants and products. It is with coefficients that we balance the reaction. The coefficients can only be placed in front of the chemical formula representing the reactants and products.

Example

$KClO_3 \rightarrow KCl + O_2$

This reaction is not balanced, it is a skeleton reaction. The reactants contain one atom of potassium, one atom of chlorine and three atoms of oxygen while the products contain one atom of potassium, one atom of chlorine and two atoms of oxygen. The oxygen's are not balanced. To balance the oxygen's we find a common factor being six so we will use coefficients to adjust the number of oxygen's on both sides of the reaction.

$2KClO_3 \rightarrow KCl + 3O_2$

Now the oxygen's are balanced but we have unbalanced the potassium's and chlorine's. The reactant side now has two of each while the product side has one of each. We can use a coefficient of two in front of the KCl on the product side to balance this inequity.

$2KClO_3 \rightarrow 2KCl + 3O_2$

This reaction is now balanced with equal numbers of each atom on both the reactant and product sides of the reaction.

Predicting the products of a chemical reaction, based upon the reaction type

In Science 10, you learned 6 different reaction types. These types can be used to predict the products of simple chemical reactions. Reaction types and examples are given below

Formation or Single Replacement

element + element $\rightarrow$ compound

$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

Simple Decomposition

compound $\rightarrow$ element + element

$2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$

Combustion

compound + $O_{2(g)}$ $\rightarrow$ most common oxides of elements of the compound

Note: most of the time your combustion reactions will be **hydrocarbon combustions** (involving compounds of the formula C_xH_y). The format now becomes hydrocarbon + oxygen $\rightarrow$ carbon dioxide + water vapour

$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

Single Replacement

element + compound $\rightarrow$ element + compound

$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

The above example illustrates 2 points:

a) Copper can form two possible ions, Cu^{2+} or Cu^{1+}. How do you know which one to use for your product? Use the more common one as indicated on your periodic table. This is Cu^{2+}. Therefore the product is $Cu(NO_3)_{2(aq)}$ as opposed to $CuNO_{3(aq)}$.

b) You will have to use your solubility chart to determine the state of your products in solution reactions. In some questions it will be necessary to know which product is the precipitate for your calculations.

Double Replacement

compound + compound $\rightarrow$ compound + compound

$2NaI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$

Both points mentioned for single replacement reactions are likewise important here.

Other

These are reactions which do not fit the above 5 classes. At this point you have no way of predicting what will happen.

Related Questions:

1. In a chemical reaction, the product can best be described as

   A. a starting substance of the reaction
   B. a variation of the reactants
   C. a new substance with new properties
   D. a chemical variation of a reactant

2. Given a closed system and a reaction involving no gases, if you took the masses of the reactants and the products you would expect them to be

   A. close to equal
   B. different
   C. identical
   D. similar

3. The following reaction can be identified as
   $H_2O \rightarrow H_2 + O_2$

   A. acid/base neutralization
   B. combustion
   C. decomposition
   D. synthesis

4. The following reaction can be identified as
   $C_3H_8O_2 \rightarrow CO_2 + H_2O$

   A. synthesis
   B. combustion
   C. decomposition
   D. single replacement

5. The following reaction can be identified as
   $Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$

   A. single replacement
   B. double replacement
   C. acid/base neutralization
   D. water producing

6. Exothermic means the reaction

   A. absorbs energy
   B. creates energy
   C. releases energy
   D. transfers energy

7. The following reaction can be described as
   $Mg_{(aq)} + O_{2(aq)} \rightarrow MgO_{(s)} \quad \Delta H - 285.8\,kJ$

   A. synthesis and endothermic
   B. synthesis and exothermic
   C. combustion and endothermic
   D. decomposition and exothermic

8. The coefficients for the following balanced reaction are
   $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$

   A. 4,1,2,4
   B. 4,2,1,4
   D. 2,1,1,2
   C. 2,1,1,1

9. The coefficients for the following balanced reaction are
   $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

   A. 4,26,16,20
   B. 3,18,12,15
   C. 2,13,8,10
   D. 1,7,4,5

CHALLENGER QUESTION

10. The reaction of $Ba(OH)_{2(aq)}$ and $H_2SO_{4(aq)}$ produces a precipitate with the chemical formula

    A. $Ba(OH)_{2(s)}$
    B. $BaSO_{4(s)}$
    C. $H_2SO_{4(s)}$
    D. $H_2O_{(s)}$

Use the following information to answer the next two questions.

Hydrochloric acid, $HCl_{(aq)}$, is produced in the stomach to aid in the breakdown of food. Excess stomach acid, however, can irritate the lower esophagus causing heartburn. Milk of Magnesia, a suspension of $Mg(OH)_{2(s)}$ in water, can usually provide effective relief from heartburn pain.

11. A balanced equation for the reaction of Milk of Magnesia with excess stomach acid is

A. $HCl_{(aq)} + Mg(OH)_{2(s)} \rightarrow HOH_{(l)} + MgCl_{(aq)}$

B. $HCl_{(aq)} + Mg(OH)_{2(s)} \rightarrow H_2O_{(l)} + MgCl_{(aq)}$

C. $2HCl_{(aq)} + Mg(OH)_{2(s)} \rightarrow 2HOH_{(l)} + MgCl_{2(aq)}$

D. $2HCl_{(aq)} + 2Mg(OH)_{(s)} \rightarrow HOH_{(l)} + 2MgCl_{(aq)}$

12. How many moles of stomach acid are neutralized for every mole of magnesium hydroxide in Milk of Magnesia?

A. 3
B. 1
C. 2
D. 4

Use the following information to answer the next question.

The Enzyme Multiplied Immunoassay Technique (EMIT) can be used to detect the presence of a particular substance in urine, serum, or plasma. The EMIT is used in laboratories to screen for the presence of illicit drugs such as marijuana, barbiturates, amphetamines, and opiates.

13. EMIT is an example of a

A. titration analysis
B. gravimetric analysis
C. quantitative analysis
D. qualitative analysis

Use the following information to answer the next question.

Carbon dioxide levels on board the space shuttle must be controlled in order to prevent the buildup of exhaled $CO_{2(g)}$. Canisters of LiOH are commonly used for this purpose. The unbalanced reaction that occurs is

$$LiOH_{(s)} + CO_{2(g)} \rightarrow Li_2CO_{3(s)} + H_2O_{(l)}$$

14. The mole ratio of lithium hydroxide to carbon dioxide in the balanced chemical equation is

A. 1:1
B. 1:2
C. 1:3
D. 2:1

Use the following information to answer the next two questions.

Sodium hydroxide has been used to make "lye soap" for centuries. The soap forms when fatty acids are heated with $NaOH_{(s)}$. Soaps made with $NaOH_{(s)}$ usually feel hard to the touch. NaOH is also used in the chemical and pulp and paper industries.

15. The concentration of $NaOH_{(s)}$ in a soap sample was established by titrating it with $HCl_{(aq)}$. The dissolved ions present in this titration were

A. $H^+_{(aq)}, OH^-_{(aq)}$ and $Cl^-_{(aq)}$

B. $Na^+_{(aq)}, Cl^-_{(aq)}$ and $H_2O_{(l)}$

C. $H_2O_{(l)}, H^+_{(aq)}$ and $OH^-_{(aq)}$

D. $Na^+_{(aq)}, OH^-_{(aq)}, H^+_{(aq)}$ and $Cl^-_{(aq)}$

Use the following information to answer the next question.

Alchemy is the precursor of modern chemistry and metallurgy. Many alchemists tried to change baser metals into gold. Gold was believed to be the perfect metal and a key to immortality. Several "aurifaction" recipes were created. One such process, which resulted in golden crystals of tin (IV) sulfide, is shown:
#1. $KAl(SO_4)_2 \rightarrow K_2SO_4 + Al_2O_3 + SO_3$
#2. $SO_3 + 2Sn \rightarrow SnS_2 + SnO_2$
(This reaction sequence was described by the Chinese alchemist Ko Hung)

16. When equation #1 is balanced, the mole ratio of potassium aluminum sulfate, $KAl(SO_4)_2$, to aluminum oxide, Al_2O_3, is

A. 1:1
B. 2:3
C. 1:2
D. 2:1

CHALLENGER QUESTION

17. The visible exhaust coming from the chimneys of homes during the winter is

A. water vapor that has condensed
B. methane that is unburnt
C. pollutants of combustion known as smog
D. carbon dioxide combined with other greenhouse gases

Use the following information to answer the next question.

Many gases, including carbon dioxide, contribute to the effects of global warming. Photosynthesis is a process that removes carbon dioxide gas from the atmosphere. A balanced equation for photosynthesis is shown below.

$6\,H_2O + 6\,CO_2 \rightarrow C_6H_2O_6 + 6\,O_2$

18. The photosynthesis reaction can be described as

A. decomposition
B. formation
C. endothermic
D. exothermic

Use the following information to answer the next 2 questions.

Limestone buildings contain large amounts of calcium carbonate. Calcium carbonate will react readily with the hydrogen ions present in acid rain to produce water, carbon dioxide, and calcium ions. The calcium may enter the water supply and contribute to the "hard water build up" sometimes observed on kettles and pots. Below is an unbalanced chemical reaction showing this process.

$CaCO_3 + H^+ \rightarrow H_2O + CO_2 + Ca^{2+}$

CHALLENGER QUESTION

19. The states of matter for the reactants and products, from left to right, are

A. s, aq, l, g, and aq
B. s, s, l, g, and s
C. aq, aq, l, g, and s
D. s, aq, l, g, and s

20. When the above equation is balanced, the coefficients from left to right are

A. 1, 2, 1, 1, and 1
B. 1, 4, 2, 1, and 1
C. 2, 2, 1, 2, and 2
D. 1, 1, 1, 1, and 1

Use the following information to answer the next question.

Butane, $C_4H_{10(g)}$, is commonly used in cigarette lighters. When the lighter is ignited, the following reaction takes place.

$$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$$

Numerical Response

CHALLENGER QUESTION

1. Which of the following descriptions classifies the above reaction? ___ ___ ___

Write your answer in ascending numerical order

1. A physical change
2. A chemical change
3. An endothermic reaction
4. An exothermic reaction
5. A formation reaction
6. A decomposition reaction
7. A single replacement reaction
8. A double replacement reaction
9. A combustion reaction

STOICHIOMETRY

4.2.1 relate the coefficients in a balanced equation to the relative number of molecules or moles (the mole ratio) of reactants and products in the chemical reaction

4.2.2 perform calculations involving reactions using any of the following:

- *number of molecules*
- *moles*
- *mass*
- *gas volume at STP*
- *solution concentration and volume*

4.2.3 perform calculations involving limiting reagent

Stoichiometry is derived from two Greek words, stoicheion meaning element and metron meaning measure. Stoichiometry is a calculation based area of chemistry which uses balanced equations to allow you to determine relative quantities of reactants to products. You can calculate how much product could be produced given known quantities of reactants or how much reactant is required to produce given amounts of product.

A very important note for any stoichiometry question is that the chemical reaction must be described by a balanced equation. The coefficients of the balanced chemical reaction relate the numbers of moles, or molecules, of reactants to one another and to the products. Consider the following:

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

This reaction tells us 2 moles of $KClO_3$ are decomposing to produce 2 moles of KCl and 3 moles of O_2, or that 2 molecules of $KClO_3$ are decomposing to produce 2 molecules of KCl and 3 molecules of O_2. The ratios are the same as equal numbers of moles contain equal numbers of molecules. These coefficients allow us to see the molar ratios between all reactants and products and thus allow for the stoichiometric calculations.

Lets consider the reaction:
$4NH_3 \rightarrow 5O_2 \rightarrow 6H_2O + 4NO$ for purposes of the following examples.

Example

If we have 3.00mol of NH_3 reacting with excess O_2 how many moles of H_2O could be produced?

In a balanced chemical equation it is possible to do calculations between different chemical species providing you have the unit mole. To do the calculation multiply the moles of NH_3 by the molar ratio between NH_3 and H_2O. The numerator of the ratio is the substance you are starting with and the denominator of the ration is the substance you are calculating for.

The molar ratio will be $\frac{6H_2O}{4NH_3}$ therefore 3.00 mol

$$NH_3 \times \frac{6H_2O}{4NH_3} = 4.50\text{mol } H_2O$$

Example

How many molecules of O_2 are required to produce 1.500×10^{10} molecules of NO?

In this example we have a known amount of product and are asked for the amount of a reactant required to produce it. Like with moles we can do direct calculations between reactants and products using a balanced chemical equation for molecules.

$$1.500\times10^{10}\ \text{molecules NO}\times\frac{5\,O_2}{4\,NO}=1.875\times10^{10}\ \text{molecules } O_2$$

↖ molar ratio

Mass to mass stoichiometry

Utilizing mass in stoichiometric calculations involves adding 1 or two more steps to the process. Recall we can calculate between substances in the balanced equation using the unit mole. Subsequently when using masses we must calculate its mole equivalent before doing a calculation between different chemical species in the equation.

Example 1

Calculate the mass of oxygen required to react with 8.00g of ammonia according to the following reaction:

$$4\,NH_3+5\,O_2\rightarrow6\,H_2O+4\,NO$$

To do these calculations first have a plan – map out the steps you must do.

Mass NH_3 $\rightarrow$ moles NH_3 $\rightarrow$ moles O_2 $\rightarrow$ mass O_2 then calculate:

$$8.00\,\text{g } NH_3\times\frac{1\,\text{mol } NH_3}{17\,\text{g } NH_3}\times\frac{5\,O_2}{4\,NH_3}\times\frac{32\,\text{g } O_2}{1\,\text{mol } O_2}$$

$$=18.8\ \text{g}\ \ O_2\ \text{required}$$

Example 2

Calculate the mass of Sodium iodide required to produce 55.0 g of Iodine according to the following reaction:

$$K_2Cr_2O_7+6NaI+7H_2SO_4\rightarrow$$
$$Cr_2(SO_4)_3+3I_2+7H_2O+3Na_2SO_4+K_2SO_4$$

Plan: mass I_2 $\rightarrow$ moles I_2 $\rightarrow$ moles NaI $\rightarrow$ mass NaI

Solution concentration and mixed problems

These problems are similar to the mass – mass questions but they are simply using different units. The procedure is the same – convert the starting unit to moles, use molar ratio to find moles of finishing substance then convert moles of finishing substance to desired unit.

Example 1

Determine the theoretical yield of Sodium sulphate in grams if 500.0 mL of a 0.200 M NaI solution reacts with excess Potassium dichromate and Sulphuric acid according to the following reaction:

$$K_2Cr_2O_7+6NaI+7H_2SO_4\rightarrow$$
$$Cr_2(SO_4)_3+3I_2+7H_2O+3Na_2SO_4+K_2SO_4$$

Plan: molarity NaI $\rightarrow$ moles NaI $\rightarrow$ moles Na_2SO_4 $\rightarrow$ mass Na2SO4.

$$0.200\,\text{M NaI}\times0.5000\,\text{L}\times\frac{3\,Na_2SO_4}{6\,NaI}\times$$

$$\frac{14.21\,\text{g } Na_2SO_4}{1\,\text{mol } Na_2SO_4}=7.11\ \text{g}\ \ Na_2SO_4$$

Stoichiometry and limiting reagents

If one reactant is completely consumed during a chemical reaction before any of the other reactants then the reaction stops and that reactant limits the yield of the products. In these sorts of problems you first determine the limiting reagent and then calculate the theoretical yield of the product(s) in question.

$$2\,C_2H_6+7\,O_2\rightarrow4\,CO_2+6\,H_2O$$

Example 1

Determine the limiting reagent if 2.0 moles of Ethane, C_2H_6, is burned with 9.0 moles of Oxygen according to the following reaction

$$2\,C_2H_6+7\,O_2\rightarrow4\,CO_2+6\,H_2O$$

Plan: moles C_2H_6 $\rightarrow$ moles O_2 to determine limiting reagent.

$$2.0\,\text{mol } C_2H_6\times\frac{7O_2}{2C_2H_6}=7.0\ \text{mol } O_2$$

We can see that if all the Ethane is consumed it will have reacted with 7.0 mol O_2. There are 9.0 moles of the Oxygen present in the reaction which is more than is required to react with all the Ethane. The Ethane is consumed first and is thus the limiting reagent.

Example 2

Chlorine is prepared by the following reaction:

$2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

Determine the theoretical volume yield at STP of Chlorine, Cl_2, if 100.0 g of Potassium permanganate, $KMnO_4$, reacts with 750.0 mL of 6.00 M Hydrochloric acid, HCl.

First plan: find limiting reagent.

Mass $KMnO_4 \rightarrow$ moles $KMnO_4 \rightarrow$ moles HCl $\rightarrow$ molarity HCl

$$100.0\text{ g } KMnO_4 \times \frac{1\text{ mol } KMnO_4}{158.0\text{ g } KMnO_4} \times \frac{16\,HCl}{2\,KMnO_4}$$

= 5.06 mol HCl/0.7500 L = 6.75 M HCl

If all the KMnO4 were to react it would require 750.0 mL of 6.75 M HCl but the HCl in question is only 6.00 M thus the HCl is the limiting reagent. We calculate the theoretical yield of product from the limiting reagent.

OR

100.0g $KMnO_4$/158.0 g $KMnO_4$

= 0.633 mol $KMnO_4$

6.00 M HCl × 0.7500L = 4.50 mol HCl

$$\frac{2\,KMnO_4}{0.633\text{ mol}} = \frac{16\,HCl}{x} \quad x = 5.06\text{ mol HCl required}$$

We don't have 5.06 mol of HCl to react with all the $KMnO_4$ therefore HCl is the limiting reagent.

Second Plan: determine theoretical yield of Cl_2 from the limiting reagent HCl.

Molarity HCl $\rightarrow$ moles HCl $\rightarrow$ moles Cl_2 $\rightarrow$ volume Cl_2

$$6.00\text{ M HCl} \times 0.7500\text{ L} \times \frac{5\,Cl_2}{16\,HCl} \times \frac{22.4\text{ L}}{1\text{ mol HCl}}$$

$$= 31.51\text{ L HCl}$$

Related Questions:

Use the following reaction for questions 21 to 23.

$Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$

21. One mole of Cu reacts with:

A. 2 moles of Ag

B. 2 moles of NO_3^-

C. 2 moles of $AgNO_3$

D. 2 moles of $Cu(NO_3)_2$

22. If 1.5×10^{-2} mol $AgNO_3$ reacted it could produce

A. 7.5×10^{-3} mol $Cu(NO_3)_2$

B. 1.5×10^{-2} mol $Cu(NO_3)_2$

C. 3.0×10^{-2} mol Ag

D. 7.5×10^{-3} mol Cu

23. 1.8mol Cu and 2.9mol $AgNO_3$ are placed in a reaction vessel. The limiting reagent in this case is

A. $AgNO_3$

B. Ag

C. $Cu(NO_3)_2$

D. Cu

24. Consider the following reaction

$2\,KNO_3 \rightarrow 2\,KNO_2 + O_2$

If 0.500 mol KNO_3 were to react, the volume of O_2 produced would be

A. 2.24 L

B. 5.60 L

C. 0.560 L

D. 0.0446 L

25. Consider the following reaction:
$2\,HCl + Ca(OH)_2 \rightarrow 2\,H_2O + CaCl_2$
If 25.0g of $Ca(OH)_2$ were to react, the mass of H_2O produced would be

A. 0.112 g
B. 1.01 g
C. 36.4 g
D. 40.1 g

26. Consider the following reaction:
$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$
If 10.0mL of 0.200M NaCl were to react, the mass of AgCl produced would be

A. 0.0287 g
B. 0.287 g
C. 2.87 g
D. 287 g

Use the following information to answer the next question.

Silverware commonly tarnishes and needs to be polished clean. The tarnish is Silver sulphide, Ag_2S .

27. What mass of Ag_2S could be produced from 0.0255g H2S according to the following reaction

$4\,Ag + 2\,H_2S + O_2 \rightarrow 2\,Ag_2S + 2\,H_2O$

A. 0.093 g Ag_2S
B. 0.370 g Ag_2S
C. 0.185 g Ag_2S
D. 0.105 g Ag_2S

Written Response

Use the following information to answer the next question.

Aluminum chloride can be produced from the reaction of Aluminum and Chlorine according to the equation

$$2\,Al + 3\,Cl_2 \rightarrow 2\,AlCl_3$$

1. What is the limiting reagent if 10.0g of Al react with 15.0 g of Chlorine and what mass of $AlCl_3$ could theoretically be produced?

TABLE OF CORRELATIONS

Atomic Theory			
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	It is expected that students will:		
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	5.1.3	identify the atomic number for an element, using a table	
	5.1.4	calculate the number of protons and electrons in an atom or ion	3, 4, 5, 6, 7, 12, 13, 18
	5.1.5	define isotope and explain it in terms of atomic structure	8, 16
	5.1.6	calculate the number of neutrons, protons, and electrons for an atom or ion of an isotope given the mass number of the isotope and the charge of the ion	9, 14, 19
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	5.2.4	identify the following families of elements: alkali metals, alkaline earth metals, halogens, noble gases	
	5.2.5	describe some properties of the alkali metals, alkaline earth metals, halogens, and noble gases	
	5.2.6	relate noble gas stability to electron arrangement within the atom	
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Atomic Theory

INTRODUCTION

5.1.1 Describe early models of the atom

5.1.2 Describe the relative position, mass, and charge for a proton, neutron, and electron

5.1.3 Identify the atomic number for an element, using a table

5.1.4 Calculate the number of protons and electrons in an atom or ion

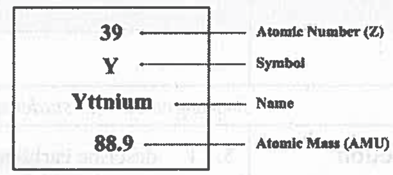
All matter consists of atoms and all atoms consist of three subatomic particles; electrons, protons and neutrons. All the protons are exactly the same and have a positive electrical charge. The neutrons are all the same and are very slightly greater on mass than the protons. Again all the electrons are identical and have a negative electrical charge. Electrons are much smaller than the other subatomic particles being about 1/1836 the size of a proton.

Mass		
Particle	**grams**	**amu**
Electron	9.109390×10^{-28}	5.485799×10^{-4}
Proton	1.672623×10^{-24}	1.007276
Neutron	1.674929×10^{-24}	1.008644

AMU is the atomic mass unit and is used for calculations involving atomic masses as the gram is much too large for convenience. 1 AMU = 1/12 of the mass of an atom of 12–C.

The nucleus is the center of the atoms shape and consists of the protons and the neutrons. The nucleus has a positive electrical charge due to the presence of the protons. The electrons are arranged on the outside of the nucleus in electron orbitals that are in different shapes at different distances out from the nucleus depending on how much energy they possess. The electrons occupy a relatively vast space given their size so most of the atom is actually empty space. Rutherford demonstrated this with his gold foil experiment. To visualize the atom think of a small pea being the nucleus in the center of your classroom. The entire class would comprise the space where the electrons orbit. It is easy to understand then that the atom is mostly empty space with a positively charged nucleus in the middle of a negatively charged space. The positive charge of the nucleus holds the negatively charged electrons in orbit about the nucleus.

The periodic table of elements contains information about the make up of the various atoms of the elements.



$$\left.\begin{matrix} AtomicMass \rightarrow 88.9 \\ AtomicNumber \rightarrow 36 \end{matrix} Y \right\}$$ This however is how it is written when not on a periodic table.

The atomic number is equal to the number of protons in the nucleus of the element. This value never changes in an element. If a proton is added or removed then the atom becomes a different element. The atomic number will also indicate the number of electrons in a neutral atom of an element. Neutral being the number of protons equals the number of electrons. Should the atom gain or lose an electron it becomes an ion and is a charged particle.

Example 1

O^{-2} this is an oxygen ion, oxide, and has a negative charge of 2. The –2 means that there are two more electrons than protons present about the atom. To determine the number of electrons present just add 2 to the atomic number, the number of protons.

Oxygen is atomic number 8 therefore electrons present are 8 + 2 = 10

Example 2

Cr^{+3} this is a chromium ion with a positive charge of + 3. The positive charge means there are three more protons than electrons. Given that the number of protons does not change ever the atom must have lost three electrons.

Chromium is atomic number 24 therefore electrons present are 24 – 3 = 21

5.1.5 Define isotope and explain it in terms of atomic structure

Atoms with an identical atomic numbers but a different mass numbers are isotopes. The atoms of an element will never vary in the number of protons they have but can vary in the number of neutrons they have. When there is a variation in the number of neutrons there will of course be a difference in atomic mass. To illustrate this we can look at Hydrogen. The majority of hydrogen atoms have one proton and no neutrons there are however two other hydrogen isotopes: Deuterium that has one proton and one neutron and then Tritium that has one proton and two neutrons.

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
Hydrogen	Deuterium	Tritium

5.1.6 Calculate the number of neutrons, protons, and electrons for an atom or ion of an isotope given the mass number of the isotope and the charge of the ion

The atomic mass is essentially equal to the sum of the masses of protons and neutrons as the mass of the electrons is relatively insignificant. The masses of the protons and neutrons can be considered to be 1 for our purposes in Chemistry 11.Therefore:

Atomic mass = number of protons + number of neutrons

Or

Atomic mass = atomic number + number of protons

From this we can determine the number of neutrons:

Number neutrons = atomic mass – number of protons (atomic number)

Example 1

Determine the number of protons, neutrons and electrons in $^{69}_{31}Ga$.

The atomic number is 31 therefore the number of protons is 31. The atom has no charge so the number of electrons must equal the number of protons, 31.

Number of neutrons = atomic mass – atomic number

Number of neutrons = 69 – 31 = 38.

Example 2

Determine the number of protons, neutrons and electrons in $^{52}_{24}Cr^{+3}$.

The atomic number is 24 therefore the number of protons is 24.

Number of neutrons = atomic mass – atomic number

Number of neutrons = 52 – 24 = 28

This ion has a charge of +3 indicating that it has three more protons than electrons. The number of electrons is then 24 – 3 = 21.

5.1.7 Calculate the average atomic mass from isotopic data

Most elements appear as mixtures of different isotopes in nature. An elements atomic mass as it appears on the periodic table is a weighted average of the naturally occurring isotopic masses of the elements isotopes. The calculation for the average atomic mass is as follows:

Atomic mass = (fraction of isotope) (mass of isotope) + (fraction of isotope)(mass of isotope) + …

It is done for as many isotopes as exists.

Example

Chlorine has two naturally occurring isotopes: 35-Cl with an abundance of 75.77% and a mass of 34.969 amu and 37-Cl with an abundance of 24.23% and a mass of 36.966% amu.

Atomic mass Cl = (fraction 35-Cl)(mass 35 - Cl) + (fraction 37-Cl)(mass 37-Cl)

Atomic mass Cl = (0.7577) (34.969 amu) + (0.2423)(36.966 amu)

Atomic mass Cl = 35.45 amu

5.1.8 Describe a simple electron arrangement for the first 20 elements

It is generally understood that the electrons of an atom orbit about the nucleus and it is the electrostatic attraction of the negative electrons to the positive protons that holds them in this orbit. In reality this is a somewhat more complicated arrangement. There are varying orbits which extend out from the nucleus. These orbits require differing amounts of energy from the electron before it will enter them. Generally the greater the amount of energy the further the orbit (or energy level) is from the nucleus.

The principle quantum number (n) represents the orbits or energy levels. Orbits are also referred to as shells by spectroscopists. These orbits, or shells, contain subshells which describe where the electron may be at any given time. The subshells are referred to here as s, p, d and f. The subshells have orbitals which we will simply say are possible positions for pairs of electrons. As the value of n increases the number of allowed orbitals increases and their size becomes larger thus allowing the electrons to be further from the atom.

Please keep in mind that is a simplified version of atomic orbital theory. So keeping this a simple as is possible.

Subshell

s type has 1s orbital and can hold 2 electrons
p type has 3p orbitals and can hold 6 electrons
d type has 5d orbitals and can hold 10 electrons

f type has 7 f orbitals and can hold 14 electrons
Orbit 1 has only s
Orbit 2 has s and p
Orbit 3 has s, p and d
Orbit 4 and up has s, p, d, and f

Each subshell in an orbit has a requisite amount of energy required for an electron to fill it. Electrons will fill the orbitals with the lowest energy requirements first unless they are excited by energy added to the system.

Energy ↑	n = 4	4 s	4 p 4 p 4 p
			3 d 3 d 3 d
	n = 3	3 s	3 p 3 p 3 p
	n = 2	2 s	2 p 2 p 2 p
	n = 1	1 s	

*** note 4s requires less energy than does 3 d.**

1s or 1↑↓ ⇐ represents on electron pair. Each orbital can hold 2 electrons that spin about the nucleus in opposite directions.

Order of electron filling

1s → 2s → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f … this continues but is beyond the requirements of Chemistry 11.

Writing electronic configurations

These are electronic configurations for elements in their ground state meaning lowest energy state. Electronic configurations show which orbits and subshells have electrons and how many are present in each. They can show relative reactivity of the elements and can be used to predict possible charge on the element if it becomes an ion.

Simply determine the number of the electrons the atom has and place them in the orbits/orbitals per the above order of increasing energy.

H has 1 electron → $1s^1$ this means orbit 1, s subshell with 1 electron.

Li has 3 electrons → $1s^2 2s^1$ this means orbit 1 has its s orbital full with 2 electrons and has 1 electron in its 2nd s orbital.

C has 6 electrons → $1s^2 2s^2 2p^2$

These can also be expressed showing the electron spins and their placement.

H 1 s ↑

Li 1 s ↑↓ 2 s ↑

C 1 s ↑↓ 2 s ↑↓ 2 p ↑ ↑ _

Mn has 25 electrons → $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$

Related Questions:

1. Who first determined that an atom contained negatively charged particles?

   A. Bohr
   B. Spray
   C. Thomson
   D. Rutherford

2. Who determined that electrons were restricted to specific orbits about the nucleus?

   A. Bohr
   B. Bond
   C. Becquerel
   D. Rutherford

3. The number of protons found in a Fe nucleus is

   A. 32
   B. 28
   C. 26
   D. 24

4. The number of electrons in orbit about a Cl atom is

   A. 17
   B. 18.5
   C. 19
   D. 35.5

5. The number of neutrons in a Mn nucleus is

   A. 79.9
   B. 30
   C. 29.9
   D. 25

6. The number of electrons in orbit about an F^- ion is

A. 8
B. 9
C. 10
D. 20

7. The number of protons found in a Y^{+2} nucleus is

A. 37
B. 39
C. 41
D. 88.9

8. Isotopes differ in mass due to different numbers of

A. electrons
B. protons
C. neutrons
D. quarks

CHALLENGER QUESTION

9. The following isotopes are found in nature:
$^{10}B = 18.9\%$ and $^{11}B = 81.1\%$
Determine the average atomic mass

A. 10.41
B. 10.5
C. 10.81
D. 11.0

10. What is the electronic configuration of Al?

A. $1s^2 2s^2 2p^6 3s^2 3p^1$
B. $1s^2 2s^2 2p^6 3s^1 3p^2$
C. $1s^2 2s^2 2p^6 3s^2 3p^3$
D. $1s^2 2s^2 2p^6 3s^2 3p^6$

11. What is the electronic configuration of O^{-2}?

A. $1s^2 2s^2 2p^2$
B. $1s^2 2s^2 2p^4$
C. $1s^2 2s^2 2p^6$
D. $1s^2 2s^2 2p^8$

12. The number of protons in As-3 is

A. 33
B. 30
C. 36
D. 15

13. The number of protons, electrons and neutrons in Rh^{+2} is

A. 45 protons, 43 electrons, 58 neutrons
B. 45 protons, 45 electrons, 58 neutrons
C. 45 protons, 47 electrons, 58 neutrons
D. 43 protons, 45 electrons, 58 neutrons

14. The relative atomic mass of the element Chlorine is 35.453 and there are two naturally occurring isotopes, ^{35}Cl and ^{37}Cl. Which of the following statements would best describe the relative abundance of the isotopes of Chlorine?

A. The isotope ^{35}Cl is about 3 times as abundant as ^{37}Cl.
B. The isotope ^{37}Cl is about 3 times as abundant as ^{35}Cl.
C. The isotopes ^{35}Cl and ^{37}Cl are in equal abundance.
D. The isotope ^{35}Cl is about 3 times as abundant as ^{38}Cl.

15. Which statement best compares a neutron and a proton?

A. Their charges are different and their masses are similar.
B. Their charges are opposite and their masses are similar.
C. Their charges are similar and their masses are different.
D. Their charges are similar and their masses are similar.

16. Consider two atoms that have slightly different masses. They could be isotopes of the same element if they have the same

A. Number of protons but different numbers of neutrons.

B. Number of neutrons but different numbers of protons.

C. Number of protons and the same number of neutrons.

D. Number of electrons but different number of protons.

17. The electronic configuration of Si is

A. $1s^2 2s^2 2p^6 3s^2 3p^2$

B. $1s^2 2s^2 2p^6 3s^2 p^2$

C. $1s^2 2s^2 2p^6 3s^1 3p^2$

D. $1s^2 2s^2 2p^6 3s^2 3p^3$

Use the following information to answer the next question.

By mass, approximately 10% of the human body is made up of hydrogen and 65% is comprised of oxygen.

18. Hydrogen has an atomic number of 1 whereas oxygen has an atomic number of 8 because each hydrogen atom has

A. one energy level whereas each oxygen atom has eight energy levels

B. one proton whereas each oxygen atom has eight protons

C. one neutron whereas each oxygen atom has eight neutrons

D. an atomic mass of one, whereas each oxygen atom an atomic mass of eight

Use the following information to answer the next question.

Carbon-14, unlike carbon-12, is a radioactive isotope of carbon. All organic matter contains a certain amount of carbon-14. After an organism dies, the carbon-14 within it begins to decay at a constant rate. Earth scientists at the University of Alberta use carbon dating to estimate the age of organic material from the last glacial era. Such information enables them to predict the location of natural resources and minerals.

19. A difference between carbon-14 and carbon-12 is that carbon-14 has

A. more protons than does carbon-12

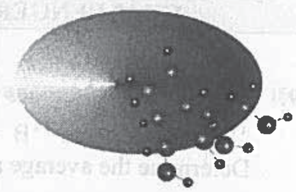
B. a lower ion charge than does carbon-12

C. more neutrons than does carbon-12

D. a higher ion charge than does carbon-12

Use the following information to answer the next question.

The Scanning Tunnelling Electron Microscope (S.T.E.M.) is a powerful tool that allows scientists to study molecules and atoms. Researchers in the Physics Department at the University of Alberta use the STEM to study the structure of thin films. In simple terms, the STEM works by mapping the surface of a material by scanning it for "bumps and dents." Bumps represent atoms, and dents represent the spaces between atoms. The structure of various molecules can be determined using this technique.



20. Which of the following statements regarding the mass of atomic structures is true?

A. Protons and electrons have a larger mass than neutrons.

B. Electrons and neutrons have a larger mass than protons.

C. Neutrons and protons have a larger mass than electrons.

D. Protons, neutrons, and electrons have approximately the same mass.

21. A potassium atom has 19 protons. A potassium atom also has 19 electrons arranged in energy levels in which of the following sequences?

A. 1, 2, 8, 8
B. 2, 8, 8, 1
C. 8, 8, 2, 1
D. 18, 1

PERIODIC TABLE

5.2.1 *Classify elements as metal, non-metal or metalloid and locate them on the periodic table*

5.2.2 *Describe the similarities and trends among elements using such properties as: melting point, ionization energy, atomic radius, chemical reactivity, ion charge and conductivity*

5.2.3 *Distinguish the ordering of elements in early periodic tables (based on atomic mass) from the ordering of elements in the modern periodic table (based on atomic number)*

5.2.4 *Identify the following families of elements: alkali metals, alkaline earth metals, halogens and noble gases*

5.2.5 *Describe some properties of the alkali metals, alkaline earth metals, halogens and noble gases*

5.2.6 *Relate noble gas stability to electron arrangement within the atom*

5.2.7 *Predict the probable electron gain or loss for elements in columns 1,2,13,15,16 and 17 to attain stability*

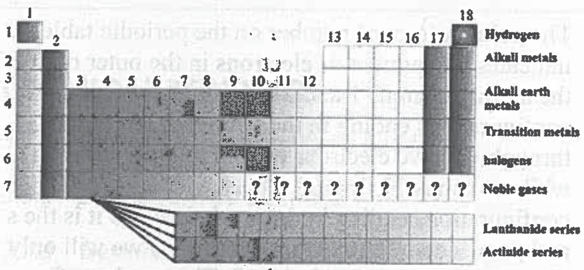
5.2.8 *Relate the observed charge of monatomic ions of metals and non-metals to numbers of electrons lost or gained*

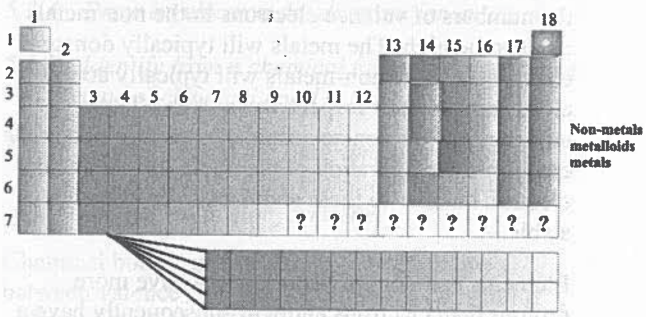
5.2.9 *Predict the characteristics of elements knowing the characteristics of another element in that family*

5.2.10 *Predict the metallic character of an element based upon its position in the table*

The modern periodic table is organized by columns (groups) and rows (periods). The groups run vertically and the periods run horizontally. The arrangement of the elements is not random and reflects the "periodic law" which states that the properties of chemical elements recur periodically when the elements are arranged from lowest to highest atomic numbers. The result is that elements of similar reactivity and with similar properties are all arranged in groups or columns. Recall that the elements of the groups all end with a similar electron configuration which accounts for the similar reactivity.

The major divisions of the periodic table are shown below with a few of their properties.





The periodic trends of the elements are shown below.

As you move across from left to right	
Atomic radius	Decreases
Metallic character	Decreases
Electro negativity	Increases
Ionization energy	Increases
reactivity	Decreases

As you move up a group	
Atomic radius	Decreases
Metallic character	Decreases
Electro negativity	Increases
Ionization energy	Increases
reactivity	Decreases

The atoms of the elements are reactive based upon their number of valence electrons. The valence electrons are the electrons found in the s and p subshells. The valence electrons that matter to the reactivity of the atom for our purposes are those electrons in the outer most orbits s and p subshells. The atom is most stable when it has a full valence thus no unfilled s and p orbitals. The noble gases are non-reactive as they have full valences. Atoms will tend to lose and gain electrons in order to become stable – have a full valence.

The column (group) number on the periodic table indicates the number of electrons in the outer orbit of the atom. Columns 1 and 2 have electronic configurations ending in the s subshell. Columns 3 through 12 have electrons ending in the d subshell while columns 13 through 18 have electronic configurations ending in the p subshell. As it is the s and p valance electrons that concern us we will only discuss columns 1, 2, 13, 15-17. The number of valence electrons in the metals tends to be low while the numbers of valence electrons in the non-metals tends to be high. The metals will typically donate electrons and the non-metals will typically accept electrons into their respective valences. Their tendency will to donate accept electrons to obtain the electronic configuration of the noble gas closest in electron number to them and thus be chemically stable.

Elements that donate electrons will have more protons than electrons and will subsequently have a positive charge. Elements that accept electrons into their valence will have more electrons than protons and will subsequently have a negative charge. If we know the column of the element we can predict its charge based upon the number of electrons it will donate or accept to become chemically stable:

Column number	Tendency	Predicted Charge
1	lose 1 electron	+1
2	lose 2 electrons	+2
13	lose 3 electrons	+3
15	gain 3 electrons	−3
16	gain 2 electrons	−2
17	gain 1 electron	−1

The numbers of valence electrons are same for all elements in any given column and the valence electrons determine how the atoms of the element will react chemically. Knowing this we can say that the elements in a column, family, all have similar chemical reactivity.

Related Questions:

22. Which element has the smallest radius based upon it periodic table position?

A. He
B. Ne
C. Ar
C. Kr

23. Which element has the smallest radius based upon it periodic table position?

A. Ca
B. K
C. Sc
D. Y

24. The ion S^{-2} has its charge because it has

A. lost 2 protons
B. gained 2 protons
C. lost 2 electrons
D. gained 2 electrons

25. Based on valence electrons, what charge would elements in group 15 have?

A. + 3
B. − 1
C. − 2
D. − 3

26. Which of the following is an Actinide element?

A. U
B. Nd
C. Rf
D. Hf

Use a periodic table to answer the next 2 questions.

27. Given that lithium (Li) reacts moderately with water and sodium (Na) reacts vigorously with water, which one of the following metals would be expected to react most vigorously with water?

A. Fe

B. Al

C. Be

D. Cs

Use the following information to answer the next question.

In 1962 at the University of British Columbia Canadian chemist Neil Bartlett synthesized a compound that most chemists thought could never be synthesized because it contained an element that was thought to be non-reactive. Bartlett's synthesis showed that although this element had very low reactivity, it could be reactive under the right circumstances.

28. Based on your knowledge of the relative reactivity of the elements in the periodic table, which of the following compounds did Bartlett most likely synthesize?

A. $RhPtF_6$

B. $KPtF_6$

C. $XePtF_6$

D. $SePtF_6$

Use the following information to answer the next question.

Internal differentiation describes a period in Earth's geological history when heavier elements moved toward Earth's centre, and lighter elements moved toward the surface. As a result of ***internal differentiation***, Earth has a core composed of heavy elements, such as iron and nickel, and an atmosphere composed of lighter elements such as nitrogen and oxygen.

A. Sulfur
B. Zinc
C. Neon
D. Gold
E. Iron
F. Sodium
G. Helium

Numerical Response

CHALLENGER QUESTION

1. The relative location of the seven elements shown above are represented in the diagram. According to the theory of *internal differentiation*, which regions could represent helium, iron, and sodium, respectively? _____, _____, _____

CHEMICAL BONDING

5.3.1 Define covalent and ionic bonding

5.3.2 Define valence electrons

5.3.3 Demonstrate a knowledge that bonding involves valence electrons

5.3.4 Draw an electron dot diagram for an atom

5.3.5 Identify from a chemical formula the probable type of chemical bond

5.3.6 Draw electron dot diagrams and structural formulae for simple molecules and ions and deduce molecular formulae

Chemical bonds are the results of interactions between valence electrons of atoms. The resulting effects of the interactions hold the atoms within a molecule together.

Typically metals interacting with a non-metal will form an ionic bond. The metal is donating electron(s) and the non-metal is accepting electron(s) resulting in positive and negative charges which hold the resulting ions together. The metal which donates the valence electrons becomes a positive ion and the non-metal which accepts the electrons becomes a negative ion. It is important to note that no two compounds will be equally ionic as there are varying degrees of electron transfer between the atoms. An extremely ionic bond has the electron completely transferred from one atom to the other but there are many ionic bonds which are "partial" as the electron is only partially transferred to from one atom to the other. It all depends on the reactivity of the atoms in question.

Covalent bonds are the result of atoms sharing a pair of electrons so that both of the atoms have some degree of "control" over the electrons. A true covalent bond has the atoms share the pair of electrons equally and this can only occur between two atoms that are identical – of the same element.

The sharing is not equal in a compound having covalent bonds. One of the atoms tends to have more control over the electron pair than the other forming a polar – covalent bond which we will still consider a covalent bond type.

The transfer of an electron form one atom to another or the unequal sharing of a pair of electrons is controlled by electronegativity. Linus Pauling defined electronegativity as "the power of an atom in a molecule to attract electrons to itself". The larger the value of the atom's electronegativity, the greater its ability to attract electrons to itself. It is the difference in the electronegativities between bonding atoms that determines the bond type.

Covalent bond electronegativity difference = 0.0-0.2
Polar covalent bond electronegativity difference = 0.2-1.7
Ionic bond electronegativity difference > 1.7

Very basically ionic and very polar covalent bonds form between metals and non-metals while weakly polar covalent and covalent bonds form between non-metals.

We can represent the bonding between atoms of some elements by using electron or Lewis dot diagrams. These diagrams are written showing the valence electrons about the atom and indicate the "pairing" of the bonding valence electrons between the atoms.

Simple Ionic compounds

» draw each atom with its valence electrons
» withdraw lone electrons from the metal to pair up with lone electrons of the non-metal and indicate charges produced

***Example* NaCl**

Sodium has one valence electron while Chlorine has seven valence electrons.

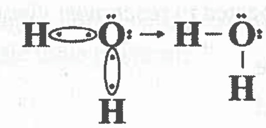
$$Na\bullet \quad \bullet\ddot{\underset{\bullet\bullet}{Cl}}: \rightarrow Na^+ \bullet \ddot{\underset{\bullet\bullet}{Cl^-}}:$$

Lewis or electron dot diagrams

The octet rules states that most atoms will tend to have 8 valence electrons about them. The exception to this is Hydrogen which only has a 1s susbshell and can only hold two valence electrons. To draw the diagram arrange the atoms symmetrically about the central atom(s) and put their valence electrons about them. Pair up the lone "bonding" electrons and redraw showing the paired up electrons as lines representing bonds.

Example

H_2O

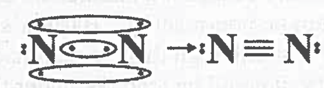


Example

SCl_2

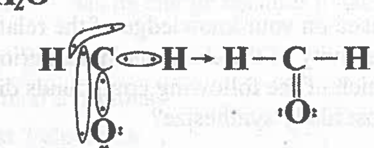
Example

N_2



Example

CH_2O



Rules for Ions and more complicated compounds

i) Count up the total number of valence electrons of the atoms of the substance in question. Add one for each negative charge and subtract one for each positive charge if the substance is an ion.

ii) Put 2 electrons between each substance to form the bond

iii) Fill the octet of the outer atoms, except hydrogen, and put the remaining electrons in pairs about the central atom.

iv)- If the central atom does not have the requisite 8 then assign a nonbonding pair about an outer atom to become a bond between it and the central atom.

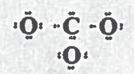
v) if the substance is an ion put square brackets about the diagram and indicate charge.

Example

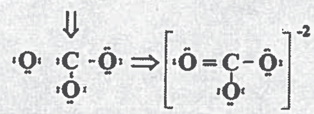
CO_3^{-2}

C O −2

$4 + (3\times6) + 2 = 24$ **valence electrons**



Carbon only has 6 electrons about it so one pair in an oxygen will become a bond



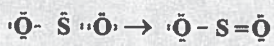
Example

SO_2

S O

$6 + (2\times6) = 18$ **valence electrons**

As S is short 2 electrons to have the octet



Related Questions:

29. A chemical bond can be described as

A. a shared pair of electrons
B. orbital electrons
C. a transfer of an electron pair
D. unequal use of two electrons

30. An ionic bond is

A. an equal sharing of an electron pair
B. the attraction between positive and negative ions
C. a metal and a non-metal joining together
D. the attraction between positive ions

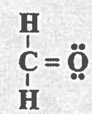
31. What type of bonding involves an electron transfer?

A. ionic bonds
B. London bonds
C. covalent bonds
D. polar covalent bonds

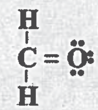
32. What type of bond is present in CF_4 ?

A. ionic
B. double
C. covalent
D. hydrogen

33. Which is the correct Lewis dot diagram for H_2CO ?

A. 

B. 

C. 

D. 

34. Which of the following has 2 lone pairs of electrons and 2 bonding electrons?

A. Iron
B. Sulphur
C. Chlorine
D. Phosphorus

35. Chemical bonds form in order to

A. transfer electron to the inner orbits
B. transfer lone electron pairs to bonds
C. have the outer orbits release electrons
D. have electrons obtain a stable configuration

36. When the atoms of Carbon an Sulphur come together to produce the compound CS_2 , the bonding electrons

A. are shared and ionic bonds are formed.
B. are shared and covalent bonds are formed.
C. are transferred and ionic bonds are formed.
D. are transferred and hydrogen bonds are formed.

Written Response

1. Use the following compounds and your knowledge of the periodic table to determine the probable formula for a Hydride of Sulphur

$KH, CaH_2, Ga_2H_6, AsH_3, H_2Se, HBr$

Organic Chemistry

		Outcomes	Questions
		It is expected that students will:	
Introduction	7.1.1	identify the multiple bonding character of carbon atoms	
	7.1.2	identify carbon as the "backbone" of organic chemistry	
	7.1.3	relate organic chemistry to products such as plastics, fuels, pharmaceutical drugs, pesticides, insecticides, solvents, synthetics	
	7.1.4	identify major sources of organic compounds	13
	7.1.5	describe a specific industrial application of organic chemistry	
Hydrocarbons	7.2.1	define hydrocarbon, alkane, alkene, alkyne, cyclic, and aromatic as they relate to organic compounds	14, 15
	7.2.2	classify a hydrocarbon as either saturated or unsaturated	1
	7.2.3	compare the geometry of single, double, and triple bonds between two carbon atoms	12
	7.2.4	compare the rotational ability in single, double, and triple bonds	
	7.2.5	name and draw structures of alkanes, alkenes, and alkynes up to C_{10}	2, 3, 8, NR3
	7.2.6	recognize and name the substituent groups methyl, ethyl, fluoro, chloro, bromo, and iodo	
	7.2.7	name and draw structures of simple substituted alkanes to C_{10}	
	7.2.8	identify cis- or trans-isomers of alkenes	7, 9, 10
	7.2.9	draw a structure of a benzene ring	4
Functional Groups	7.3.1	describe the term functional group and relate it to classes of compounds	
	7.3.2	identify a compound as an alcohol, aldehyde, ketone, ether, organic acid, ester, amine, or amide when given a structural diagram	5, 6, 11, NR2, NR4 NR1
	7.3.3	name and draw structures for simple alcohols	
	7.3.4	describe how an ester can be prepared through the reaction of an alcohol and an organic acid and how it can be detected (by its aroma)	

Organic Chemistry

INTRODUCTION

7.1.1 Identify the multiple bonding character of carbon atoms

7.1.2 Identify carbon as the ³backbone² of organic chemistry

7.1.3 Relate organic chemistry to products such as plastics, fuels, pharmaceutical drugs, pesticides, insecticides, solvents, synthetics

7.1.4 Identify major sources of organic compounds

7.1.5 Describe a specific industrial application of organic chemistry

Organic Chemistry is the branch of chemistry dealing with studies of organic compounds. As the names imply, organic compounds were thought to be the compounds responsible for maintaining and reproducing life and inorganic compounds the compounds present in minerals.

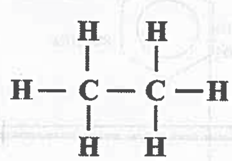
Today these definitions have broadened to the place where organic compounds are **molecular** compounds of carbon, therefore containing covalent bonds within molecules. All other compounds, including ionic compounds of carbon such as carbonates, bicarbonate, and cyanides are inorganic. Oxides of carbon such as CO and are also inorganic, even though they are molecular compounds of carbon.

Carbon is having a special ability to make bonds with itself to form chains. This property is called catenation. The property of catenation serves as a backbone of organic chemistry.

Carbon is also having a special character of forming multiple bonds i.e. forming bonds by using more than one pair of electron. For example,

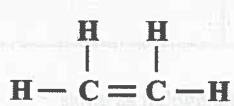
Single Bond

Ethane



Double Bond

Ethene



Triple Bond

Ethyne

H — C ≡ C — H

Carbon is present in DNA. Fossil fuels such as lignite coal and oil are the major sources of organic compounds. A wide variety of things that we use daily contains carbon-rich organic compounds, for example wool, silk, cotton, nylon, polyester. Organic compounds are also used in various types of plastics, pharmaceutical drugs, solvents and synthetics.

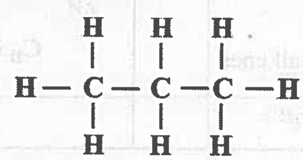
Related Questions:

7.2.1 Define hydrocarbon, alkane, alkene, alkyne, cyclic, and aromatic as they relate to organic compounds

7.2.2 Classify a hydrocarbon as either saturated or unsaturated

Hydrocarbons are the compounds that contain only carbon and hydrogen. They are classified as aliphatic or aromatic compounds. Hydrocarbons can be saturated or unsaturated.

Saturated hydrocarbon is an organic compound in which carbon atom forms single bond with another atom. For example, C_3H_8 Propane



Unsaturated hydrocarbon is an organic compound in which chain or ring are not completely satisfied and forms double or triple bonds. The unsaturated are more reactive than saturated compounds. For example, C_2H_2 Ethyne

H — C ≡ C — H

7.2.3 Compare the geometry of single, double, and triple bonds between two carbon atoms

7.2.4 Compare the rotational ability in single, double, and triple bonds

7.2.5 Name and draw structures of alkanes, alkenes, and alkynes up to C10

7.2.6 Recognize and name the substituent groups methyl, ethyl, fluoro, chloro, bromo, and iodo

7.2.7 Name and draw structures of simple substituted alkanes to C10

7.2.8 Identify cis- or trans-isomers of alkenes

7.2.9 Draw a structure of a benzene ring

Table 4-1

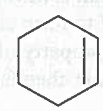
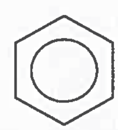
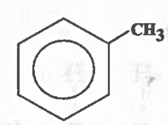
Aliphatics			
Family Name	**General Formula**	**Recognizable Characteristic**	**Examples**
Alkanes	C_nH_{2n+2}	saturated; all single bonds between carbons	2-methylpentane $CH_3-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-CH_2-CH_2-CH_3$
Alkenes	C_nH_{2n}	unsaturated; 1 double bond	2-butene $CH_3-CH=CH-CH_3$
Alkynes	C_nH_{2n-2}	unsaturated; 1 triple bond	propyne $CH\equiv C-CH_3$
Cycloalkanes	C_nH_{2n}	saturated; ring structure	cyclobutane 
Cycloalkenes	C_nH_{2n-2}	1 double bond; ring structure	cyclohexene 
Aromatics			
	C_6H_6 or $C_6H_{<6}R$	contain benzene, C_6H_6: 	methyl benzene 

Table 4-2

Family	Functional Group	General Formula	General Name	Example
alcohols	**hydroxyl:** $-O-H$	$R-O-H$	~ anol	CH_3-O-H methanol
aldehydes	**carbonyl:** $-\overset{\displaystyle O}{\overset{\|}{C}}-$	$R(H)-\overset{\displaystyle O}{\overset{\|}{C}}-H$	~ anal	$CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-H$ ethanal
ketones	**carbonyl:** (as above)	$R-\overset{\displaystyle O}{\overset{\|}{C}}-R'$	~ anone	$CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-CH_3$ propanone
carboxylic acids	**carboxyl:** (combination of carbonyl and hydroxyl) $-\overset{\displaystyle O}{\overset{\|}{C}}-O-H$	$R(H)-\overset{\displaystyle O}{\overset{\|}{C}}-O-H$	~ anonic acid	$CH_3-\overset{\displaystyle O}{\overset{\|}{C}}-O-H$ ethanoic acid
organic halides	halogen atom:	R-X, where X represents a halogen atom	NA	$CH_3-\underset{\displaystyle Br}{\underset{\|}{CH}}-CH_3$ 2-bromopropane
amines	**amine:** (see general formula)	$R-\underset{\displaystyle R''(H)}{\underset{\|}{N}}-R'(H)$ **	~ ylamine	$CH_3-\underset{\displaystyle H}{\underset{\|}{N}}-H$ methylamine
amides	**amide:** (see general formula)	$R(H)-\underset{\displaystyle O}{\underset{\|}{C}}-NH_2$	~ anamide	$CH_3-\underset{\displaystyle O}{\underset{\|}{C}}-\underset{\displaystyle H}{\underset{\|}{N}}-H$ ethanamide
esters	**ester:** (see general formula)	$R(H)-\overset{\displaystyle O}{\overset{\|}{C}}-O-R'$	~ yl ~ anoate	$H-\overset{\displaystyle O}{\overset{\|}{C}}-O-CH_2-CH_3$ Ethylmethanoate

* same as halogenated hydrocarbons
** means R or H
R refers to any hydrocarbon group – for example, CH_3-, or CH_3-CH_2-

*Hydrocarbons, C_xH_y, can be divided into the classes as shown in **Table 4-1**. (see next page)*

Abbreviated nomenclature rules, with examples will be provided here.

Number prefixes:

1 *meth*	6 *hex*
2 *eth*	7 *hept*
3 *prop*	8 *oct*
4 *but*	9 *non*
5 *pent*	10 *dec*

Aliphatics:

1. Identify the longest continuous chain of carbon atoms (longest continuous chain, containing the double or triple bond, in alk*enes* or alk*ynes*). Write the root name using the number prefix and the ending *ane*, *ene*, or *yne*, depending on whether it is an alkane, alkene, or alkyne. This will come at the **end** of the name.
2. Number the chain from the end closest to the first branch point (from the end closest to the double or triple bond in alk*enes* or alk*ynes*).
3. The position of the double bond is indicated with a number for alkanes and alkenes.
4. Name branches with the appropriate prefix and a *yl* ending.
5. If there is more than one branch, branches should be listed in al*phabetical order*.
6. Complete the name with branches in alphabetical order, numbers to locate the branches, and the root name. Always have the lowest set of numbers for alkanes. If possible, have the numbers in ascending order.

In the following examples, the long chain is identified with a shaded line.

[**Note:** In the last example, you could also find the long chain by going straight across.]

Cycloalkanes and cycloalkenes:

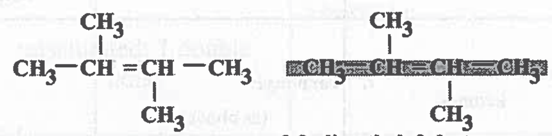
These compounds are like alkanes and alkenes except that at one point the chain is joined to itself to form a ring.

To name, count the number of carbon atoms in the ring, give it the alkane or alkene name, and add the prefix *cyclo*.

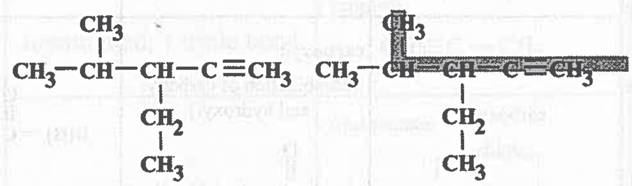
Examples:

propane

2–methylpentane



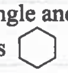
2,3–dimethyl–2–butene



3–ethyl–4–methyl–1–pentyne

Aromatics:

As referred to in the chart of section 1.2 of these notes, aromatic compounds all contain the compound benzene, C_6H_6.

Benzene is thought to be composed of a ring of 6 carbons with delocalized hybrid bonds between carbons. Each carbon could be pictured as bonding to its neighbouring carbon by a bond, which is a cross between a single and a double bond. We draw benzene as  , not as  , which would be cyclohexane.

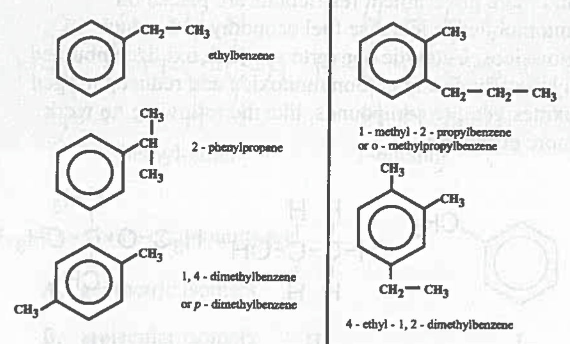
Aromatic nomenclature will be shown by example with these 3 notes:

1. Disubstituted aromatics may be named used the following prefixes:

   ortho, *o* meaning 1, 2
   meta, *m* meaning 1, 3
   para, *p* meaning 1, 4

2. Lowest set of numbers is the major priority in numbering.

3. The name *phenyl* is used when the benzene ring is labeled as a branch. Examples:



Note: even though ethyl comes first in the name by alphabetical order it gets a higher number because if it was 1, the methyls would be at 3 and 4 leading to a higher total set of numbers. Also note that the di does not count when considering alphabetical order.

Related Questions:

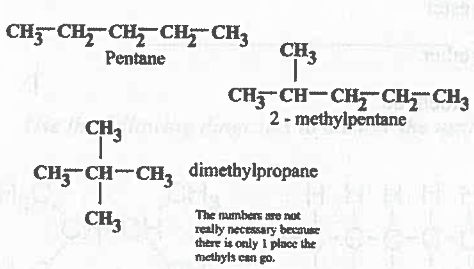
Isomerism

The phenomenon of existence of chemical compounds that have the same molecular formulae but different molecular structures or different arrangements of atoms in space is called isomerism.

Structural Isomerism

Structural *isomers* are compounds with the same chemical formula but different bonds. Other types of isomers are *optical isomers* and *geometrical isomers*, but we will only consider structural isomers here.

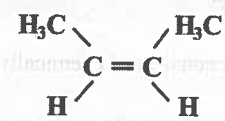
Consider the compound . It would appear initially to have many different structural isomers, but actually there are only 3. You will see this if you try to draw others and carefully find the real long chain. The 3 isomers are:



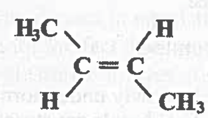
Cis-Trans Isomerism

Cis-trans isomers are the compounds with same chemical formula but different position of groups with respect to double bond.

C is Isomer



Trans Isomer



Related Questions:

7.3.1 *Describe the term functional group and relate it to classes of compounds*

7.3.2 *Identify a compound as an alcohol, aldehyde, ketone, ether, organic acid, ester, amine, or amide when given a structural diagram*

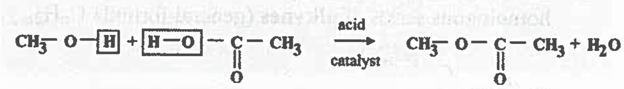
7.3.3 *Name and draw structures for simple alcohols*

7.3.4 *Describe how an ester can be prepared through the reaction of an alcohol and an organic acid and how it can be detected (by its aroma)*

(see also Table 4-2 on the previous page)

Esterification

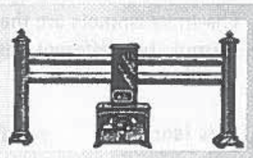
Esterification occurs between *alcohols* and *carboxylic acids*. As the alcohol and carboxylic acid react, H comes off the alcohol and OH off the acid to form water. The remaining fragments join together to make the ester. An acid catalyst and heat are required. Esters often have pleasant odours.



Related Questions:

Use the following information to answer the next question.

Paraffin is another name for alkanes. The name is derived from the Latin terms *parum* and *affino*, which together mean "chemically inactive".



1. Alkanes might be considered chemically inactive because they

   A. are capable of accepting more hydrogen

   B. are unsaturated

   C. are polyunsaturated

   D. react extremely slowly under normal conditions because their C-H bonds are strong

Use the following information to answer the next two questions.

Computers are playing an expanding role in modern chemical research. Graduates from computer colleges, such as CDI College of Business and Technology, can find opportunities to work in this field if they have a background in chemistry. These graduates play an important role in creating software applications for controlling chemical devices and creating databases.

2. A CDI student has designed a program to enumerate the members of both alkenes and alkynes. Her program indicates that the third member of the homologous series of **alkenes** (general formula C_nH_{2n}) is

   A. C_3H_6

   B. C_4H_8

   C. C_5H_{10}

   D. C_6H_{12}

3. Her program would indicate that the third member of homologous series of **alkynes** (general formula C_nH_{2n-2}) is

   A. C_3H_4

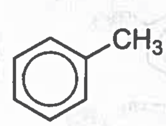
   B. C_4H_6

   C. C_5H_{10}

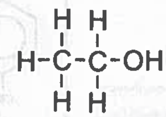
   D. C_6H_{12}

Use the following information to answer the next three questions.

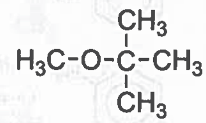
Gasoline must meet more stringent requirements as more and more government restrictions are placed on automobiles to increase fuel economy and reduce emissions. Catalytic converters, which oxidize unburned hydrocarbons and carbon monoxide and reduce nitrogen oxides, require compounds, like the following, to work more efficiently.



I



II



III

4. Compound I is

   A. phenol

   B. benzene

   C. methyl-benzene

   D. cyclohexene

5. Compound II is

   A. isoethane

   B. ethanol

   C. ethane

   D. ethene

6. Compound III is

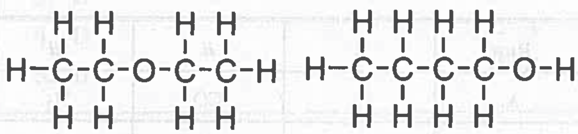
   A. an alcohol

   B. an ester

   C. an ether

   D. an aldehyde

Use the following diagrams to answer the next question.



diethyl-ether 1-butanol

7. The above compounds are

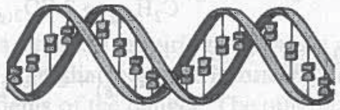
A. geometric isomers

B. molecular isomers

C. structural isomers

D. optical isomers

Use the following information to answer the next question.



Recently, it was discovered that carbon-containing molecules can survive the frigid (50° K and lower) high vacuum conditions of outer space. In laboratory tests, DNA has been shown to survive these interstellar conditions. The unusual molecules C_3 and C_5 have been identified by astronomers who study microwave radiation from space.

8. If these chains were saturated with hydrogen, they would be known as, respectively

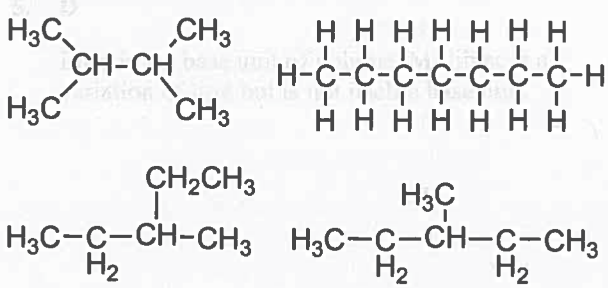
A. propane and pentane

B. propane and butane

C. butane and pentene

D. propane and pentene

Use the following diagrams to answer the next question.



9. The number of different structural isomers represented above is

A. 1

B. 2

C. 3

D. 4

Use the following information to answer the next question.

Louis Pasteur discovered that two crystals of sodium ammonium tartrate, present in equal amounts in a mixture, were alike except for the fact that they were non-superimposable mirror images of each other.

10. Pasteur was describing molecules we now classify as

A. polymers

B. monomers

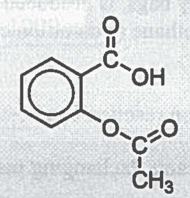
C. isomers

D. hydrocarbons

Use the following information to answer the next question.

Acetylsalicylic acid (ASA), found in aspirin, is a derivative of the active ingredient in willow bark tea, namely salicylic acid. Interestingly enough, Compound W™, a proprietary wart remover, is a concentrated aqueous salicylic acid solution. In aspirin, however, it was found necessary to modify salicylic acid to ASA to avoid the burning sensation salicylic acid causes when taken internally.

The structure of aspirin is



11. Two functional groups present in aspirin are

A. amide and amine

B. carboxylic acid and alcohol

C. ester and carboxylic acid

D. ketone and carboxylic acid

Use the following information to answer the next four questions.

1. propanol	2. propanal	3. propene
4. methyl propenoate	5. propanamide	6. chloropropane
7. propanoic acid	8. propyne	9. propanone

Numerical Response

1. The number of Cs, Hs, Os, and Ns, respectively, in the only nitrogen containing compound in the list above is ___, ___, ___, and ___ .

Numerical Response

2. The first four compounds (reading from 1-9) that contain a carbonyl group are ___, ___, ___, and___.

Numerical Response

3. In the list above, the alkyne, alkene, alcohol (alkanol) and aldehyde (alkanal), respectively, is ___, ___, ___, and___ .

Numerical Response

4. The molar mass of the haloalkane in the list is ________. (Record your answer to three digits.)

Use the following information to answer the next question.

Ethanol, used in the manufacture of poly(ethene), the plastic in grocery bags, is produced by the elimination of hydrogen from ethane (a common component of natural gas in Alberta).

12. An elimination reaction can be characterized by

A. two reactants exchanging parts to give two products

B. one reactant reacts so as to lose atoms, usually as a small molecule

C. two reactants combining to form a single product

D. a reorganization of bonds and atoms to give a different product with the same chemical formula

13. The complete combustion of methane would require _*i*_ and would primarily yield _*ii*_ and _*iii*_. The blanks in the above statement are filled by

Row	*i*	*ii*	*iii*
A.	O_2	CO	H_2
B.	H_2	CO	H_2O
C.	O_2	CO	H_2O
D.	O_2	CO_2	H_2O

14. The reaction equation that correctly represents the combustion of butane is

A. $2\ C_4H_{10(g)} + O_{2(g)} \rightarrow 2\ C_4H_{8(g)} + 2H_2O_{(g)}$

B. $2\ C_4H_{10(g)} + 3\ O_{2(g)} \rightarrow C_2H_{6(g)} + 2\ CO_{2(g)} + 2\ H_2O_{(g)}$

C. $2\ C_4H_{10(g)} + 9\ O_{2(g)} \rightarrow C_2H_{4(g)} + 6\ CO_{2(g)} + 6\ H_2O_{(g)}$

D. $2\ C_4H_{10(g)} + 13\ O_{2(g)} \rightarrow 8\ CO_{2(g)} + 10\ H_2O_{(g)}$

Use the following information to answer the next question.

Naphthalene ($C_{10}H_{10(s)}$) is an aromatic compound that is commonly used in mothballs.

15. The complete combustion reaction of naphthalene is represented by

A. $C_{10}H_{10} + O_2 \rightarrow CO_2 + H_2O$

B. $C_{10}H_{10} + 12\frac{1}{2}O_2 \rightarrow 10\ CO + 5\ H_2O$

C. $C_{10}H_{10} + 12\ O_2 \rightarrow 10\ CO_2 + 5\ H_2O$

D. $2\ C_{10}H_{10} + 25\ O_2 \rightarrow 20\ CO_2 + 10\ H_2O$

Practice Examination One

PRACTICE

EXAM

Practice Examination One

1. Which of the following is not an SI unit?

   A. s
   B. mL
   C. g
   D. m

2. How many significant digits are present in 56070?

   A. 4
   B. 5
   C. 3
   D. 2

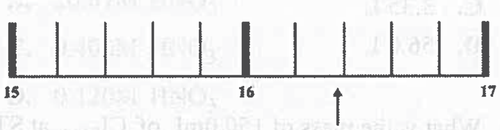
3. Considering significant figures, the correct answer for $\left(\dfrac{6.02\times10^{23}}{0.02457}\right)$ is

   A. 1.48×10^{22}
   B. 2.450×10^{25}
   C. 2.45×10^{24}
   D. 2.45×10^{25}

4. If a chemical reaction produces 0.0150g of CO_2 in 2.00 min, the rate would best be described in which unit?

   A. g/s
   B. g • min
   C. g/s^{-1}
   D. g^{-1}/s

5. What is the measure indicated on the scale below?



   A. 16.3
   B. 16.29
   C. 16.35
   D. 15.16

6. Matter can best be defined as anything that has ____________ and occupies __________.

   A. mass, area
   B. volume, mass
   C. atoms, volume
   D. mass, volume

7. Matter can be identified by its ___________.

   A. properties
   B. shape
   C. molecular structure
   D. coreolis effect

8. New substances with new properties produced describes

   A. a chemical change
   B. a physical change
   C. vaporization
   D. sublimation

9. Which of the following cannot chemically be separated into simpler substances?

   A. element
   B. compound
   C. solution
   D. polyatomic ion

10. The correct name for $KClO_4$ is

    A. Potassium Chlorate
    B. Potassium I Perchlorate
    C. Potassium Chlorine tetraoxide
    D. Potassium Perchlorate

11. The correct name for $Mg_2(PO_4)_2$ is

    A. Magnesium Phosphate
    B. Manganese Phosphate
    C. Magnesium II Phosphate
    D. Manganese II Phosphate

12. The correct name for SCl_2 is

A. Sulphur Chloride
B. diSulphur Chloride
C. Sulphur diChloride
D. Sulphur II Chloride

13. The correct name for PbO_2 is

A. Lead Oxide
B. Lead II Oxide
C. Lead dioxide
D. Lead IV Oxide

14. The correct chemical formula for Ammonium Oxalate is

A. $NH_4C_2O_4$
B. $(NH_4)_2C_2O_4$
C. $(NH_4)C_2O_4$
D. $NH_4(C_2O_4)_2$

15. The correct chemical formula for diPhosphorus pentachloride is

A. P_2O_5
B. P_5Cl_2
C. P_2C_5
D. P_2Cl_5

16. The unit used to count the numbers of atoms, molecules or ions present is

A. the mole
B. the atomic number
C. the isotope
D. the joule

17. The molar mass of Yttrium is

A. 88.9 g
B. 88.9 AMU
C. 39 g
D. 39 AMU

18. The molar mass of $Mn_2(CO_3)_3$ is

A. 289.8 g
B. 289.8 AMU
C. 239.9 g
D. 239.9 AMU

19. How many moles are in 32.0 g CH_2O ?

A. 960 mol
B. 0.938 mol
C. 1.07 mol
D. 0.107 mol

20. How many molecules are present in 0.150 moles of RbCl?

A. 9×10^{22}
B. 2.49×10^{-23}
C. 9.03×10^{22}
D. 9.03×10^{22}

21. How many molecules are in 1.63×10^{-3} g of KCl?

A. 7.32×10^{-22} molecules
B. 7.32×10^{22} molecules
C. 2.02×10^{-25} molecules
D. 1.32×10^{19} molecules

22. What volume would 75.0 mol $NO_{(g)}$ occupy at STP?

A. 1680 L
B. 168.0 L
C. 3.35 L
D. 56.0 L

23. What is the mass of 150.0mL of $Cl_{2(g)}$ at STP?

A. 26.67 g
B. 238.5 g
C. 475.4 g
D. 0.4754 g

24. Phentylfloroform is 57.54% C, 3.45% H and 39.1% F. What is the empirical formula of the compound?

A. $C_7H_5F_3$

B. $C_3H_2F_1$

C. $C_5H_7F_3$

D. $C_7H_3F_3$

25. Molarity is a measure of

A. molar ability

B. molar volume

C. moles per mass

D. molar concentration

26. The molarity of a solution prepared by dissolving 0.350 mol KI in 2.75 L of water is

A. 1.27 M KI

B. 0.127 M KI

C. 0.963 M KI

D. 9.63 M KI

27. The mass of RbCl required to prepare 2.000L of a 0.4500 M RbCl solution is

A. 134.4 g

B. 7.34×10^{-3} g

C. 108.9 g

D. 18.9 g

28. If 150.0mL water is added to 250.0 mL of a 1.00 M HNO_3 solution the new molarity of HNO_3 is

A. 0.625M HNO_3

B. 26.67M HNO_3

C. 0.400M HNO_3

D. 0.120M HNO_3

29. How do the masses of reactants and products relate if a reaction occurs in a closed system?

A. they are equal

B. they are proportionate

C. the mass of products decreases

D. the masses do not change

30. The following reaction can be identified as $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

A. combustion

B. single replacement

C. synthesis/combination

D. double replacement

31. The following reaction can be identified as $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

A. combustion

B. single replacement

C. acid/base neutralization

D. double replacement

32. The following reaction can be identified as $KClO_3 \rightarrow KCl + O_2$

A. acid/base

B. synthesis

C. composition

D. decomposition

33. The coefficients for the following reaction to be properly balanced would be $Zn + HCl \rightarrow ZnCl_2 + H_2$

A. 1,2,1,1

B. 2,4,2,2

C. 2,4,2,1

D. 1,2,2,1

34. The following reaction can be identified as $H_2 + Cl_2 \rightarrow 2HCl + 185\ kJ$

A. synthesis and exothermic

B. synthesis and endothermic

C. decomposition and exothermic

D. single replacement and endothermic

35. Consider the following reaction:
$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(l)}$
How many moles of O_2 are required to produce 6.0 moles of NO?

A. 4.8 moles
B. 7.5 moles
C. 6.0 moles
D. 5.0 moles

36. Consider the following reaction:
$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$
There is 16.0 g Cu and 42.0g $AgNO_3$ present in a reaction vessel. Which reactant is the limiting reagent?

A. Ag
B. Cu
C. $AgNO_3$
D. $Cu(NO_3)_2$

37. How many neutrons are present in the nucleus of a Neodymium atom?

A. 84
B. 60
C. 84.2
D. 90

38. The number protons in the nucleus of an Iridium atom are

A. 115
B. 192
C. 77
D. 78

39. How many electrons are there in orbit about the P^{-3} ion?

A. 18
B. 15
C. 31
D. 16

40. What is the electronic configuration of Titanium?

A. $1s^2 2s^2 2p^6 3s^2 3p^6$
B. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
C. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
D. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

41. What is the electronic configuration of Mn^{+4}?

A. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
B. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
C. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
D. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^1$

42. Column 14 elements have an electronic configuration that ends in

A. s^2p^2
B. s^2p^3
C. s^2p^1
D. s^2d^2

43. Which of the following elements is a member of the Alkaline Earth metals?

A. Sr
B. Rb
C. Y
D. Xe

44. Which of the following is a property of a Halogen?

A. conducts electricity
B. very soft and malleable
C. most are hard and dense
D. diatomic in gaseous phase

45. Which element will tend to gain 3 electrons to have a noble gas configuration?

A. C
B. P
C. S
D. Cl

46. What type of bonding involves a partial electron transfer?

A. ionic
B. covalent
C. polar covalent
D. London's

47. What type of bond is present in PCl_3

A. covalent
B. ionic
C. double
D. hydrogen

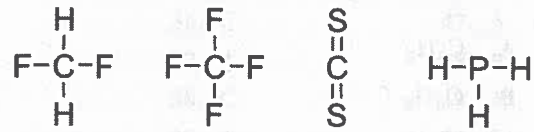
Use the following information to answer the next question.

Chlorine (Cl_2) is used in great quantity as a bleach for the pulp and paper industry and as a disinfectant for municipal water supplies. Iodine is another chemical commonly used for water disinfection on a smaller scale. The boiling point of chlorine is −34.6° C while the boiling point of iodine is 184° C.

48. The electrons involved in the bonding within chlorine molecules are

A. completely transferred from one atom to another
B. equally shared
C. unequally shared
D. the inner shell electrons only

Use the following diagram to answer the next question.



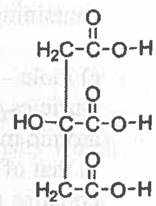
I II III IV

49. The polar molecule is

A. I
B. II
C. III
D. IV

Use the following information to answer the next question.

Acids have a characteristic sour taste. Foods such as lemons, oranges, and tomatoes contain citric acid which is, in part, responsible for their particular flavours. The structure of citric acid is to the right.



50. The functional groups within citric acid are

A. ketones and alcohols
B. aldehydes and carboxylic acids
C. esters and alcohols
D. alcohols and carboxylic acids

Use the following information to answer the next question.

Carvone is a compound whose molecules have **two structures** with the same formula ($C_{10}H_{14}O_2$) One geometric arrangement provides the characteristic odour of caraway seeds, while the other form gives the pleasant aroma of spearmint oil.

51. These two forms of carvone can best be described as

A. inorganic compounds
B. monomers
C. polymers
D. isomers

52. The third member of the homologous series of cycloalkanes (general formula C_nH_{2n}) is

A. C_3H_6
B. C_4H_8
C. C_5H_{10}
D. C_6H_{12}

Written Response

1. Define **a)** atom – answer = the smallest possible unit of an element possible which will still retain the properties of the element.

b) solution – answer = a homogenous mixture containing 2 or more substances.

c) mole – answer = the mole is the number of particles of Carbon in 12 grams of C-12. The atomic masses of all other elements are relative to that of C-12 so one mole of any element contains the same numbers of particle as one mole of C-12.

2. State Avogadro's hypothesis – answer = Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles.

3. What is the percent composition by mass of H_2CO_3?

4. Determine the molecular formula of a compound given its empirical formula is C_2H_6O and its molar mass is 138.0 g

5. Consider the following reaction
$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$
What mass of Carbon is required to react completely with 100.0 g of the Iron III Oxide?

6. Magnesium has three naturally occurring isotopes: 23-Mg = 23.804 amu and has an abundance of 11.01%, 24-Mg = 23.985 amu and has an abundance of 79.99% and 25-Mg = 24.986 amu and has an abundance of 10.00%. Determine the atomic mass of Magnesium as it appears on the periodic table.

7. Showing all work draw the electron (Lewis) dot diagram for
a) $CSCl_2$
b) PO_4^{-3}

8. Jacqueline dilutes the concentration of Chloride ions present in 125.0 mL of a 0.150 M NaCl solution by adding 75.0mL of 0.195M NaI. She has diluted the concentration of chloride ions but has she diluted the concentration of sodium ion? Show all work.

9. Draw the condensed structure for 4- ethyl – 3 – methyl – 2- heptene.

10. Name the compound

Solutions – Practice Examination One

1. B	12. C	23. D	34. A	45. B
2. A	13. D	24. A	35. B	46. C
3. D	14. B	25. D	36. C	47. A
4. A	15. D	26. B	37. A	48. B
5. C	16. A	27. C	38. C	49. A
6. D	17. A	28. A	39. A	50. D
7. A	18. A	29. A	40. D	51. D
8. A	19. C	30. A	41. A	52. C
9. A	20. C	31. D	42. A	
10. D	21. A	32. D	43. A	
11. A	22. D	33. A	44. D	

1. **B**

   Litre is the base SI unit of volume and mL is variation on it.

2. **A**

   There are 4 significant figures present in **56070** – the final zero is a place value and is not indicated to be significant as there is no decimal place present. The zero between the 6 and the 7 is significant.

3. **D**

   In a multiplication or division involving significant figures the value with fewest sig figs is used to determine the number of sig figs in the answer. 6.02×10^{23} has 3 sig figs while 0.02457 has 4 so the answer will have 3.

4. **A**

   To determine rate we would divide the mass by the time and end up with a rate in g/min or g/s. The latter is more correct as the SI unit of time is the second.

5. **C**

   We know the point is between 16.2 and 16.4 and appears to be more than halfway between 16.2 and 16.4 so is at least 16.3. We cannot see how far past 16.3 the measure is as there is no marker so we can estimate and say half way in this case thus 16.35. The five is a somewhat certain digit and is considered to have significance but us recognized as not precise.

6. **D**

   All matter is considered to have mass and occupy volume.

7. **A**

   All matter will have defining properties such as melting point, boiling point and density which can be used to identify it.

8. **A**

   A chemical change alters the molecular composition of the matter which results in new substances which will have their own distinct properties.

9. **A**

   An element consists of only one type of atom and therefore cannot be chemically reduced into any simpler substance. Compounds, solutions and polyatomic ions all consist of more than one type of atom and thus can be chemically altered into a simpler substance.

10. **D**

    This is a metal – non-metal compound thus is an ionic compound. K is the symbol for Potassium which is always +1 and requires no roman numeral to indicate its charge. ClO_4^- is the Perchlorate ion and as no roman numerals are required the names are simply written together **Potassium Perchlorate**.

11. A

This again is an ionic compound. Mg is the symbol for Magnesium which always has a charge of +2 and thus does not require a roman numeral in the name to indicate its charge. PO_4^{-3} is the phosphate ion and as no roman numerals are required the names are simply written together Magnesium Phosphate.

12. C

This is a covalent compound as it is a composed of two non-metals. There is only a single Sulphur present so no prefix is required to indicate number of them. There are two chlorides present so a prefix, di, is required to indicate the number of chloride atoms in the compound – **Sulphur dichloride**.

13. D

Pb is Lead which does not have a consistent charge so a roman numeral is required to indicate its charge in the compound. O^{-2} is the oxide ion of which there are 2 in the compound. This gives us a total negative charge of 4 and as the sum of the charges in an ionic compound is zero there must be a positive charge of 4. There is only one Lead atom in the compound so it must have a charge of +4 – **Lead IV Oxide**.

14. B

This is an ionic compound despite NH_4^+ not being a metal – it is a positively charged polyatomic ion. Ammonium has a +1 charge while Oxalate has a -2 charge and as the charge sum of the compound equals zero we require two Ammonium ions to have a +2 to match the -2. There is a subscript 2 placed after the Ammonium to indicate that there are two of them present in the compound - $(NH_4)_2C_2O_4$.

15. D

This is a covalent compound as it is comprised of non-metals. The prefixes indicate the number of each atom present in the compound thus we would have 2 phosphorus atoms and 5 Chlorine atoms in the compound. The subscripts written indicate the number of atoms of each element that precedes them - P_2Cl_5.

16. A

The mole is the arbitrary unit used to relate the relative masses and quantities of atoms, molecules and ions to one another.

17. A

The molar mass of an element is equal to its atomic mass in grams. The atomic mass of Yttrium is 88.9 AMU so it molar mass would be **88.9 g**.

18. A

The molar mass is equal to the sum of the atomic masses of all the atoms present in the compound.

$$Mn(2\times54.9\,g)+C(3\times12.0\,g)+O(9\times16.0\,g)$$
$$=289.8g$$

19. C

This is a unitary conversion form mass to moles.

$$32.0\,g\ CH_2O\times\frac{1\,mol}{30.0\,g\ CH_2O}$$
$$=1.07\,mol\ CH_2O$$

20. C

Unitary conversion from moles to molecules

$$0.150\,mol\ RbCl\times\frac{6.02\times10^{23}\ molecules}{1\,mol}$$
$$=9.03\times10^{22}\ molecules\ RbCl$$

21. A

Unitary conversion from mass to molecules. To do this we must first convert the mass to moles and then the moles can be converted to molecules

Plan: mass KCl $\rightarrow$ moles KCl $\rightarrow$ molecules KCl

$$1.63\times10^{-3}\ g\ KCl\times\frac{1\,mol}{74.6\,g\ KCl}\times\frac{6.02\times10^{23}\ molecules}{1\,mol}$$
$$=8.07\times10^{22}\ molecules\ KCl$$

22. D

Unitary conversion from moles of NO to Litres at STP

$$75.0\,\text{mol NO} \times \frac{22.4\,\text{L}}{1\,\text{mol}}$$

$= 1680\text{L NO}$

23. D

This is a unitary conversion between mass and volume of a substance. Mass cannot directly be converted into volume and we must go through the unit mole. Any of the measures can be converted into moles and moles can be converted into any measure.

First change the volume into Litres – 150.0mL = 0.1500L

Plan $\text{L Cl}_2 \rightarrow \text{mol Cl}_2 \rightarrow \text{mass Cl}_2$

$$0.1500\,\text{L Cl}_2 \times \frac{1\,\text{mol}}{22.4\,\text{L}} \times \frac{32.0\,\text{g}}{1\,\text{mol}} = 0.4754\ \text{g Cl}_2$$

24. A

Step 1: we assume the mass of the compound analyzed to be 100.0g so that the percent by mass of the elements is easily determined: 57.45g C, 3.45g H and 39.1g F. Then we follow the procedure for finding empirical formula: grams → mol: find ratio and ratio is empirical formula.

Step 2: $57.45\,\text{g C} \times \frac{1\,\text{mol}}{12.0\,\text{g}} = 4.788\,\text{mol}$

Step 3: C/2.06 = 2.33

Step 4: ×3 = 7
All ratio numbers are multiplied by 3 to make them whole numbers as the C ratio value was 2.33 and the H ration number is 1.67. The result is 7C:5H:3F and empirical formula is $C_7H_5F_3$.

25. D

Molarity is a measure of concentration and is equal to the numbers of moles in a given volume M = mol/L. Molar concentration is another way of stating molarity

26. B

Molarity = mol/L so to determine the molarity of the KI we simply divide the mol KI by the volume it is dissolving in.

$$\text{Molarity} = \frac{\text{mol}}{\text{L}} = \frac{0.350\,\text{mol}}{2.75\,\text{L}}$$

$= 0.127\text{M KI}$

27. C

We use the formula $MV = \frac{\text{mass}}{\text{g/mol}}$

$$(0.4500\text{M RbCl})(2.000\,\text{L}) = \frac{\text{mass RbCl}}{121.0\,\text{g/mol}}$$

$\rightarrow$ mass RbCl = (0.4500)(2.000)(121.0)

mass RbCl = 108.9 g

28. A

When volumes are mixed the molarity of a solution will change. It changes because the molarity is volume dependent as it is a measure of the number of moles in a given volume. To determine the new molarity we us the dilution formula, $M_iV_i = M_fV_f$ where M is molarity, V is volume, i is initial and f is final. This basically finds the moles of the substance and divides it by the new volume.

$M_i = 1.00\,\text{M}$

$V_i = 1.2500\,\text{L}$

$M_f = ?$

$V_f = 2.500\,\text{L} + 0.1500\,\text{L}$

$= 0.4000\,\text{L}$

$(1.00\,\text{M})(0.2500\,\text{L}) = M_f(0.4000\,\text{L})$

$$M_i = \frac{(1.00\,\text{M})(0.2500\,\text{L})}{(0.4000\,\text{L})}$$

$= 0.625\ \text{M HNO}_3$

29. A

The reaction is in a closed system meaning nothing is entering or leaving the reaction vessel. It involves no gases being consumed or produced so the mass of the reaction vessel and its contents can be taken before and after the reaction accurately. We cannot measure the mass of a gas directly. A chemical reaction involves the reactant atoms rearranging to form the products so the mass of the reactants should, given the Law of Conservation of Mass, equal the mass of the products.

30. A

This reaction is a combustion reaction. They can be identified by a Hydrocarbon compound reacting with Oxygen to form Carbon Dioxide and Water. These components are always present in a combustion reaction even though it may contain other substances as well.

31. D

This is a double replacement reaction. They can be recognized as the reactants are both ionic compounds and the products are formed by the metals of the reactants switching places with the respective non-metals. Note in the above reaction that the Ag^+ is originally with the NO_3^- as a reactant and finishes with the Cl^- as a product while the Na^+ starts with the Cl^- as a reactant and finishes with the NO_3^- as a product.

32. D

This reaction is a decomposition and they are easily recognizable as they have only one reactant compound which forms two or more products.

33. A

A balanced chemical reaction is necessary to reflect the Law of Conservation of Mass. Given that the reactants are rearranging to form the products there must be the same number of atoms of each element on the reactant side of the reaction as there is on the product side of the reaction. This of course is the reason that the mass of the products will equal the mass of the reactants as per the Law of conservation of Mass. We can only use coefficients to change the numbers of atoms of the elements as changing any of the subscripts will change what the actual substance is.

34. A

Firstly this is a synthesis and/or combination reaction. These reaction types can be recognized by there being two lone elements as the reactants, H_2 and Cl_2 in this reaction, which combine to form a single compound as the product, HCl. The change in enthalpy for this reaction is -185kJ. It is a negative value because it appears as a product indicating the products have less enthalpy, potential energy, than the reactants. When energy is lost in the system it has been released into the environment and the reaction is called Exothermic.

35. B

We can determine the moles of O_2 required directly as it is possible to do calculations within a balanced chemical reaction as long as you have the unit mole.

$$6.0\,\text{mol NO} \times \frac{5\,\text{mol O}_2}{4\,\text{mol NO}}$$

$$= 7.5\text{ mol O}_2$$

36. C

The limiting reagent is the reactant that is consumed first and thus stops the reaction and limits the amount of the product that can be produced. We can use molar ratio numbers to see how many moles of $AgNO_3$ are required to react with the Cu in the question.

First we need to find moles of each reactant

$$16.0\,\text{g Cu} \times \frac{1\,\text{mol}}{63.5\,\text{g Cu}} = 0.252\text{mol Cu } 42.0\,\text{g}$$

$$AgNO_3 \times \frac{1\,\text{mol}}{169.9\,\text{g AgNO}_3} = 0.247\text{ mol AgNO}_3$$

$$\frac{\text{Cu}}{0.252\,\text{mol}} = \frac{2\,\text{AgNO}_3}{x}$$

x = the moles $AgNO_3$ required to react with 0.252mol Cu

$x = 0.504\text{mol AgNO}_3$ as we only have 0.247 mol $AgNO_3$ it is the limiting reagent.

37. A

The neutrons and the protons make up the mass of the atom and have essentially the same mass equal to one. Therefore the mass of the atom minis the number (thus mass) of the protons will equal the number (and mass) of the neutrons. Recall number of protons is equivalent to the atomic number.

Number neutrons = atomic mass – atomic number

Number neutrons = 144.2 – 60 = 84.2 = 84 neutrons

38. C

The number of protons in the nucleus of any element's atom is equivalent to the element's atomic number. The atomic number of Iridium is 77 therefore it has 77 protons in the nucleus.

39. A

P^{-3} is an ion with a -3 charge. This indicates there are three more electrons present than protons, electrons having a negative charge and protons having a positive charge. The number of protons never changes therefore the atom has gained and electron and the total will be the number of protons (atomic number) plus three. Number electrons = atomic number + 3 = 15 + 3 = 18 electrons

40. D

Ti is a neutral atom therefore the number of electrons is equal to the number of protons. The number of protons is equivalent to the atomic number which is 22. The electrons are placed into the orbits and sub shells filling the lowest energy levels first: Order of electron filling: 1s → 2s → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f … remembering that s holds 2 electrons, p holds 6 electrons, d holds 10 electrons and f holds 14 electrons.

41. A

Mn+4 is an ion therefore the number of electrons is not equal to the number of protons. The number of protons is equivalent to the atomic number which is 25 for Manganese. The +4 charge means there is 4 more protons present than electrons indicating the atom has lost 4 electrons. The number of electrons is then 25 – 4 = 21. The electrons are placed into the orbits and sub shells filling the lowest energy levels first: Order of electron filling: 1s → 2s → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f … remembering that s holds 2 electrons, p holds 6 electrons, d holds 10 electrons and f holds 14 electrons.

42. A

Column 14 elements are in the P – block and thus the final valance electrons are in the p subshell. Column 14 is the second column in the p block and all elements in this column have 2 electrons in the p subshell while in ground state.

43. A

The alkaline earth metals are those elements in column 2 of the periodic table and Sr is in column (group) 2.

44. D

Halogens are non-metals and do not have any metallic properties. The properties in b,c and d are all those of metals while diatomic in the gaseous phase is a specific property of the halogens.

45. B

The elements in column 15 end in s^2p^3 and require 3 more electrons to have the electronic configuration of the noble gases which is full valence - s^2p^6.

46. C

When there is an electronegativity difference between the substances of between 0.2 and 1.7 there will be a partial electron transfer from the substance with the lower electronegativity to the substance with the higher electronegativity. This results with a substance having a partial positive charge and a substance with a partial negative charge. The partially charged atoms are called dipoles.

47. A

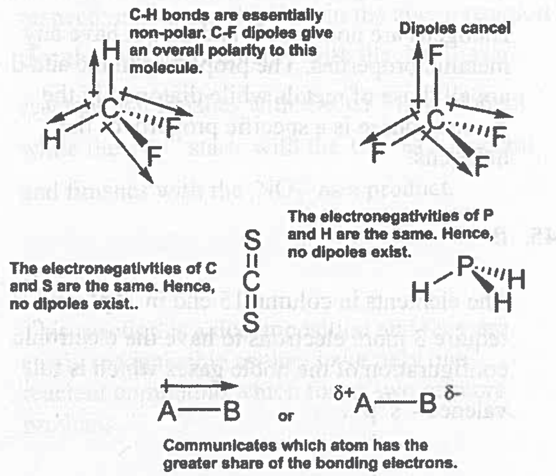
Phosphorus and chlorine are both non-metals so this will not be an ionic bond. We do not have the electronegativities of the elements so it is not possible to determine the exact nature of the bond, polar covalent versus covalent therefore the best answer is a covalent bond.

48. B

Chlorine is a molecular element held together by a single covalent bond. Since each bonded atom has the same electronegativity, the electrons are shared equally.

49. A

Only molecules that contain polar bonds can be polar and then only if the bond dipoles do not cancel as a consequence of molecular geometry. Thus the V.S.E.P.R. shape of a molecule is necessary to predict that molecule's polarity.



50. D

The structural feature of an organic molecule that allows us to classify that molecule by its reactivity is called a functional group.

Some important functional groups are shown in the adjacent column.

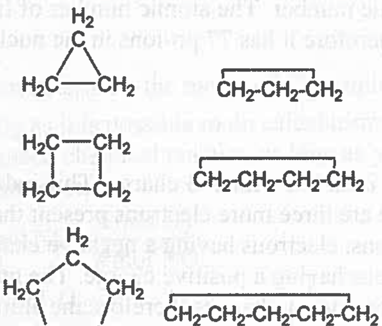
The important functional groups in citric acid are those of alcohols and carboxylic acids.

51. D

Isomers are compounds that have the same chemical formula but different chemical structures. More often than not isomers have different chemical and/or physical properties. Without knowing the precise structures of carvone (other than that it is an organic compound), its two forms meet the qualification for isomers.

52. C

The first cycloalkene is cyclopropane, C_3H_6, the second is cyclobutane, C4H8, and the third member is cyclopentane, C_5H_{10}. These are shown below.



Written Response

1. Define **a)** atom – = the smallest possible unit of an element possible which will still retain the properties of the element.

   b) solution – = a homogenous mixture containing 2 or more substances.

   c) mole – = the mole is the number of particles of Carbon in 12 grams of C-12. The atomic masses of all other elements are relative to that of C-12 so one mole of any element contains the same numbers of particle as one mole of C-12.

2. State Avogadro's hypothesis – = Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles.

3. What is the percent composition by mass of H_2CO_3 ?

   To determine the percent composition by mass of all the elements in the compound, H_2CO_3, we must divide the sum of the masses for each element by the molar mass of the compound and multiply by 100.

   Mass H = 2 × 1.0 = 2.0 g
   Mass C = 1 × 12.0 = 12.0 g
   Mass O = 3 × 16.0 = 48.0 g
   Molar mass H_2CO_3 = 2.0 g
   H + 12.0 g
   C + 48.0 g
   O = 62.0 g H_2CO_3

% by mass H = (2.0g/62.0g) × 100 = 3.2% H
% by mass Mn = (12.0g/62.0g) × 100 = 19.4% C
% by mass O = (48.0g/62.0g) × 100 = 77.4% O

4. Determine the molecular formula of a compound given its empirical formula is C_2H_6O and its molar mass is 138.0 g

$(C_2H_6O)n$ = 138.0 g therefore

$(46.0\,g)_n = 138.0$ g and $n = 138.0/46.0 = 3$

So $(C_2H_6O)_3 = C_6H_{18}O_3$

5. Consider the following reaction:

$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$

What mass of Carbon is required to react completely with 100.0g of the Iron III Oxide?

Plan: mass

$Fe_2O_3 \rightarrow$ moles $Fe_2O_3 \rightarrow$ moles $C \rightarrow$ mass C

$$100.0\,g\,Fe_2O_3 \times \frac{1\,mol}{159.6\,g\,Fe_2O_3} \times \frac{3\,mol\,C}{1\,mol\,Fe_2O_3} \times \frac{12.0\,g\,C}{1\,mol}$$

= 22.56 g C

6. Magnesium has three naturally occurring isotopes: 23-Mg = 23.804amu and has an abundance of 11.01%, 24-Mg = 23.985amu and has an abundance of 79.99% and 25-Mg = 24.986amu and has an abundance of 10.00%. Determine the atomic mass of Magnesium as it appears on the periodic table.

The atomic mass as it appears on the periodic table is a relative average of the isotopes of the element.

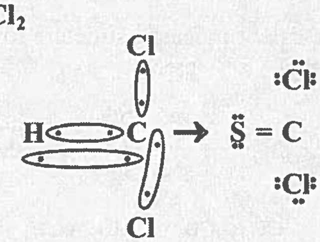
Atomic mass = (fraction of isotope) (mass of isotope) + (fraction of isotope)(mass of isotope) + …

Atomic mass = (0.1101 × 23.804) + (0.7999 × 23.985) + (0.1000 × 24.986)
= 24.305 AMU

7. Showing all work draw the electron (Lewis) dot diagram for

a) $CsCl_2$

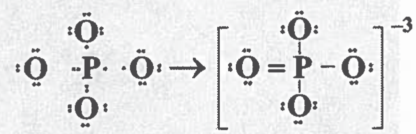
CSCl₂



b) PO_4^{-3}

$\mathbf{PO_4^{-3}}$

5 + 4 × 6 + 3 = 32
P O



8. Jacqueline dilutes the concentration of Chloride ions present in 125.0mL of a 0.150M NaCl solution by adding 75.0mL of 0.195M NaI. She has diluted the concentration of chloride ions but has she diluted the concentration of sodium ion? Show all work.

First calculate the molarities of the compounds after the dilution using the dilution formula

$M_iV_i = M_fV_f$

$$\text{Diluted [NaCl]} = \frac{(0.150\,M)(0.1250\,L)}{(0.2000\,L)}$$
= 0.0938 M NaCl

$$\text{Diluted [NaI]} = \frac{(0.195\,M)(0.0750\,L)}{(0.2000\,L)}$$
= 0.0731 M NaI

Secondly write the dissociation equations for the compounds and from them determine the concentration of the ions.

$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

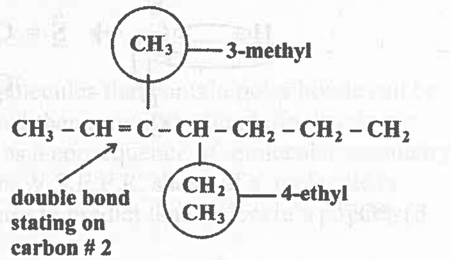
0.0938M so 0.938M 0.0938M

$NaI_{(s)} \rightarrow Na^+_{(aq)} + I^-_{(aq)}$

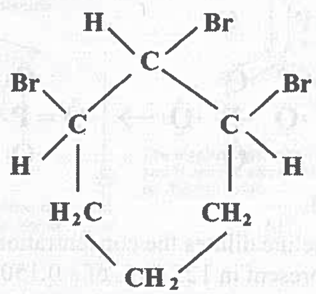
0.0731M so 0.0731M 0.0731M

$[Na^+] = 0.0938 + 0.0731 = 0.167M$
Concentration of Na+ went up to 0.167M from 0.150M in the original solution.

9. Draw the condensed structure for 4- ethyl – 3 – methyl – 2- heptene.



10. Name the compound



= 1,3 – dibromo – 2 – methyl - cyclohexane