Name

Regents Chemistry:

Workbook 2

Chemistry 200

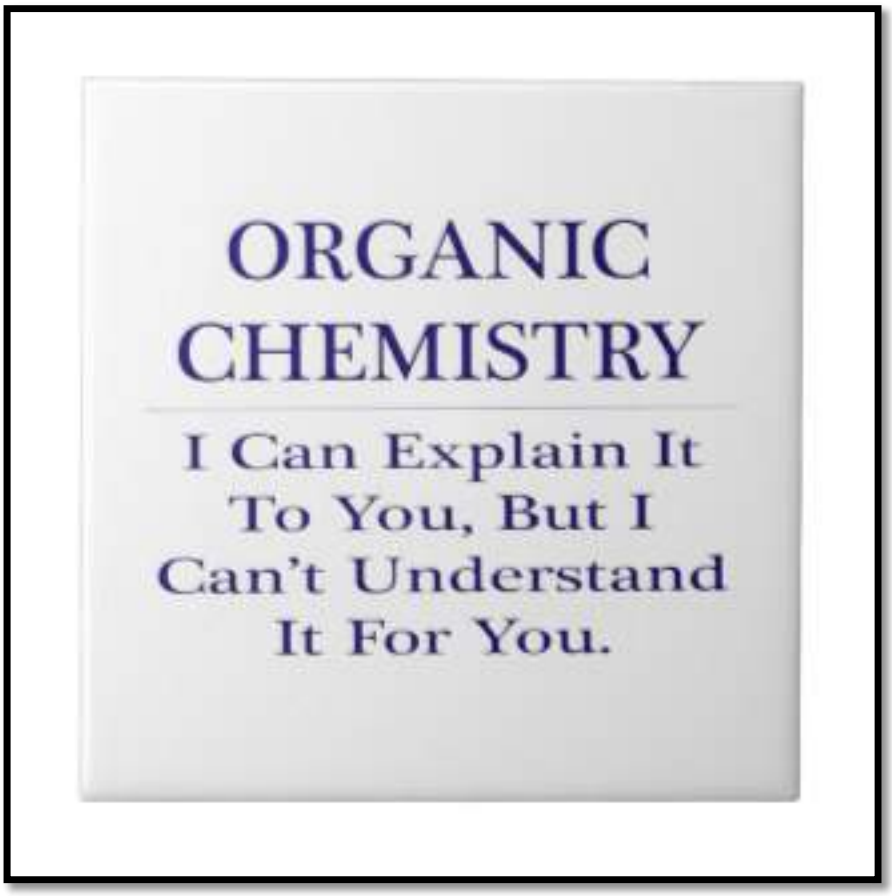
KEEP CALM AND STUDY CHEMISTRY

Name:_____

Regents Chemistry:

Practice Packet

Chapter 8:OrganicChemistry



Chapter 8: Organic Chemistry

Alkane – a hydrocarbon containing only single covalent bonds – saturated hydrocarbon

Alkyl group – a hydrocarbon substituent, the methyl group ($-CH_3$) is an alkyl group

Alkenes – a hydrocarbon containing one or more carbon- carbon double bonds

Alkynes - a hydrocarbon containing a carbon-carbon triple bond

Alkyl halides – a halocarbon in which one ore more halogen atoms are attached to the carbon atoms

Alcohol – an organic compound having an –OH (hydroxyl) group

Aldhyde – an organic compound in which the carbon of the carbonyl group is joined to at least one hydrogen

Addition reaction – a reaction in which a substance is added at the double bond of an alkene or at the triple bond of an alkyne.

Branched-chain alkane – an alkane with one ore more alkyl groups attached to the parent structure

Carbonyl group – a functional group having a carbon atom and an oxygen atom joined by a double bond

Carboxylic acid – an organic acid containing a carboxyl group

Carboxyl group - a functional group consisting of a carbonyl group attached to a hydroxyl group

Condensed structural formula – a structural formal that leaves out some bonds and/or atom; the presence of these atoms or bonds is understood

Ether - an organic compound in which oxygen is bonded to two carbon groups

Esters – a derivative of a carboxylic acid in which the - OH of the carboxyl group has been replaced by the –OR from an alcohol.

Esterfication - An ester is an organic compound where the hydrogen in the compound's carboxyl group is replaced with a hydrocarbon group.

Fatty acids - the name given to continuous-chain carboxylic acids that were first isolated from fats.

Functional group – a specific arrangement of atoms in an organic compound that is capable of characteristics chemical reactions

Fermentation – the production of ethanol from sugars by the action of yeast or bacteria

Hydrocarbon – contain only hydrogen and carbon

Halocarbons - any member of a class of organic compounds containing covalently bonded fluorine, chlorine, bromine or iodine

Hydroxyl group – the –OH functional groups present in alcohols.

Homologous series - a group of compounds in which there is a constant increment of change in molecular structure form one compound in the series to the next

Isomers – compounds that have the same molecular formula buy different molecular structures

Ketone - an organic compound in which the carbon of the carbonyl groups is joined to two other carbons

Monomer - a simple molecule that repeatedly combines to form a polymer

Polymer – a very large molecule formed by the covalent bonding of repeating small molecules, known as monomers

Substituent - an atom or group of atoms that can take the place of a hydrogen atom on a parent hydrocarbon

Saturated compounds – an organic compound in which all carbon atoms are joined by single covalent bonds

Straight-chain alkanes – a saturated hydrocarbon that contains any number of carbons atoms arranged one after the other

Substitution reaction – a common type of organic reaction, which involves the replacement of an atom or group of atoms by another atom or group of atoms

Soaoponification – the hydrolysis of fats or oils by a hot aqueous alkali-metal hydroxide, the making of soaps

Unsaturated compounds – an organic compound with one or more double or triple carbon-carbon bonds

Organic Chemistry Introduction

8.1

Objective:

What is a hydrocarbon and the properties of organic molecules?

How do we use table P and Q to write structural and molecular formulas for hydrocarbons?

Organic versus Hydrocarbon

- Organic molecules must have the element C.
- Hydrocarbons can only have the elements H and C.
- Therefore hydrocarbons are organic but not all organic compounds are hydrocarbons:
  - CH_4 is a hydrocarbon and is organic
  - CCl_4 is organic but not a hydrocarbon
  - O_2 is neither organic nor a hydrocarbon

Organic molecules

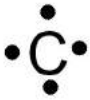
- Found in fossil fuels, plants and animals.
- Examples include gasoline, oil, kerosene, butane, propane…

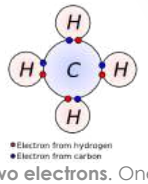
Hydrocarbon Properties

- Mostly insoluble
- Non-electrolytes (do not conduct electricity)
- React very slowly
- As size increases, the melting point and boiling point of the hydrocarbons increase.
- Small hydrocarbons may be gases and large hydrocarbons may be solids at room temperature.

Carbon

- Remember carbon has four valence electrons. Therefore carbon will bond four times to achieve an octet.

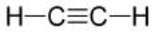




- **Each bond shown shares two electrons.** One from C and one from H.

Organic molecules

- Sometimes double and triple bond will be needed to fulfill all octets.

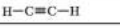


- A double bond shares four electrons.
- A triple bond shares six electrons.

Organic Molecules

- All single bonded hydrocarbons are in the same family known as alkanes.
- All double bonded hydrocarbons are in the same family known as alkenes.
- All triple bonded hydrocarbons are in the same family known as alkynes.
- Refer to table Q

Table Q
Homologous Series of Hydrocarbons

Name	General Formula	Examples	
		Name	Structural Formula
alkanes	C_nH_{2n+2}	ethane	
alkenes	C_nH_{2n}	ethene	
alkynes	C_nH_{2n-2}	ethyne	

Note: n = number of carbon atoms

Organic Molecules

- Notice the picture shows you the number of bonds.
- Table Q also shows that if you know how many carbon atoms are present in a molecule and what type of bonds it has, you can CALCULATE the number of hydrogen atoms using the general formula. But drawing it out might be easier.

Organic Molecules

Table P
Organic Prefixes

Prefix	Number of Carbon Atoms
meth-	1
eth-	2
prop-	3
but-	4
pent-	5
hex-	6
hept-	7
oct-	8
non-	9
dec-	10

Table P shows prefixes to determine how many Carbon atoms a compound has.

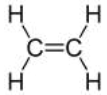
Give the prefix for the following:

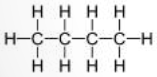
C_2H_6	Eth
C_3H_6	Prop
C_4H_6	But
C_5H_{12}	Pent
C_6H_{12}	Hex
C_7H_{14}	Hept
C_8H_{18}	Oct
C_9H_{16}	Non
$C_{10}H_{20}$	Dec

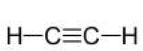
Organic Compounds

- Putting P and Q together we can name simple hydrocarbons:


Methane


Ethene


Butane


Ethyne

Saturation

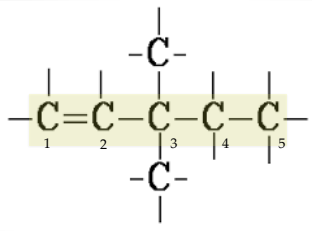
- Compounds like fats that are saturated have many hydrogen atoms. This requires single bonds. Alkanes are saturated with single bonds.
- Compounds that are unsaturated have double and triple bonds, therefore, they have less hydrogen atoms. Alkenes and alkynes are unsaturated.

Branched Hydrocarbons

Video 8.2

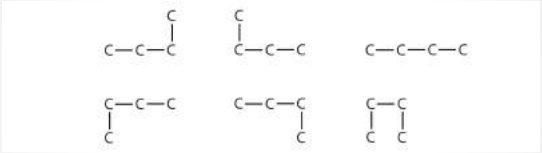
Branched hydrocarbons

- When naming branched hydrocarbons, name the longest continuous chain and use that as the 'last name.' Making sure the multiple bond is part of that chain. Then name the shorter chains, specifying the position of each branch. Also make sure that your branches are numbered as low as possible.



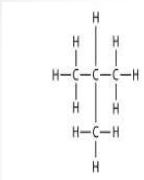
3,3 dimethyl 1 pentene

Branched hydrocarbons

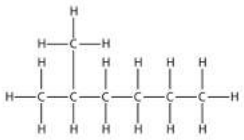


These are all just butane!!!

Branched hydrocarbons



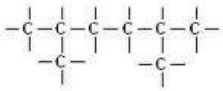
2 methyl propane



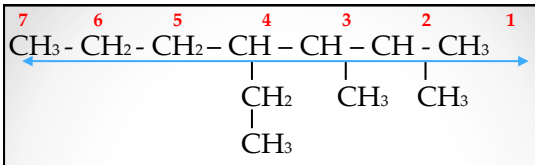
2 methyl hexane

The little branches are known as alkyl groups which is why they have a "yl" ending.

Branched hydrocarbons



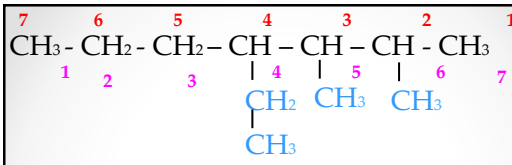
2,5 dimethyl hexane



Find the longest continuous chain of carbons. This is the parent chain. Look at all bonds between carbons to determine type of hydrocarbon . Count from the side with the alkyl groups

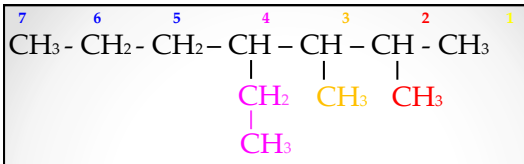
All single bonds so ending is… ane.

There are 7 continuous carbons, so the parent chain is heptane



Number the carbons in the main sequence starting with the end that will give the alkyl groups the smallest #.

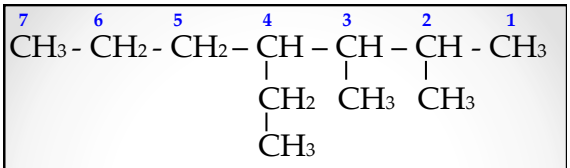
The chain is numbered from right to left because it gives the attached groups the lowest possible number



Add numbers to the names of the groups to identify their positions on the chain. These are prefixes with a "yl" ending.

In this ex. the positions are:

2 - methyl, 3 - methyl, 4 - ethyl



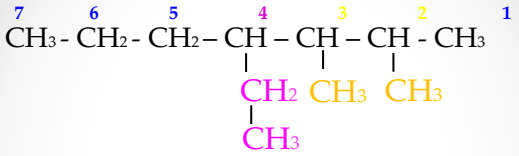
Use prefixes to indicate the appearance of a group more than once in the structure. And list them in alpha order

Di = twice

Tri = three times

Tetra = four times

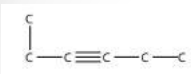
Penta= five times



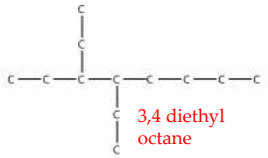
- The name of this compound is:

2,3-dimethyl – 4-ethyl heptane

Branched hydrocarbons

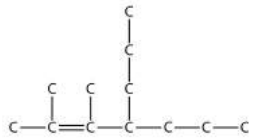


3 hexyne



3,4 diethyl octane

Branched hydrocarbons

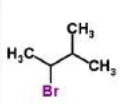


2, 3 dimethyl 4 propyl 2 heptene

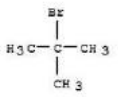
Other groups

- More groups can be on the ring/branch such as:
  - F fluoro
  - Cl chloro
  - Br bromo
  - I iodo

Branched hydrocarbons



2 bromo 3 methyl butane



2 bromo 2 methyl propane

Isomers

Video Lesson 8.3

Objectives

- **Identify structural isomers.**

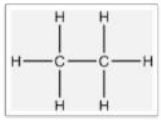
Isomers

- Isomers are compounds that have the same simple molecular formula, but different structures.

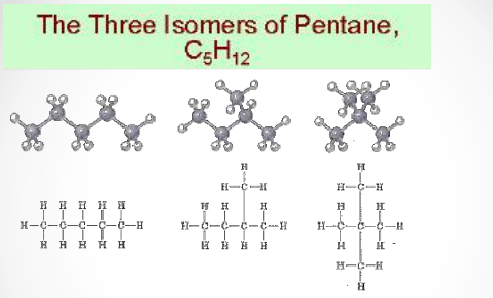
Formulas

1. Molecular Formula: shows the number of atoms of each element in a compound.
2. Structural Formula: diagram of the molecular shape of a compound.
3. Condensed Structural Formula: each carbon is written separately followed by atoms bonded to it.

C_2H_6

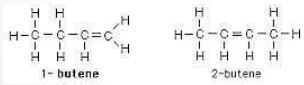


CH_3CH_3



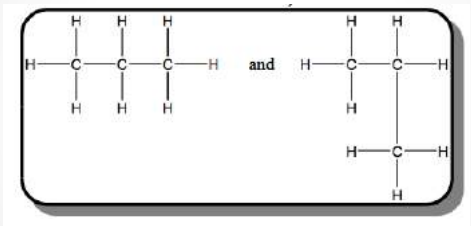
Isomers

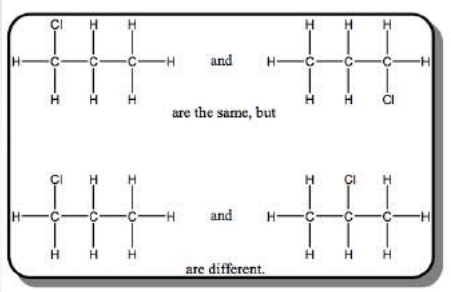
- Isomers have the same molecular formula but rearranged in a different structure with different properties.
- Draw two isomers of butene:



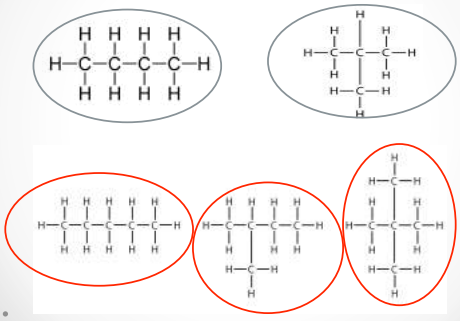
- Why not 3-butene?

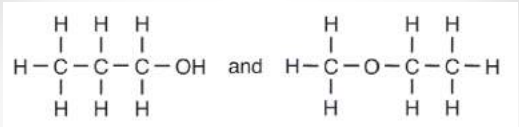
Not Isomers





Circle the isomers:





C_3H_8O C_3H_8O

Functional groups

Video 8.4

Objective:

How do we use Table R to recognize structural and molecular formulas for organic molecules containing functional groups?

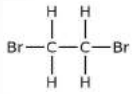
TAKE OUT TABLE R!!!!!

Table R
Organic Functional Groups

Class of Compound	Functional Group	General Formula	Example
halide (halocarbon)	—F (fluoro-) —Cl (chloro-) —Br (bromo-) —I (iodo-)	$R-X$ (X represents any halogen)	$CH_3CHClCH_3$ 2-chloropropane
alcohol	$-OH$	$R-OH$	$CH_3CH_2CH_2OH$ 1-propanol
ether	$-O-$	$R-O-R'$	$CH_3OCH_2CH_3$ methyl ethyl ether
aldehyde	$-C(=O)-H$	$R-C(=O)-H$	$CH_3CH_2C(=O)-H$ propanal
ketone	$-C(=O)-$	$R-C(=O)-R'$	$CH_3C(=O)CH_2CH_2CH_3$ 2-pentanone
organic acid	$-C(=O)-OH$	$R-C(=O)-OH$	$CH_3CH_2C(=O)-OH$ propanoic acid
ester	$-C(=O)-O-$	$R-C(=O)-O-R'$	$CH_3CH_2C(=O)OCH_3$ methyl propanoate
amine	$-N-$	$R-N-R'$	$CH_3CH_2CH_2NH_2$ 1-propanamine
amide	$-C(=O)-NH-$	$R-C(=O)-NH-$	$CH_3CH_2C(=O)-NH_2$ propanamide

Note: R represents a bonded atom or group of atoms.

Halides

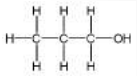


1, 2 dibromo ethane

- Have one of the halogens as a branched group. Names as we did in the previous lessons.

halide (halocarbon)	—F (fluoro-) —Cl (chloro-) —Br (bromo-) —I (iodo-)	$R-X$ (X represents any halogen)	$CH_3CHClCH_3$ 2-chloropropane

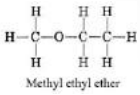
Alcohol



- Suffix: -ol
- Flammable, soluble

alcohol	$-OH$	$R-OH$	$CH_3CH_2CH_2OH$ 1-propanol

Ether

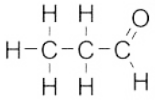


Methyl ethyl ether

- Name small chain, large chain, suffix: -ether
- Anesthetic, soluble

ether	$-O-$	$R-O-R'$	$CH_3OCH_2CH_3$ methyl ethyl ether

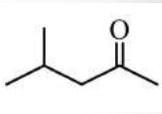
Aldehyde



- Suffix: -al
- Soluble, reactive

aldehyde	$-C(=O)-H$	$R-C(=O)-H$	$CH_3CH_2C(=O)-H$ propanal

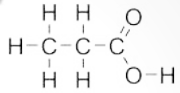
Keytone



- Suffix: -one
- Somewhat soluble, needs at least 3 carbons

ketone	$-C(=O)-$	$R-C(=O)-R'$	$CH_3CCH_2CH_2CH_3$ 2-pentanone

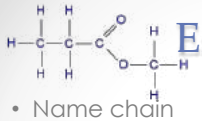
acids



- Suffix: -anoic acid
- Also known as carboxylic acids: weak acids/ **weak electrolytes**

organic acid	$-\overset{O}{\overset{\|\|}{C}}-OH$	$R-\overset{O}{\overset{\|\|}{C}}-OH$	$CH_3CH_2\overset{O}{\overset{\|\|}{C}}-OH$ propanoic acid

Esters



- Name chain adjacent to double bonded O last, Suffix: -anoate
- Smells great (perfumes, foods)

ester	$-\overset{O}{\overset{\|\|}{C}}-O-$	$R-\overset{O}{\overset{\|\|}{C}}-O-R'$	$CH_3CH_2\overset{O}{\overset{\|\|}{C}}OCH_3$ methyl propanoate

Amine

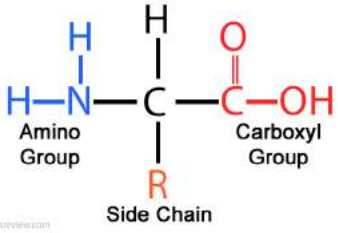
Nitogen is present

- Prefix: amino-
- Basic, used in dyes, found in proteins: DNA

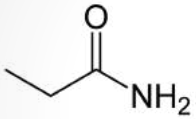
amine	$-\overset{\|}{N}-$	$R-\overset{R'}{\overset{\|}{N}}-R''$	$CH_3CH_2CH_2NH_2$ 1-propanamine

Amine

Amino Acid Structure



Amide



- **Suffix: -amide**
- **Used in dyes**

amide	$-\overset{O}{\overset{\|\|}{C}}-\overset{\|}{N}H$	$R-\overset{O}{\overset{\|\|}{C}}-\overset{R'}{\overset{\|}{N}}H$	$CH_3CH_2\overset{O}{\overset{\|\|}{C}}-NH_2$ propanamide

Polymers

- Long Carbon chains
- $(C_2H_4)_n$
- Used in all plastics, runner, nylons…

Organic Reactions

Video Lesson 8.5

Objectives

- Describe and classify different types of organic reactions.

7 Types of Organic Reactions

1. Combustion
2. Substitution
3. Addition
4. Esterification
5. Saponification
6. Fermentation
7. Polymerization

Combustion

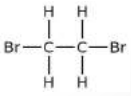
- An alkane is burned in the presence of oxygen to produce water and carbon dioxide (O_2 is always a reactant!)

Addition

- Similar to synthesis reactions: one product forms
- $C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$
- Notice the first compound is unsaturated. The bond breaks to allow new Bromine atoms into the chemical.

Before:

After:



Substitution

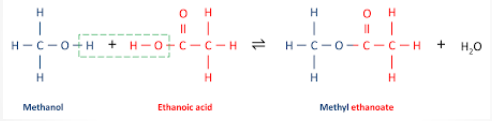
- 1 or more hydrogen atom in a SATURATED ALKANE is replaced by another atom/group

R-H + X_2 → R-X + HX

Alkane Halogen Halocarbon Hydrogen halide

Esterification

organic acid + alcohol → ester + water



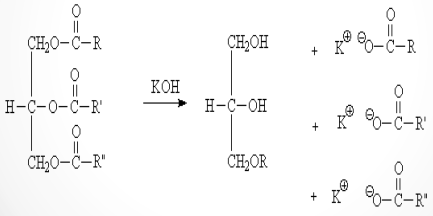
Saponification

- Making soap

Ester + base → alcohol + soap

Saponification

Saponification: Bases added to fats yield glycerol and soap. Look for large molecules and bases!



Fermentation

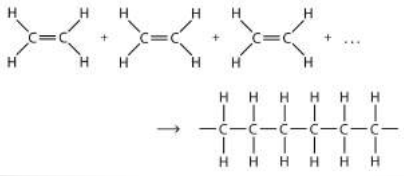
- The production of alcohol (ethanol)

Sugar $\xrightarrow{enzyme}$ alcohol + CO_2

Example: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

Polymerization

- Formation of large molecules called polymers
- $C_2H_4 + C_2H_4 \rightarrow (C_2H_4)_2$



Sketch Notes

Sketch Notes

Chapter 8: Organic Chemistry

1	Methane CH_4		
2	Ethane C_2H_6	Ethene C_2H_4	Ethyne C_2H_2
3			
4			
5			
6			
7			
8			
9			
10			

Video 8.1 Hydrocarbons

1. How many carbon atoms are in each compound?

a. Methane ____	f. Hexane ____	k. Decane ____
b. Ethane ____	g. Ethyne ____	l. Butyne ____
c. Ethene ____	h. Propane ____	m. Butane ____
d. Pentane ____	i. Heptane ____	n. Propyne ____
e. Propene ____	j. Octane ____	o. Butene ____

2. For each compound fill in each blank:

	Number of Carbon atoms	Series	Formula
a. Methane	_____	_____	_____
b. Butane	_____	_____	_____
c. Propyne	_____	_____	_____
d. Pentane	_____	_____	_____
e. Octane	_____	_____	_____
f. Heptene	_____	_____	_____
g. Propene	_____	_____	_____
h. Butyne	_____	_____	_____
i. Decane	_____	_____	_____
j. Nonane	_____	_____	_____
k. Heptane	_____	_____	_____
l. Ethyne	_____	_____	_____
m. Hexyne	_____	_____	_____
n. Ethane	_____	_____	_____
o. Propane	_____	_____	_____
p. Decene	_____	_____	_____
q. Octyne	_____	_____	_____

3. How many times does carbon bond and why?

Answer the following questions.

1. ___All organic compounds must contain:
    1. hydrogen
    2. nitrogen
    3. carbon
    4. oxygen

2. ___ Which element is composed of atoms that can form more than one covalent bond with one another?
    1. hydrogen
    2. helium
    3. carbon
    4. calcium

3. ___ What is the total number of valence electrons in a carbon atom in the ground state
    1. 12
    2. 2
    3. 6
    4. 4

4. ___ Which property is generally characteristic of an organic compound?
    1. low melting point
    2. high melting point
    3. mainly polar
    4. mainly nonpolar

5. ___ In general, which property do organic compounds share?
    1. high melting points
    2. high electrical conductivity
    3. readily soluble in water
    4. slow reaction rate

6. ___ A hydrocarbon molecule containing one triple bond is classified as an:
    1. alkene
    2. alkane
    3. alkyne
    4. alkadience

7. ___ What is the total number of hydrogen atoms in a molecule of butane?
    1. 10
    2. 6
    3. 8
    4. 4

8. ___ By how many carbon atoms does each member of a homologous series differ from the previous member?
    1. 1
    2. 2
    3. 3
    4. 4

9. ___ Which of the following is a saturated hydrocarbon?
    1. ethene
    2. ethyne
    3. propene
    4. propane

10. ___Which compound is a member of the same homologous series as C_3H_6?
    1. C_2H_4
    2. C_2H_6
    3. C_3H_4
    4. C_3H_8

11. ___ Which hydrocarbon is a member of the series with the general formul C_nH_{2n-2}?
    1. ethyne
    2. ethane
    3. butane
    4. benzene

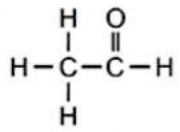
12. ___ Which compound belongs to the alkene series?
    1. C_2H_2
    2. C_2H_4
    3. C_6H_6
    4. C_6H_{14}

13. ___ Which type of bond occurs in a saturated hydrocarbon molecule?
    1. single covalent
    2. double covalent
    3. triple covalent
    4. ionic

14. ___In which group could the hydrocarbons all belong to the same homologous series?
    1. C_2H_2, C_2H_4, C_2H_6
    2. C_2H_4, C_3H_4, C_4H_8
    3. C_2H_4, C_2H_6, C_3H_6
    4. C_2H_4, C_3H_6, C_4H_8

15. ___ Which formula represents butane?
    1. CH_3CH_3
    2. $CH_3CH_2CH_3$
    3. $CH_3CH_2CH_2CH_3$
    4. $CH_3CH_2CH_2CH_2CH_3$

16. ___ Which formula represents an unsaturated hydrocarbon?

A)


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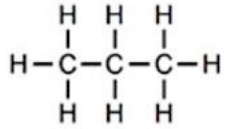

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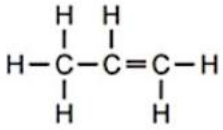

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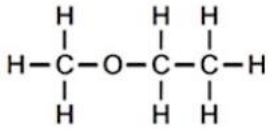
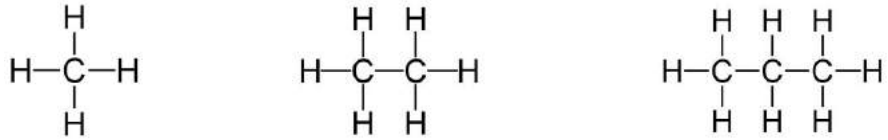


Video Lesson 8.1

Background: Structural formulas show the arrangement of the atoms within the molecules as far as which atoms are bonded to which and whether single, double or triple bonds are used.

Figure 1:

Structural formulas for alkanes:



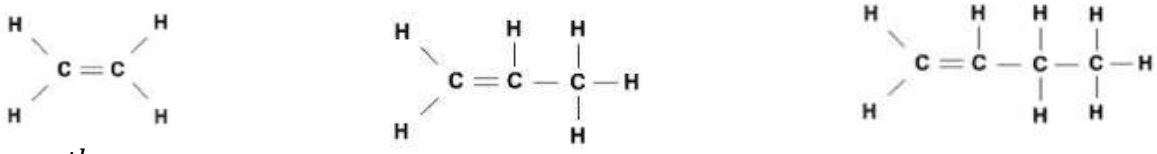
methane *ethane* *propane*

1. Using Tables P and Q in your reference table, draw the structural formula for the following alkanes. **Name each compound.**

a. C_4H_{10} b. C_5H_{12} c. C_6H_{14}

Figure 2:

Structural Formulas for Alkenes



ethene *propene* *butene*

1. Based upon Figure 2 and your knowledge of alkenes, why does the compound methene not exist?

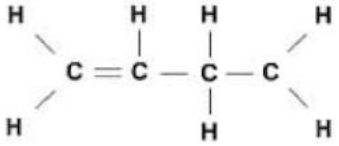
2. Why do the carbon atoms with the double bond contain 1 less Hydrogen atoms then carbon atoms that contains a single bond?

3. Using Tables P and Q, draw the structural formula for the following alkenes. Name each compound.

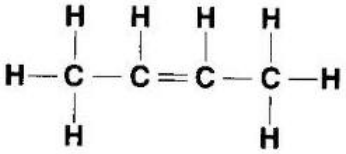
   a. C_5H_{10} b. C_6H_{12} c. C_7H_{14}

When naming alkenes you must give the location of the double bond in the name when there are more than 3 carbon atoms in the compound. You do this by numbering the carbon atoms and stating which number carbon the double bond is on. You can number the carbon atoms ***from left to right*** or ***right to left*** which ever gives the double bond the lowest possible numbered location. This is because compounds are not stationary in the "real world" and are therefore constantly moving. See Figure 3 below.

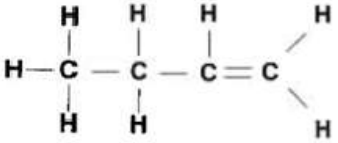
Figure 3:



1-butene

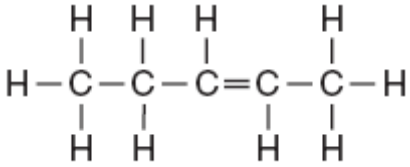


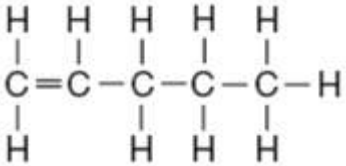
2-butene

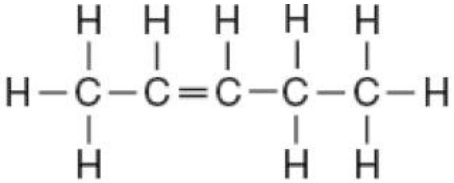


1-butene

1. Using Figure 3 and reference tables P & Q name the following compounds:



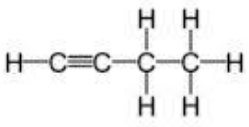




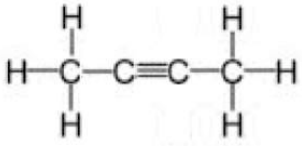
Drawing structural formulas for alkynes is exactly the same as alkenes except they contain a triple bond instead of a double bond.

Figure 4:

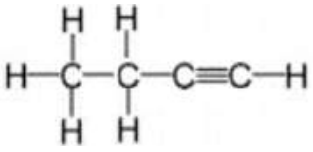
Structural Formulas for alkynes



1- butyne



2-butyne

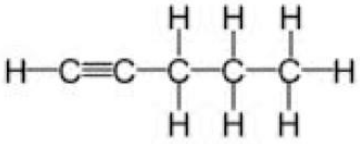


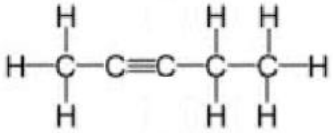
1-butyne

1. Why do the carbons with the triple bond contain no bonded hydrogen atoms?

2. Using Reference Tables P and Q, draw the structural formula for the following alkynes. Name each compound.
   a. C_5H_8 b. C_6H_{10} c. C_7H_{12}

3. Name the following compounds:





Practice: ***Draw the structural formula for the following compounds:***

a. C_8H_{16} b. C_4H_6 c. 2-hexene

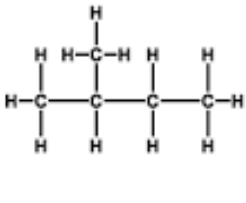
d. 2-heptyne e. 3-hexene f. 1-heptyne

Video Lesson 8.2

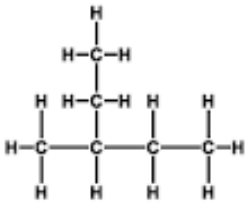
Chemists use a system developed by the IUPAC (International Union of Pure and App lied Chemistry) system for naming isomers.

1. Identify the longest *continuous* carbon chain. This chain is called the parent chain and forms the basis for the name of the hydrocarbon.
2. Identify all of the substituents (groups branching from the parent chain). The substituents are named using the proper prefix (meth-, eth-, etc) and a –yl ending.
3. Number the carbons of the parent chain from the end that gives the substituents the lowest numbers.
4. If the same substituent occurs more than once, the location of each point on which the substituent occurs is given. In addition, the number of times the substituent group occurs is indicated by a prefix (di, tri, tetra, etc.).
5. If there are two or more different substituents they are listed in alphabetical order using the base name (ignore the prefixes).

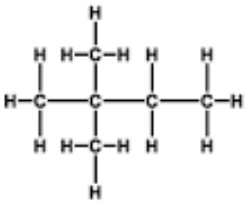
The following examples will illustrate this:



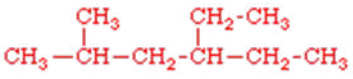
2-methylbutane



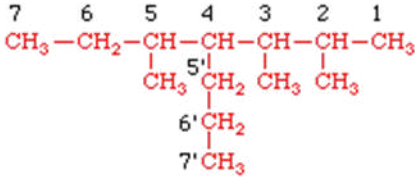
3-methylpentane
(not 2-ethylbutane)



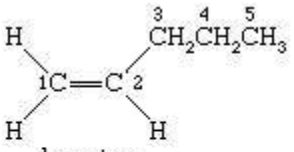
2,2-dimethylbutane



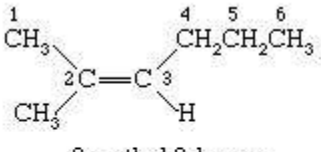
4-ethyl-2-methylhexane



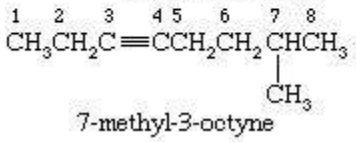
2,3,5-trimethyl-4-propylheptane



1-pentene



2-methyl-2-hexene



7-methyl-3-octyne

When numbering, the parent chain determines the lowest number.

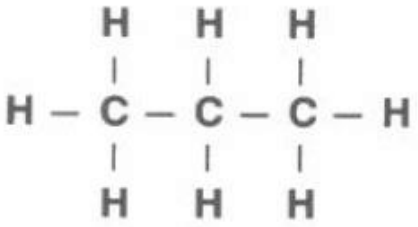
Draw the structural formula for 3-ethyl-5-methyl-3-heptene.

Structure of Hydrocarbons

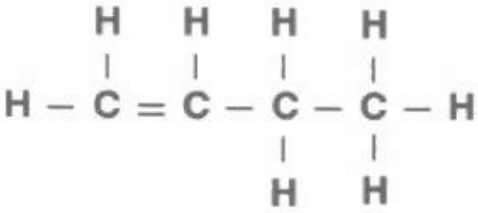
1. ethane	5. ethyne
2. propene	6. 3,3-dimethyl pentane
3. 2-butene	7. 2,3 –dimethyl pentane
4. methane	8. 2-butyne

Naming Hydrocarbons

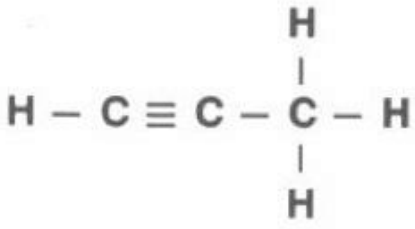
1.



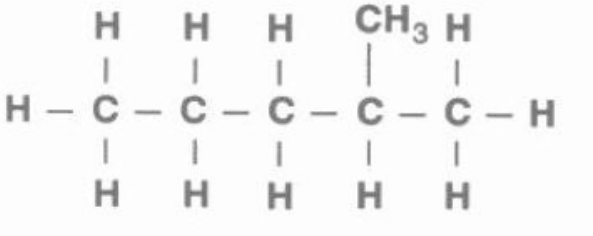
2.



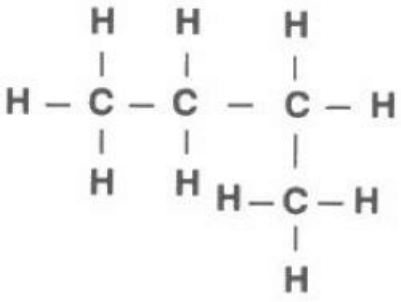
3.



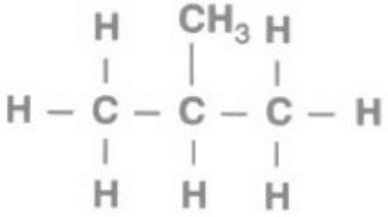
4.



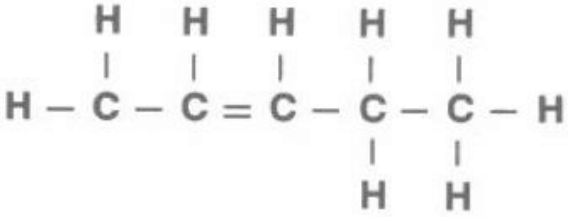
5.



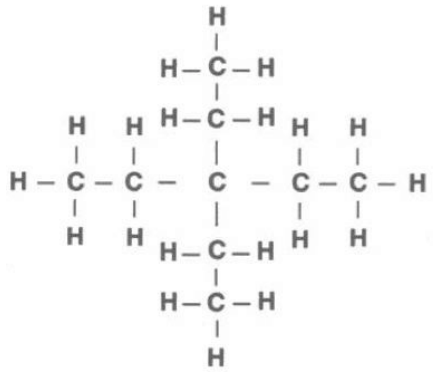
6.



7.



8.

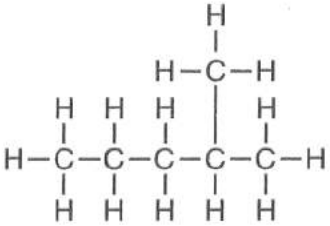


Organic Structural Formulas

1. Which element is present in all organic compounds?

   1) carbon 2) hydrogen
   3) nitrogen 4) oxygen

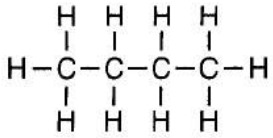
2. What is the IUPAC name of the organic compound that has the formula shown below?



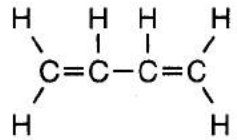
   1) 1,1-dimethylbutane
   2) 2-methylpentane
   3) hexane
   4) 4-methylpentane

3. Which formula represents 2-butene?

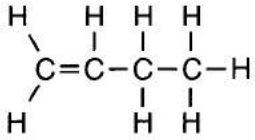
   1)



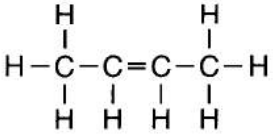
   2)



   3)



   4)

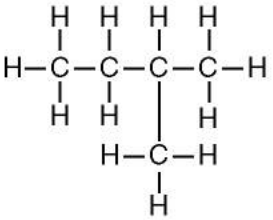


4. Which formula represents propyne?

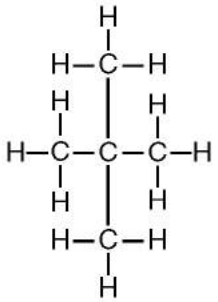
   1) C_3H_4 2) C_3H_6
   3) C_5H_8 4) C_5H_{10}

5. Which structural formula represents 2-pentyne?

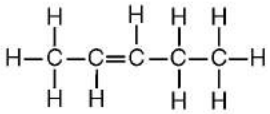
   1)



   2)



   3)



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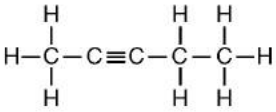
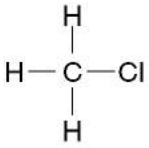
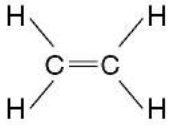


6. Which structural formula is *incorrect*?

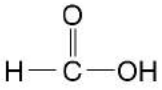
   1)



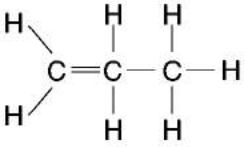
   2)



   3)

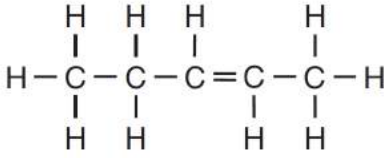


   4)



7. In the space below, draw a structural formula for a molecule of 2,2,4-trimethylpentane.

8. Given the formula representing a compound:



   What is a chemical name of this compound?

   1) 2-pentene 2) 2-pentyne
   3) 3-pentene 4) 3-pentyne

9. Which condensed structural formula represents an unsaturated compound?

1) $CH_3CHCHCH_3$
2) $CH_3CH_2CH_3$
3) CH_3CH_3
4) CH_4

10. Given the structural formula for ethyne:

What is the total number of electrons shared between the carbon atoms?

1) 6 2) 2 3) 3 4) 4

Video Lesson 8.3: Isomers

1. Record the Structural formula, molecular formula, and condensed formula for the following:

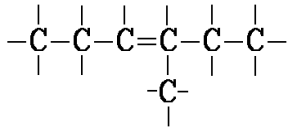
Name	Structural	Molecular	Condensed
2, 3-dimethyl butane			
2, 2-dimethyl butane			
2-heptyne			
3-hexene			
2-methyl 1-pentene			

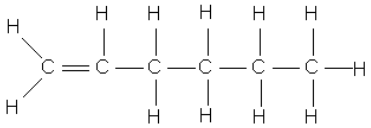
2. Where any of the above isomers? Explain your answer. _____

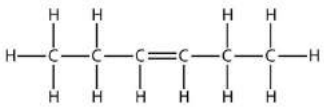
_____

3. Draw an isomer of 2-heptyne below. Give the name of your isomer: _____

4. Name the following and identify the isomers.





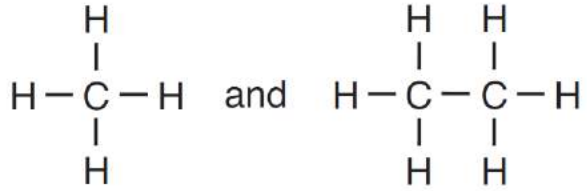


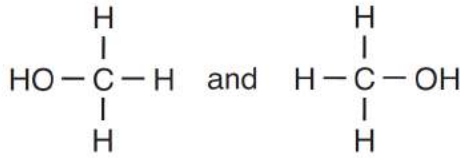
_____ _____ _____

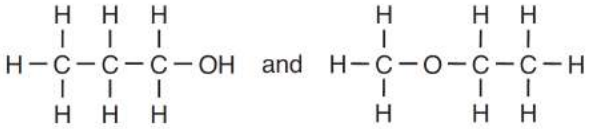
5. Which of the hydrocarbons in the table above were saturated?

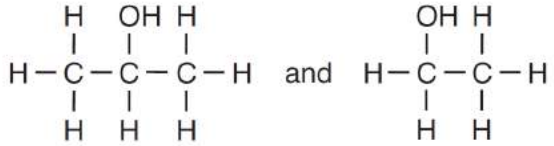
Isomers 8.3

____ 1. Which formulas represent compounds that are isomers of each other?

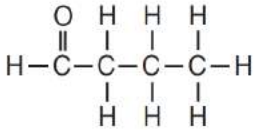
1) 

2) 

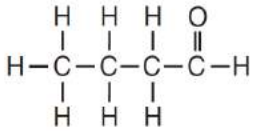
3) 

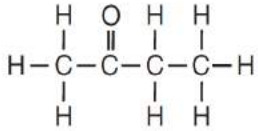
4) 

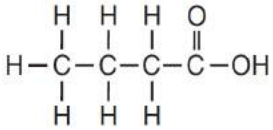
____ 2. Given a formula representing a compound:

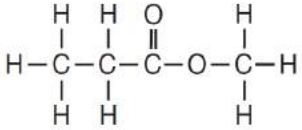


Which formula represents an isomer of this compound?

1) 

2) 

3) 

4) 

____ 3. The two isomers of butane have different

1) formula masses
2) empirical formulas
3) molecular formulas
4) structural formulas

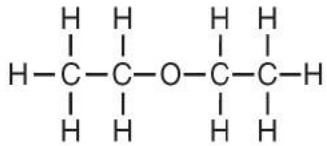
____ 4. The isomers butane and methylpropane differ in their

1) molecular formulas
2) structural formulas
3) total number of atoms per molecule
4) total number of bonds per molecule

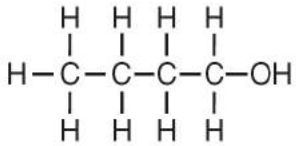
____ 5. Which two compounds are isomers of each other?

1) CH_3CH_2COOH and $CH_3COOCH_2CH_3$
2) CH_3CH_2CHO and CH_3COCH_3
3) $CH_3CHBrCH_3$ and $CH_2BrCHBrCH_3$
4) $CH_3CHOHCH_3$ and $CH_3CHOHCH_2OH$

____ 6. Given the formulas for two compounds:



and



These compounds differ in

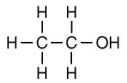
1) gram-formula mass
2) molecular formula
3) percent composition by mass
4) physical properties at STP

____ 7. Two substances have different physical and chemical properties. Both substances have molecules that contain two carbon atoms, one oxygen atom, and six hydrogen atoms. These two substances must be

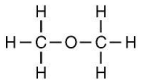
1) isomers of each other
2) isotopes of each other
3) the same compound
4) the same hydrocarbon

8. Given the structural formulas:

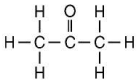
Formula A



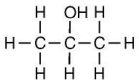
Formula B



Formula C



Formula D



Which two formulas represent compounds that are isomers of each other?

1) *A* and *B*
2) *A* and *C*
3) *B* and *D*
4) *C* and *D*

9. The compounds CH_3OCH_3 and CH_3CH_2OH are isomers of each other. These two compounds must have the same

1) density
2) reactivity
3) melting point
4) molecular formula

10. Which pair of compounds are isomers?

1) NO_2 and N_2O_4
2) P_2O_5 and P_4O_{10}
3) HCOOH and CH_3COOH
4) CH_3OCH_3 and C_2H_5OH

Chapter 8: Organic Chemistry

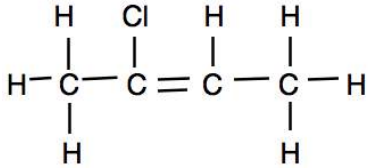
VideoLesson 8.4: Functional Groups

For the following compounds, determine the family and draw the compound:

Name	Family	Structural Formula	Condensed Formula
Butanoic acid			
Methanal			
Butanamide			
3-iodo octane			
Methyl pentanoate			
Ethanol			
2-heptanone			
Diethyl ether			
2-pentanol			
Ethanoic acid			
2-propanamine			
Hexanal			
Ethyl methanoate			

Video&Lesson 8.4:

Classify each of the following structural formulas and write each name



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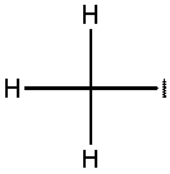

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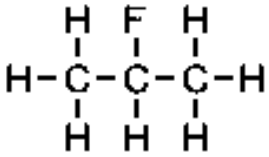

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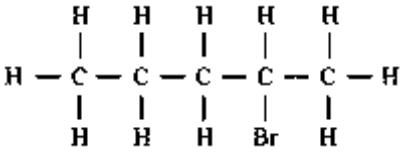

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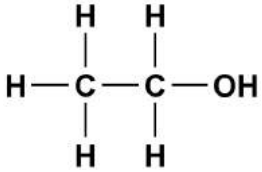

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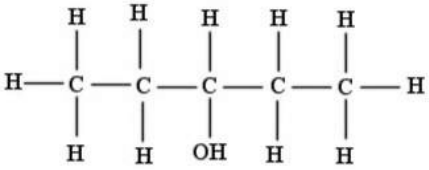

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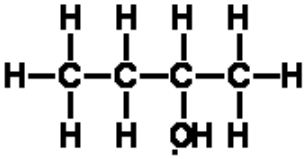

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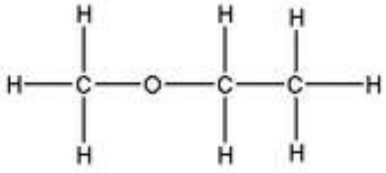

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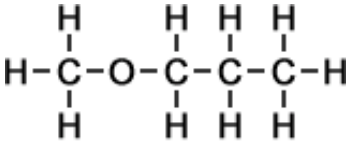

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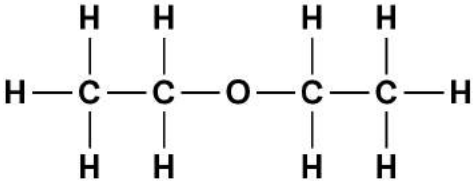

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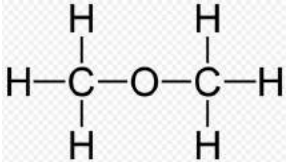

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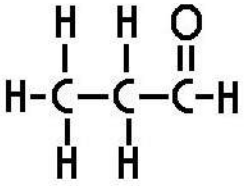

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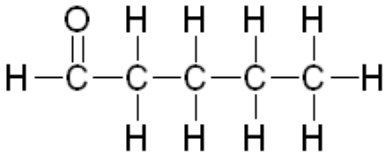

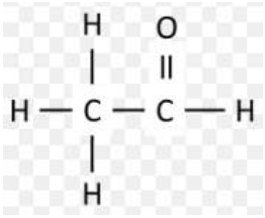

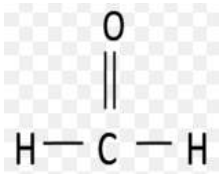

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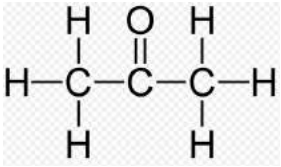
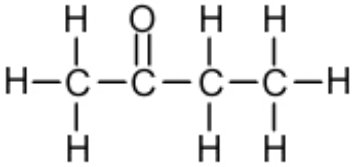
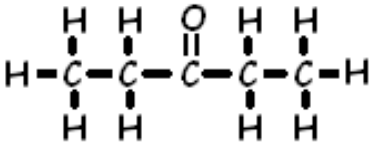






_____ _____

Classify each name and draw the structural formula

2 hexanol

ethyl methyl ether

3 heptanol

2 hexanone

butanal

2 pentanone

Video&Lesson 8.5: Organic Reactions

Combustion:Many organic compounds react with excess oxygen to form carbon dioxide and water. On Table I of your reference, the first 6 reactions are combustion reactions. Write a balance reaction for the combustion of:

1. Ethane:

2. Pentane:

Substitution: Saturated hydrocarbons(ALKANES) may replace a hydrogen atom in the molecule with another element usually a halogen.

Example $C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$

Draw the structural formulas for the above reaction:

Name the product C_2H_5Cl_____
Write a balanced reaction for the substitution of bromine onto propane.

Draw the structure of and name two possible halocarbon isomers formed in the above reaction.

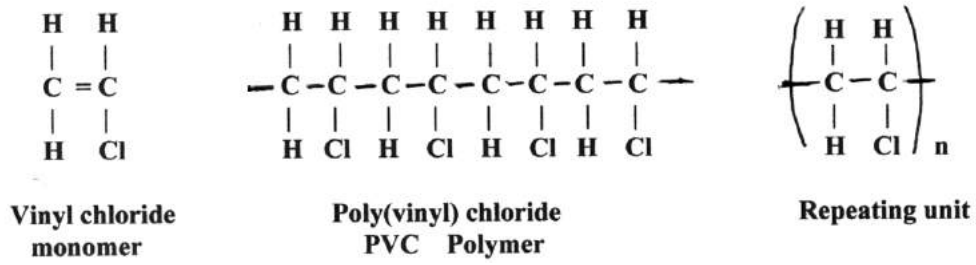
Addition:Unsaturated hydrocarbons (ALKENES or(ALKYNES) can add a atom of hydrogen or of a halogen at the site of a double or triple bond. When hydrogen is added, the process is called HYDROGENATION. When a halogen is added, the process is called HALOGENATION.

$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

Name the product_____

Now write structural formulas for the addition of Cl_2 onto 2 butene. Name the product. Notice that, unlike substitution, only one product is possible!
When hydrogen is added to propene, what is the name of the new hydrocarbon thae forms?Write a balanced equation to illustrate this reaction.

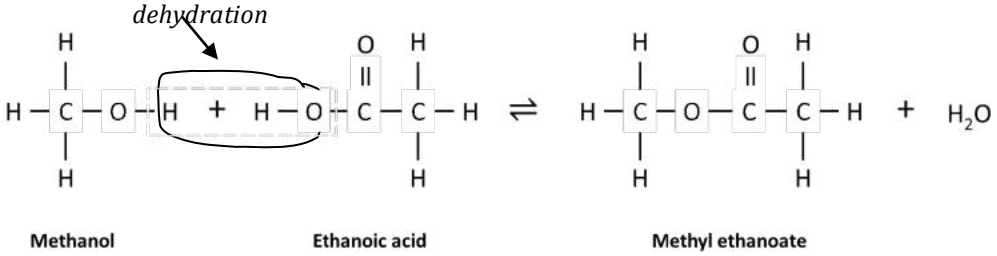
Polymerization: Large molecules can form when individual units of molecules (***monomers***) are chained together to form a ***polymer***. If the individual monomer is an unsaturated hydrocarbon, ***addition polymerization*** my occur as the double (or triple) bond is "broken open" and a chain is formed:(



Vinyl chloride monomer | **Poly(vinyl) chloride PVC Polymer** | **Repeating unit**

Esterification: Esters are compounds which have pleasant odors. They are formed by the reaction between organic acids and alcohols.

Ethanoic acid and methanol will react to form methyl ethanoate. The structural formulas for this reaction are shown below.



Methanol | **Ethanoic acid** | **Methyl ethanoate**

Now draw the structures, determine the products and name eac reactant and organic product in the following esterification reactions:

$C_2H_5COOH + C_2H_5OH \rightarrow$

$HCOOH + C_3H_7OH \rightarrow$

$C_3H_7COOH + CH_3OH \rightarrow$

Fermentation: In the fermentation process, enzymes found in living things, such as yeast, convert carbohydrates usually sugar into carbon dioxide and alcohol.

Glucose($C_6H_{12}O_6$) is fermented in the presence of the enzyme *zymase* in yeast to form ethanol and carbon dioxide. Write a balanced equation to represent this reaction:

Saponification: The hydrolysis of fats by basis is saponification or *soap-making*. This process was made"famous" by a scene from the(movie "Fight(Club". The main(characters in the film steal human fat from a liposuction clinic and react it with lye (NaOH) to form soap.

The reaction looks like this:(

$C_{17}H_{35}COO_3C_3H_5 + 3NaOH \rightarrow C_3H_5(OH)_3 + 3C_{17}H_{35}COONa$

The presence of the Na and the NaOH makes this reaction very recognizable! Occasionally, KOH is used instead of NaOH….

Video Lesson 8.5: Organic Reactions

Match the reaction to its name:

___ 1. Addition
___ 2. Substitution
___ 3. Combustion
___ 4. Cracking
___ 5. Polymerization
___ 6. Fermentation
___ 7. Esterification
___ 8. Saponification

a. $C_{13}H_{28} \rightarrow C_8H_{18} + C_2H_4 + C_3H_6$
b. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_20$
c. $(C_{17}H_{35}COO)_3C_3H_5 + 3\ NaOH \rightarrow C_3H_5(OH)_3 + 3C_{17}H_{35}COONa$
d. $C_6H_{12}O_6 \rightarrow C_2H_5OH + CO_2$
e. $n(CH_2CH_2) \rightarrow (CH_2CH_2)_n$
f. $C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$
g. $C_3H_6COOH + C_2H_5OH \rightarrow C_3H_6COOC_2H_5 + H_2O$
h. $C_3H_6 + I_2 \rightarrow C_3H_6I_2$

Name the reaction:

9. A saturated alkane reacts with fluorine _____
10. Small alkene chains connect to form larger alkane chains _____
11. Sugar is decomposed to form an alcohol _____
12. Large hydrocarbons are heated and break into smaller fragments _____
13. An unsaturated hydrocarbon reacts with bromine _____
14. An alcohol and an organic acid are reacted _____
15. A base is added to a fat molecule to form a soap _____
16. Hydrocarbons are burned in the presence of oxygen _____
17. Another name for hydrogenation* _____
18. Another name for halogenation* _____

Draw all organic reactants and products. Then name and give the formula for the missing substance in the reaction. Give the reaction type.

19. $C_2H_4 + F_2 \rightarrow$ _____ Rxn: _____

20. $C_3H_6 + H_2 \rightarrow$ _____ Rxn: _____

21. $C_2H_6 + Cl_2 \rightarrow$ _____ + HCl Rxn: _____

22. $C_4H_{10} + Br_2 \rightarrow$ _____ + HBr Rxn: _____

23. $CH_4 + O_2 \rightarrow$ _____ + H_2O Rxn: _____

24. $C_3H_8 + O_2 \rightarrow CO_2 +$ _____ Rxn: _____

25. $C_6H_{12}O_6 \rightarrow 2CO_2 + 2$_____ Rxn: _____

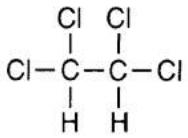
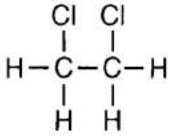
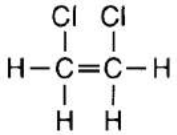
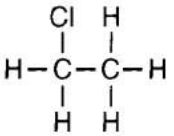
26. $C_8H_{18} \rightarrow C_6H_{12} +$ _____ Rxn: _____

27. $C_2H_5OH + C_3H_7COOH \rightarrow H_2O +$ _____ Rxn: _____

28. $C_5H_{10} + F_2 \rightarrow$ _____ Rxn: _____

8.5 Organic Reactions

1. Which formula represents the product of the addition reaction between ethene and chlorine, Cl_2 ?

1) 
2) 
3) 
4) 

2. Given the balanced equation for an organic reaction:

$C_2H_2 + 2Cl_2 \rightarrow C_2H_2Cl_4$

This reaction is best classified as

1) addition
2) esterification
3) fermentation
4) substitution

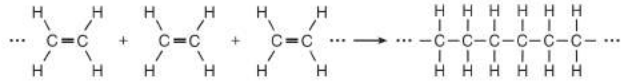
3. Given the equation:

$C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$

This reaction is best described as

1) addition involving a saturated hydrocarbon
2) addition involving an unsaturated hydrocarbon
3) substitution involving a saturated hydrocarbon
4) substitution involving an unsaturated hydrocarbon

4. Given the equation:



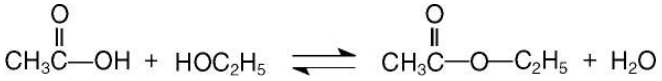
Which type of reaction is represented by this equation?

1) combustion
2) esterification
3) polymerization
4) substitution

5. The reaction that joins thousands of small, identical molecules to form one very long molecule is called

1) esterification
2) fermentation
3) polymerization
4) substitution

6. Given the reaction:



This reaction is an example of

1) fermentation
2) saponification
3) hydrogenation
4) esterification

7. When butane burns in an excess of oxygen, the principal products are

1) CO_2 and H_2O
2) CO_2 and H_2
3) CO and H_2O
4) CO and H_2

8. Which reaction results in the production of soap?

1) esterification
2) fermentation
3) polymerization
4) saponification

9. Which formula correctly represents the product of an addition reaction between ethene and chlorine?

1) CH_2Cl_2
2) CH_3Cl
3) $C_2H_4Cl_2$
4) C_2H_3Cl

10. Given the balanced equation representing a reaction:

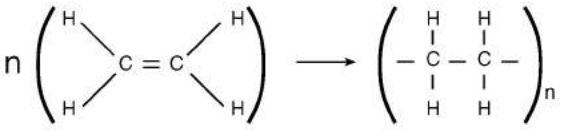
$CH_3CH_2CH_3 + Br_2 \rightarrow CH_3CH_2CH_2Br + HBr$

This organic reaction is best classified as

1) an addition reaction
2) an esterification reaction
3) a polymerization reaction
4) a substitution reaction

11. Which type of reaction is represented by the equation below?

Note: n and n are very large numbers equal to about 2000.



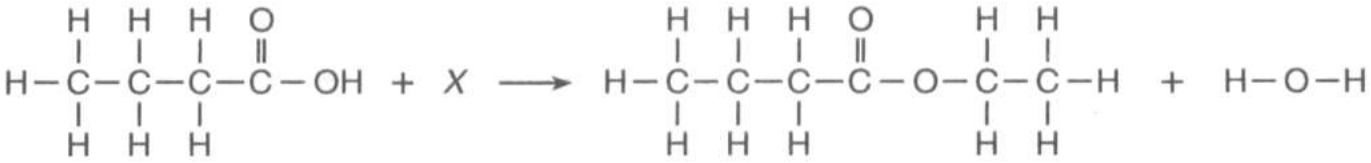
1) esterification
2) fermentation
3) saponification
4) polymerization

12. Which reaction produces ethanol?

1) combustion
2) esterification
3) fermentation
4) polymerization

Base your answers to questions **13** and **14** on the following information.

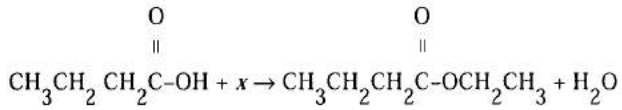
The equation below represents the reaction between butanoic acid and an unidentified reactant, *X*.



___ 13 Draw a structural formula for the unidentified reactant, *X*, in the equation.

___ 14. Identify the type of organic reaction represented by the equation.

___ 15. Given the incomplete reaction:



Which compound is represented by *x*?

1) CH_3CH_2OH 2) $CH_3C{-}H$

3) $CH_3OCH_2CH_3$ 4) CH_3CCH_3 (O)

___ 16. What are the two main products of a fermentation reaction?

1) ethanol and carbon dioxide
2) ethanol and water
3) sugar and carbon dioxide
4) sugar and water

___ 17. Which reaction best represents the complete combustion of ethene?

1) $C_2H_4 + HCl \rightarrow C_2H_5Cl$
2) $C_2H_4 + Cl_2 \rightarrow C_2H_4Cl_2$
3) $C_2H_4 + 3\ O_2 \rightarrow 2\ CO_2 + 2\ H_2O$
4) $C_2H_4 + H_2O \rightarrow C_2H_5OH$

Organic Review

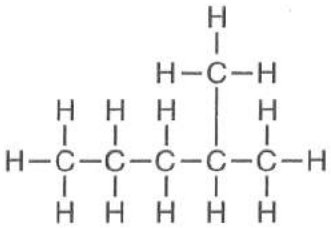
1. Which compound is a saturated hydrocarbon?

   A) propanal B) propane
   C) propene D) propyne

2. Which compound is a member of the same homologous series as C_3H_8?

   A) CH_4 B) C_4H_8
   C) C_5H_8 D) C_5H_{10}

3. What is the IUPAC name of the organic compound that has the formula shown below?



   A) 1,1-dimethylbutane
   B) 2-methylpentane
   C) hexane
   D) 4-methylpentane

4. Hydrocarbons are compounds that contain

   A) carbon, only
   B) carbon and hydrogen, only
   C) carbon, hydrogen, and oxygen, only
   D) carbon, hydrogen, oxygen, and nitrogen, only

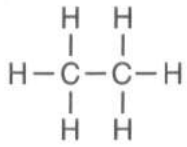
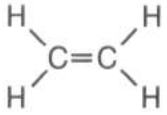
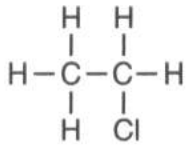
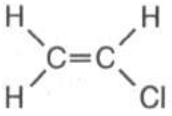
5. A molecule of a compound contains a total of 10 hydrogen atoms and has the general formula C_nH_{2n+2}.
   Which prefix is used in the name of this compound?

   A) but- B) dec- C) oct- D) pent-

6. Which compound is a saturated hydrocarbon?

   A) CH_2CH_2 B) CH_3CH_3
   C) CH_3CHO D) CH_3CH_2OH

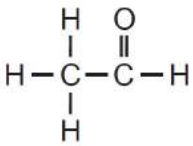
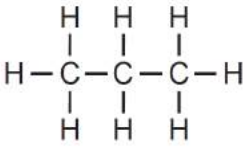
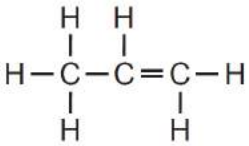
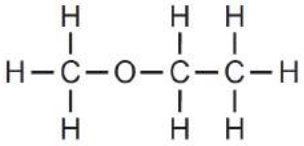
7. Which formula represents an unsaturated hydrocarbon?

   A) 
   B) 
   C) 
   D) 

8. Which formula represents an unsaturated hydrocarbon?

   A) C_5H_{12} B) C_6H_{14}
   C) C_7H_{16} D) C_8H_{14}

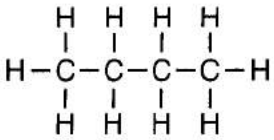
9. Which formula represents an unsaturated hydrocarbon?

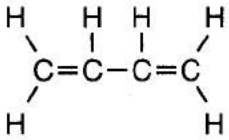
   A) 
   B) 
   C) 
   D) 

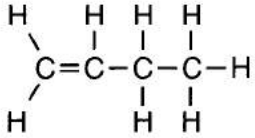
10. A straight-chain hydrocarbon that has only one double bond in each molecule has the general formula

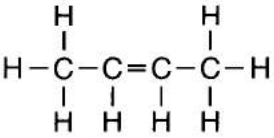
    A) C_nH_{2n-6} B) C_nH_{2n-2}
    C) C_nH_{2n} D) C_nH_{2n+2}

11. Which formula represents 2-butene?

A) 

B) 

C) 

D) 

12. A carbon-carbon triple bond is found in a molecule of

A) butane B) butanone
C) butene D) butyne

13. Which compound is an alkyne?

A) C_2H_2 B) C_2H_4
C) C_4H_8 D) C_4H_{10}

14. Which general formula represents the compound CH_3CH_2CCH?

A) C_nH_n B) C_nH_{2n}
C) C_nH_{2n-2} D) C_nH_{2n+2}

15. Which compound is an unsaturated hydrocarbon?

A) hexanal B) hexane
C) hexanoic acid D) hexyne

16. Which element is present in all organic compounds?

A) carbon B) hydrogen
C) nitrogen D) oxygen

17. Butanal and butanone have different chemical and physical properties primarily because of differences in their

A) functional groups
B) molecular masses
C) molecular formulas
D) number of carbon atoms per molecule

18. Ethanol and dimethyl ether have different chemical and physical properties because they have different

A) functional groups
B) molecular masses
C) numbers of covalent bonds
D) percent compositions by mass

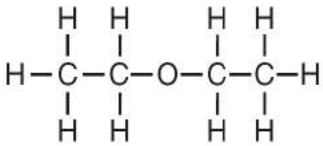
19. Which two compounds have the same molecular formula but different chemical and physical properties?

A) CH_3CH_2Cl and CH_3CH_2Br
B) CH_3CHCH_2 and $CH_3CH_2CH_3$
C) CH_3CHO and CH_3COCH_3
D) CH_3CH_2OH and CH_3OCH_3

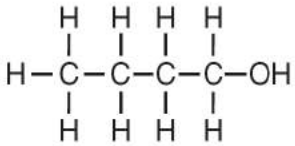
20. The isomers butane and methylpropane differ in their

A) molecular formulas
B) structural formulas
C) total number of atoms per molecule
D) total number of bonds per molecule

21. Given the formulas for two compounds:



and

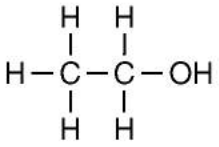


These compounds differ in

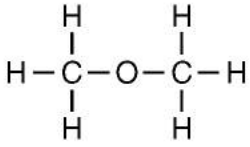
A) gram-formula mass
B) molecular formula
C) percent composition by mass
D) physical properties at STP

22. Given the structural formulas:

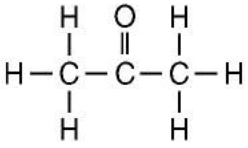
Formula A



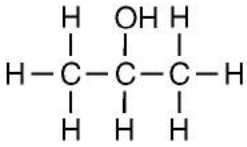
Formula B



Formula C



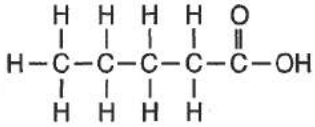
Formula D



Which two formulas represent compounds that are isomers of each other?

A) *A* and *B* B) *A* and *C* C) *B* and *D* D) *C* and *D*

23. Given the formula for an organic compound:



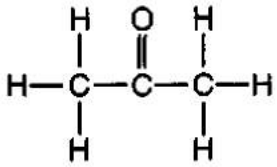
This compound is classified as an

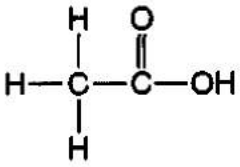
A) aldehyde B) amine
C) ester D) organic acid

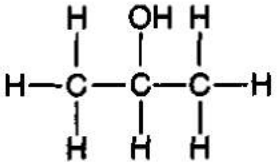
24. What is the total number of carbon atoms in a molecule of ethanoic acid?

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

25. Given the three organic structural formulas shown below:



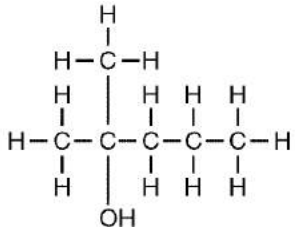


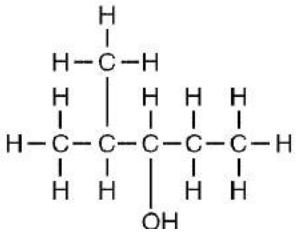


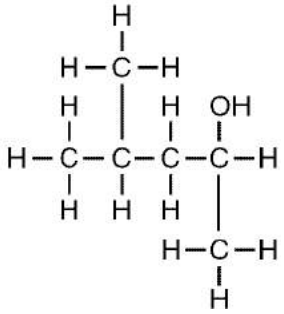
Which organic compound classes are represented by these structural formulas, as shown from left to right?

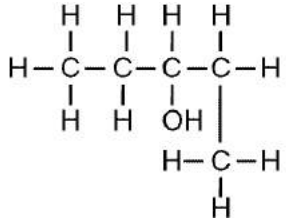
A) ester, organic acid, ketone B) ester, aldehyde, organic acid
C) ketone, aldehyde, alcohol D) ketone, organic acid, alcohol

26. Which structural formula is correct for 2-methyl-3-pentanol?

A) 

B) 

C) 

D) 

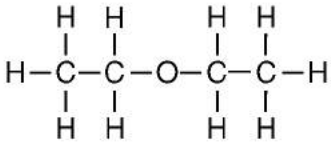
27. What is the total number of pairs of electrons shared between the carbon atom and the oxygen atom in a molecule of methanal?

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

28. The reaction between an organic acid and an alcohol produces

A) an aldehyde B) a ketone
C) an ether D) an ester

29. Given the structural formula:



The compound represented by this formula can be classified as an

A) organic acid B) ether
C) ester D) aldehyde

30. Given the balanced equation for an organic reaction:

$$C_2H_2 + 2Cl_2 \rightarrow C_2H_2Cl_4$$

This reaction is best classified as

A) addition B) esterification
C) fermentation D) substitution

31. Given the equation:

$$C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$$

This reaction is best described as

A) addition involving a saturated hydrocarbon
B) addition involving an unsaturated hydrocarbon
C) substitution involving a saturated hydrocarbon
D) substitution involving an unsaturated hydrocarbon

32. The reaction that joins thousands of small, identical molecules to form one very long molecule is called

A) esterification B) fermentation
C) polymerization D) substitution

33. Given the reaction:

This reaction is an example of

A) fermentation B) saponification C) hydrogenation D) esterification

34. What are the two main products of a fermentation reaction?

A) ethanol and carbon dioxide
B) ethanol and water
C) sugar and carbon dioxide
D) sugar and water

35. Which reaction results in the production of soap?

A) esterification B) fermentation
C) polymerization D) saponification

Base your answers to questions **36** through **38** on the information below.

Gasoline is a mixture composed primarily of hydrocarbons such as isooctane, which is also known as 2,2,4-trimethylpentane.

Gasoline is assigned a number called an octane rating. Gasoline with an octane rating of 87 performs the same as a mixture that consists of 87% isooctane and 13% heptane.

An alternative fuel, E-85, can be used in some automobiles. This fuel is a mixture of 85% ethanol and 15% gasoline.

36. In the space below, draw a structural formula for a molecule of 2,2,4-trimethylpentane.

37. State the octane rating of a gasoline sample that performs the same as a mixture consisting of 92% isooctane and 8% heptane.

38. Identify the functional group in a molecule of ethanol in the alternative fuel E-85.

Base your answers to questions **39** through **43** on the information below.

Biodiesel is an alternative fuel for vehicles that use petroleum diesel. Biodiesel is produced by reacting vegetable oil with CH_3OH. Methyl palmitate, $C_{15}H_{31}COOCH_3$, a compound found in biodiesel, is made from soybean oil. One reaction of methyl palmitate with oxygen is represented by the balanced equation below.

$$2C_{15}H_{31}COOCH_3 + 49O_2 \rightarrow 34CO_2 + 34H_2O + \text{energy}$$

39. Identify the type of organic reaction represented by the balanced equation.

40. Identify the class of organic compounds to which methyl palmitate belongs.

41. Explain, in terms of *both* atoms and molecular structure, why there is no isomer of CH_3OH.

42. Write the IUPAC name for the compound that reacts with vegetable oil to produce biodiesel.

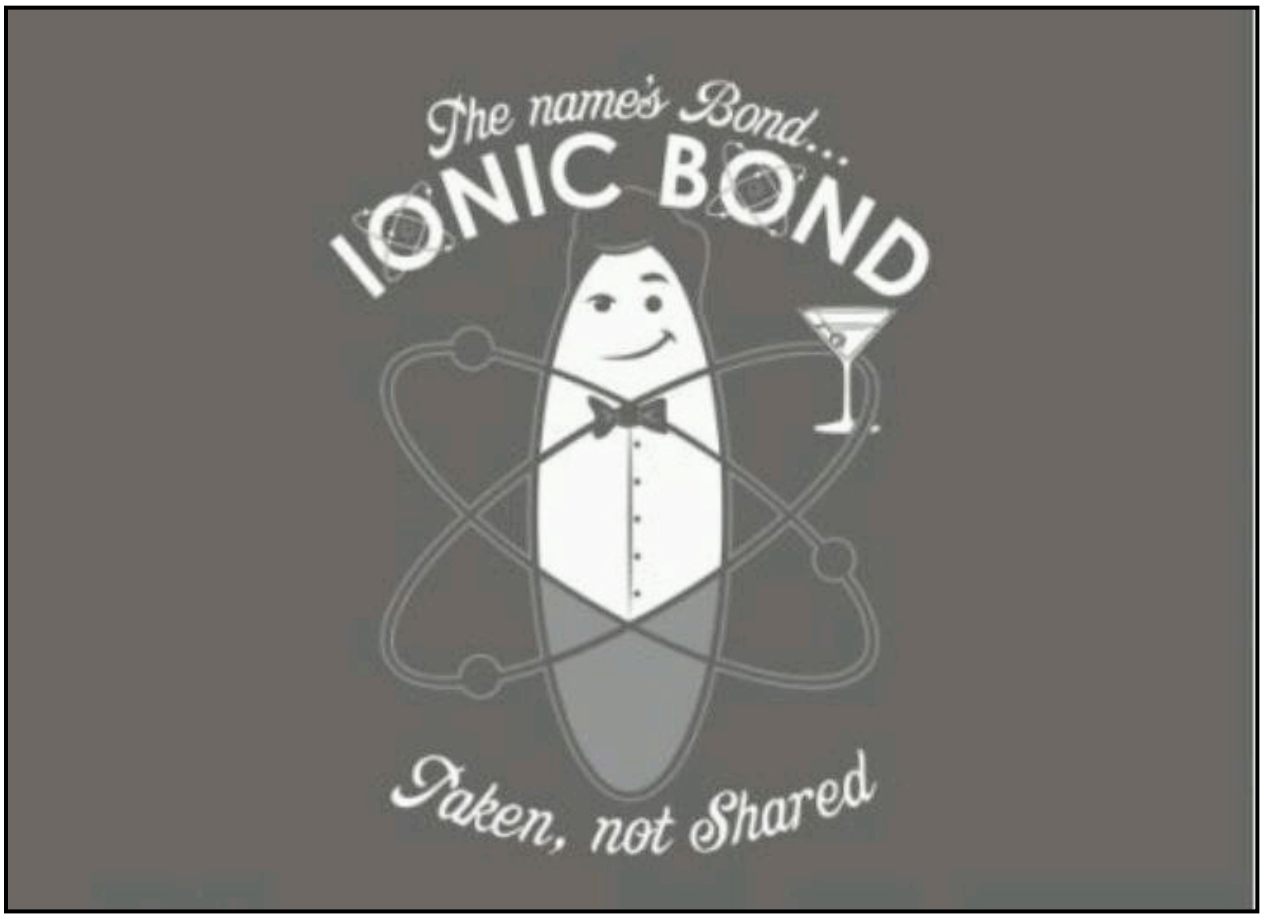
43. State evidence from the balanced equation that indicates the reaction is exothermic.

Name:

Regents Chemistry:

Practice Packet

Chapter 9: Chemical Bonding



Chapter 9: Bonding Vocabulary

Molecule – a COVALENTLY bonded substance; can be atoms of the same element (Ex: diatomic elements/molecules); molecular substance = covalent substanc

Compound – a substance composed of two or more atoms from different elements CHEMICALLY bonded together

Bond – forces of attraction that hold atoms together in a molecule or compound

Octet Rule – atoms bond together in order to have 8 electrons in their valence shell

Exothermic – energy is RELEASED as a product of a chemical reaction

Endothermic – energy is CONSUMED as a product of a chemical reaction

Ionic Bond – chemical bond involving the TRANSFER of electrons between a metal and nonmetal atom (metals lose, nonmetals gain); electronegativity difference between elements typically GREATER than 1.7

Covalent Bond – chemical bond involving the SHARING of electrons between two nonmetal atoms; electronegativity difference between elements typically LESS than 1.7

Oxidation number – the "charge" an element has within a compound

Polyatomic ions – atoms of two or more elements chemically bonded together and having a NET CHARGE

Polar molecule – a covalent molecule with an unequal sharing of electrons; contains atoms of two different nonmetal elements (all covalent compounds that are NOT diatoms)

Nonpolar molecules – a molecule with symmetrical/equal sharing of electrons

Intermolecular forces (IMF's) – weak forces between molecules that hold the molecules to one another; not actually chemical bonds

Chemical Bonding

Video Lesson 91.

Objectives

- Describe the 2 major types of chemical bonds in terms of electrons.
- Describe the properties of ionic and covalent bonding

Chemical Bonds

- Chemical bond is the force between atoms or ions
- Atoms will **gain**, **lose**, or **share electrons** to achieve the same electron configurations as the noble gases

Bond Energy – Bond Formation

- Bond formation is spontaneous
  - Energy is release when bonds are formed
    - EXOTHERMIC
    - Atoms go from high energy to lower energy
    - Creating bonds creates stability

$$A + B \rightarrow AB + \text{Energy (heat) (560 kJ)}$$

Bond Energy – Bond Breaking

- Breaking bonds is not spontaneous
  - Energy is required/absorbed/consume
  - ENDOTHERMIC

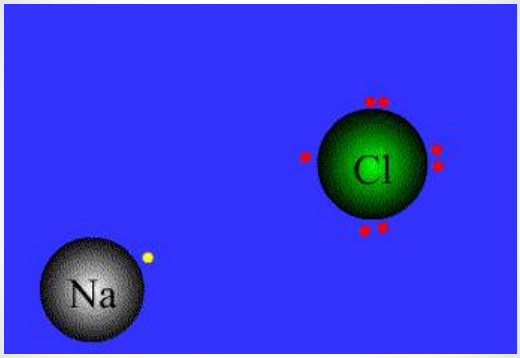
$$BC + \text{energy (heat) (600 kJ)} \rightarrow B + C$$

Bond Types

- Ionic Bonds
  - Bonds that involve IONS
- Covalent Bonds
  - Bonds between nonmetals
  - Molecules
- Metallic Bonds
  - Bonds that hold metals together.

Ionic Bonds

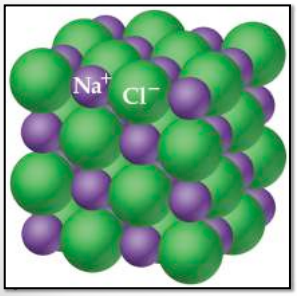
- Bond formed between ions
  - Metals bonded to nonmetals
  - Compound that contain polyatmoic ions
- Involves **a transfer of electrons**
- Electronegativity difference typically greater than 1.7 (subtraction)
  - The greater the difference the more ionic character



Properties of Ionic Compounds

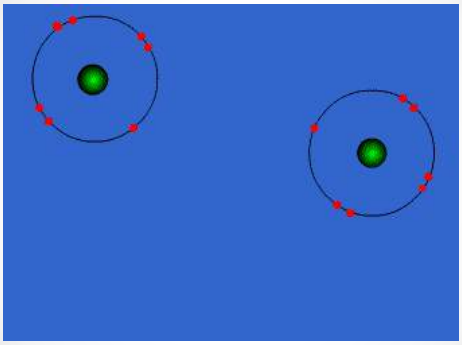
- Very strong bond
- High melting points and high boiling points
- Hard
- Form crystal lattice shape (regular geometric pattern)
- Electrolytes
  - Can conduct electricity in solution (aq)

Crystalline Structure



Covalent Bonds

- Found in molecular substances.
- **Involves a sharing of electron pairs**.
- Nonmetals bonded to nonmetals
  - Or diatomic elements
    - Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2
- Similar electronegativities



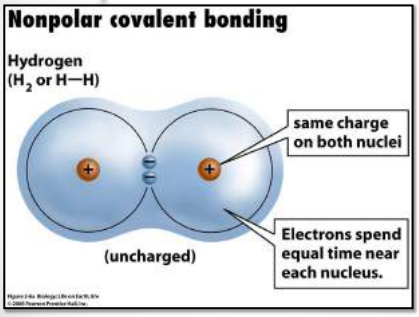
Properties of Covalent Bonds

- Weaker bonds than ionic bonds
- Low melting and low boiling points
- Do not form electrolytes
- Typically soft

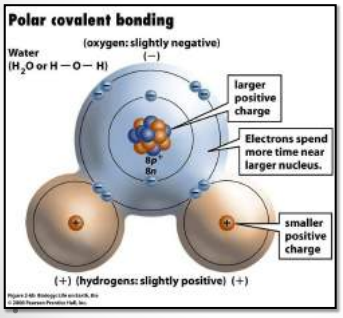
Types of Covalent Bonds

Nonpolar Covalent Bonds	Polar Covalent Bonds
• Form between atoms with the same electronegativity • Electrons are shared equally • H_2	• Form between nonmetals with different electronegativity • Electrons are shared unequally

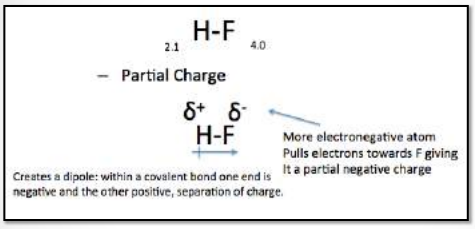
Nonpolar Covalent Bonds



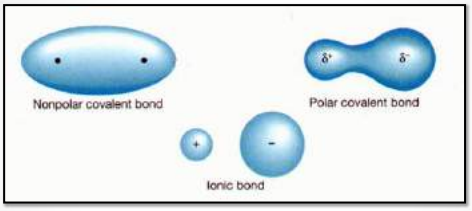
Polar Covalent Bond



Partial Charge



Summary



Lewis Dot Diagrams

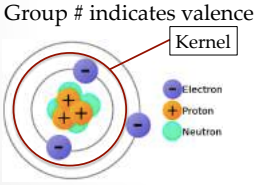
Chemistry 200
Video Lesson 9.2

Objective:

How do we show the arrangement of electrons for Ionic and Molecular substances using Lewis Dot Diagrams? (Lewis Structures)

Valence electrons – electrons in the outermost shell of an atom or ion

- Establish chemical characteristics of elements
- The only electrons shown in Lewis electron dot structures
- Group # indicates valence e^- except for Helium

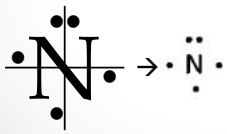


Lewis Structures or Dot Diagrams

Atoms

- Determine valence electron number
- Write the symbol for the element
- Place a dot on each side of the symbol for each electron

Nitrogen → 5 valence e^-



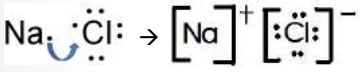
Ions

- ***Cations***(+) lose valence e^- & ***Anions***(-) gain valence e^-
- Place [ ] around the symbol & indicate the charge :

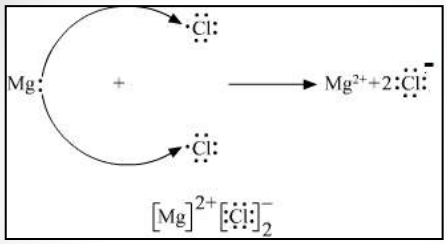
Creating Lewis Dot Diagrams for molecules

Ionic

- create dot diagrams for each ion and put them together
- use subscripts if more than one cation or anion



Lewis dot diagram for **$MgCl_2$**



Molecular

Step 1

Obtain the sum of the **valence** electrons from all of the atoms. Do not worry about keeping track of which electrons come from which atoms. It is the total number of valence electrons that is important. Be attentive to the charge if applicable.

Step 2

Use one pair of electrons to form a bond btwn each pair of bound atoms. For convenience, a line (instead of a pair of dots) is generally used to indicate each pair of bonding electrons. The atom w/ the smallest electronegativity is usually in the middle. Oxygen tends not to be the central atom. Hydrogen is **never** the central atom because it only forms one bond.

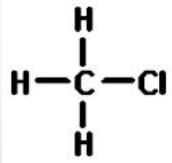
Step 3

Arrange the remaining electrons to satisfy the **duet** rule for hydrogen & the **octet** rule for each of the other atoms. If each atom does not have an octet of electrons around it & there are still electrons to be assigned, consider a multiple bond.

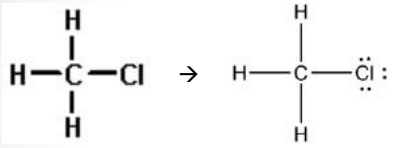
Draw a Lewis Dot structure for **CH_3Cl**

1. Total number of valence e^- for the atoms in the compound.
   C = 4
   3H = 3 → Total 14 e^-
   Cl = 7

   must have this amount on your diagram when completed
2. Put most electronegative element in the center & connect all atoms to it using bond lines, one line for each e^- pair.



3. Complete diagram using the Octet & Duet Rules



Draw a Lewis dot diagram for O_2

Each oxygen atom has 6 valence e^-, therefore O_2 has a total of 12 valence e^-.

1. O – O This leaves 10 e^- left to use
2.  There are a total of 12 valence e^- which is all that can be used. Apply Octet rule.
3.  Now we have an Octet around each oxygen atom

ALWAYS COUNT BONDS & ELECTRON PAIRS WHEN FINISHED!!! PLEASE

Molecular Shapes & Polarity

Chemistry 200
Video Lesson 9.3

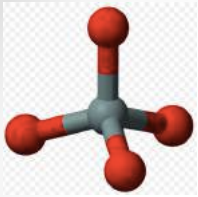
Objective:

How do we determine the shape and polarity of a molecule?

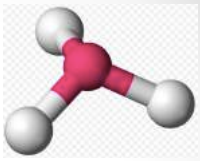
Shapes of Molecules (VSEPR - Valence Shell Electron Pair Repulsion)

VSEPR Theory is just a fancy model used to identify the shape of a molecule in 3 dimensions. The theory is based on the fact that atoms AND unbonded (lone) pairs of electrons found on the CENTRAL atom repel each other. As a result, the 3 dimensional shape of a molecule is simply the result of electron clouds getting as far away from each other as possible while still being bonded to a central atom. This should make sense since electrons are negative & repel each other

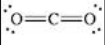
Name of Shape	Number of atoms bonded to the central atom	Number of unshared pairs of electrons on the central atom	Shape	Examples
Tetrahedral	4	0		CH_4 CH_3I CCl_4
Trigonal Pyramidal	3	1		NH_3 PH_3



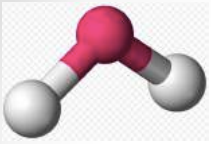
Tetrahedral



Trigonal Pyramidal

Name of Shape	Number of atoms bonded to the central atom	Number of unshared pairs of electrons on the central atom	Shape	Examples
Bent	2	2		H_2O H_2S
Linear	2	0		CO_2

* H-F is linear



Bent

Linear

Polarity within a Molecule

Polar Molecule: A molecule that has an overall slight (+) & (-) side. For a molecule to be polar, it must meet 2 important criteria:

1. It must contain polar covalent bonds
   AND
2. It must have an asymmetrical (uneven) charge distribution

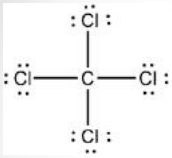
Failing either of these (or both) means the molecule is **NONPOLAR**

Nonpolar Molecule: A molecule that lacks an overall (+) &(-) side. Nonpolar molecules are created when a molecule lacks polar bonds or has a symmetrical (even) charge distribution

Determinining Polarity w/in a Molecule

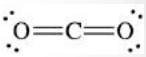
- Ionic molecules are always polar
- look at the shape & central atom of the molecule
- the shape **MUST** permit a net displacement of charge (one end positive & one end negative) to be polar
- if all polar bonds w/in a molecule are **equal**, the molecule is nonpolar

Nonpolar Molecule



Polar bonds

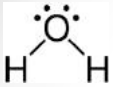
Symmetrical charge distribution

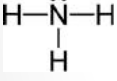


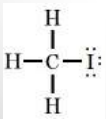
Nonpolar bonds

Symmetrical charge distribution

Polar Molecules







Polar bonds

Asymmetrical charge distribution

Intermolecular Forces

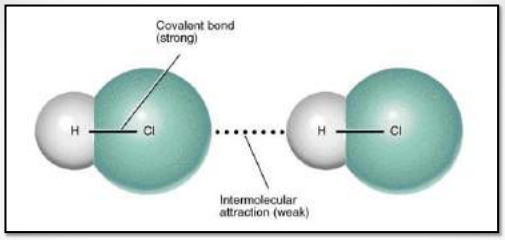
Video Lesson 9.4

Objectives

- Evaluate the strength and type of intermolecular forces of attraction.

Intermolecular Forces (IMF's)

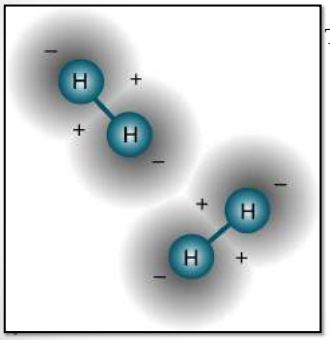
- Only in covalent molecules
- Weak forces that act BETWEEN molecules
- Only exist in Gas and liquid phase
- Much weaker than chemical bonds
- IMF's ARE NOT BONDS!!!



Types of Intermolecular Forces

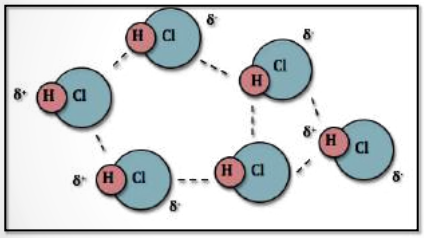
- Van der Waal's Attraction
  - Weakest of All IMF's
  - Nonpolar molecules only
  - H-H
- Dipole – Dipole Interaction
  - Two poles – positive and negative
  - Hydrogen Bonding
    - Strongest
    - HFON

Van der Waals Attraction

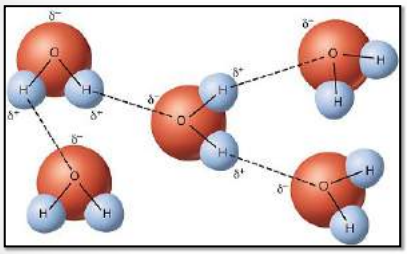


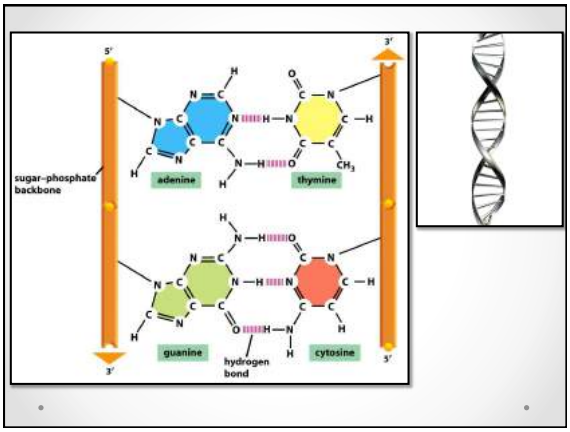
Temporary dipoles

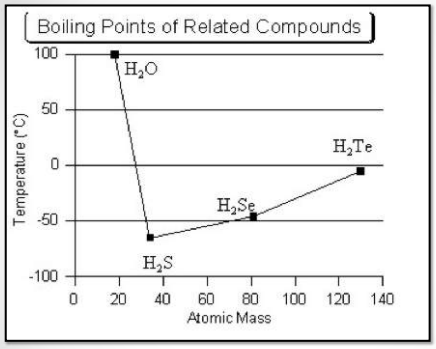
Dipole – Dipole Interactions



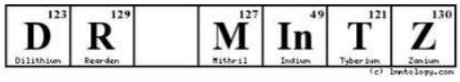
Hydrogen Bonding







Sketch Notes



Video Lesson 9.1: Bonding

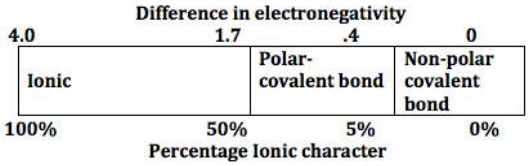
Ionic Bond	between a Metal and Non-Metal	(M + NM)
Covalent Bond	between a Non-Metal and Non-Metal	(NM + NM)
Metallic Bond	between a Metal and Metal	(M+ M)

Determine if the elements in the following compounds are metals or non-metals. Describe the type of bonding that occurs in the compound.

Compound	**Element 1 (metal or non-metal?)**	**Element 2 (metal or non-metal?)**	**Bond Type**
NO_2	N = non-metal	O = non-metal	covalent
NaCl			
SO_2			
PO_4^{3-}			
$MgBr_2$			
CaO			
H_2O			
K_2O			
Cu-Zn alloy			
O_2			
$CuCl_2$			
NO_2^-			
TiO_2			
HF			
Rb_2S			
Au-Ag mixture			
Fe_2O_3			

Types of Chemical Bonds:
Classify the following compounds as ionic (metal + nonmetal), covalent (nonmetal + nonmetal) or both (compound containing a polyatomic ion).

1. $CaCl_2$ _____
2. CO_2 _____
3. H_2O _____
4. $BaSO_4$ _____
5. K_2O _____
6. NaF _____
7. Na_2CO_3 _____
8. CH_4 _____
9. SO_3 _____
10. LiBr _____
11. MgO _____
12. NH_4Cl _____
13. HCl _____
14. KI _____
15. NaOH _____
16. NO_2 _____
17. $AlPO_4$ _____
18. $FeCl_3$ _____
19. P_2O_5 _____
20. N_2O_3 _____



Bonding between	More Electronegative element and value	Less Electronegative element and value	Difference in electronegativity	Bond Type
Sulfur and Hydrogen				
Sulfur and cesium				
Chlorine and bromine				
Calcium and chlorine				
Oxygen and hydrogen				
Nitrogen and hydrogen				
Iodine and iodine				
Copper and sulfur				
Hydrogen and fluorine				
Carbon and oxygen				

Bond Types

1. Two molecules of HBr collide and then form H_2 and Br_2. During the collision, the bonds in the HBr molecules are

   1) broken as energy is absorbed
   2) broken as energy is released
   3) formed as energy is absorbed
   4) formed as energy is release

2. Given the balanced equation representing a reaction:

   $$O_2 \rightarrow O + O$$

   What occurs during this reaction?

   1) Energy is absorbed as bonds are broken.
   2) Energy is absorbed as bonds are formed.
   3) Energy is released as bonds are broken.
   4) Energy is released as bonds are formed.

3. What occurs as two atoms of fluorine combine to become a molecule of fluorine?

   1) A bond is formed as energy is absorbed.
   2) A bond is formed as energy is released.
   3) A bond is broken as energy is absorbed.
   4) A bond is broken as energy is released.

4. Which formulas represent one ionic compound and one molecular compound?

   1) N_2 and SO_2
   2) Cl_2 and H_2S
   3) $BaCl_2$ and N_2O_4
   4) $NaOH$ and $BaSO_4$

5. Which element forms an ionic compound when it reacts with lithium?

   1) K 2) Fe 3) Kr 4) Br

6. Which type of substance can conduct electricity in the liquid phase but *not* in the solid phase?

   1) ionic compound
   2) molecular compound
   3) metallic element
   4) nonmetallic element

7. A molecular compound is formed when a chemical reaction occurs between atoms of

   1) chlorine and sodium
   2) chlorine and yttrium
   3) oxygen and hydrogen
   4) oxygen and magnesium

8. Which compound has both ionic and covalent bonding?

   1) $CaCO_3$ 2) CH_2Cl_2
   3) CH_3OH 4) $C_6H_{12}O_6$

9. Which formula represents a molecular compound?

   1) Kr 2) LiOH
   3) N_2O_4 4) NaI

10. Which characteristic is a property of molecular substances?

    1) good heat conductivity
    2) good electrical conductivity
    3) low melting point
    4) high melting point

11. Which type of bond is found between atoms of solid cobalt?

    1) nonpolar covalent
    2) polar covalent
    3) metallic
    4) ionic

12. A solid substance is an excellent conductor of electricity. The chemical bonds in this substance are most likely

    1) ionic, because the valence electrons are shared between atoms
    2) ionic, because the valence electrons are mobile
    3) metallic, because the valence electrons are stationary
    4) metallic, because the valence electrons are mobile

Video Lesson 9.2: Lewis Structures

Lewis Dot Diagrams for Molecular substances

Step 1
Obtain the sum of the **valence** electrons from all of the atoms. Do not worry about keeping track of which electrons come from which atoms. It is the total number of valence electrons that is important. Be attentive to the charge if applicable.

Step 2
Use one pair of electrons to form a bond btwn each pair of bound atoms. For convenience, a line (instead of a pair of dots) is generally used to indicate each pair of bonding electrons. The atom w/ the smallest electronegativity is usually in the middle. Oxygen tends not to be the central atom. Hydrogen is **never** the central atom because it only forms one bond.

Step 3
Arrange the remaining electrons to satisfy the **duet** rule for hydrogen & the **octet** rule for each of the other atoms. If each atom does not have an octet of electrons around it & there are still electrons to be assigned, consider a multiple bond.

Draw Lewis Dot Diagrams for the following substances:

CH_4	**CH_3Cl**	**NH_3**	**H_2O**
HBr	**F_2**	**O_2**	**N_2**
CO_2	**C_2H_2**	**CCl_4**	**PCl_3**

PO_4^{-3} **ClO_3^{-1}** **NH_4^{+1}**

NaF **Li_2S** **$MgCl_2$** **Al_2O_3**

Video Lesson 9.3: Molecular Shapes and Polarity

Read This!

The **VSEPR (Valence Shell Electron Pair Repulsion) Theory** helps predict the shapes of molecules and is based on the premise that electrons around a central atom repel each other. **Electron domains** are areas of high electron density such as bonds (single, double or triple) and lone-pairs of electrons. In simple terms VSEPR means that all electron bonding domains and electron nonbonding domains around a central atom need to be positioned as far apart as possible in *three-dimensional* space.

1. VSEPR theory specifies "valence shell" electrons. Explain why these are the most critical electrons for determining molecular shape.

2. In the VSEPR theory, what is repelling what?

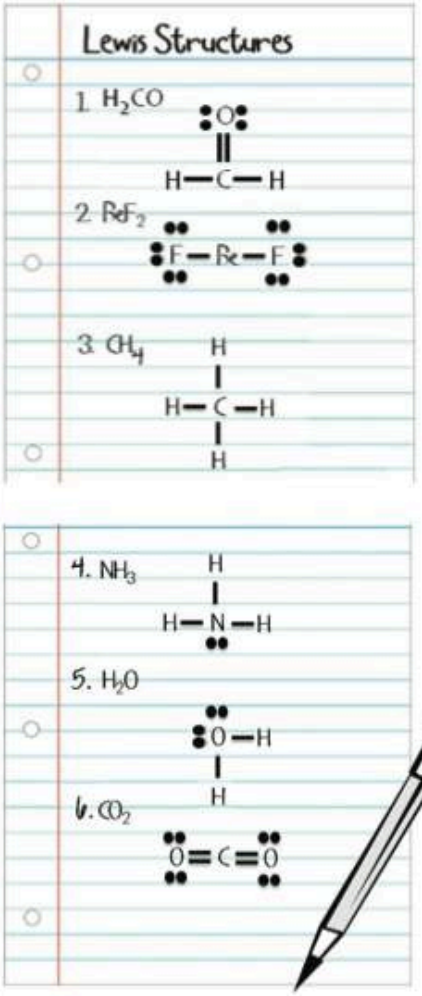
Based on the information in the *Read This!* section, sketch one of the molecular shapes shown below in each of the boxes provided in Model 1.

Three-Dimensional Molecular Shapes	Linear 180°	Trigonal planar 120°
Tetrahedral 109.5°	Pyramidal 107°	Bent 104.5°

Why?

When you draw a Lewis structure for a molecule on paper, you are making a two-dimensional representation of the atoms. In reality however, molecules are not flat—they are *three*-dimensional. The true shape of a molecule is important because it determines many physical and chemical properties for the substance. In this activity you will learn how to predict molecular shapes.

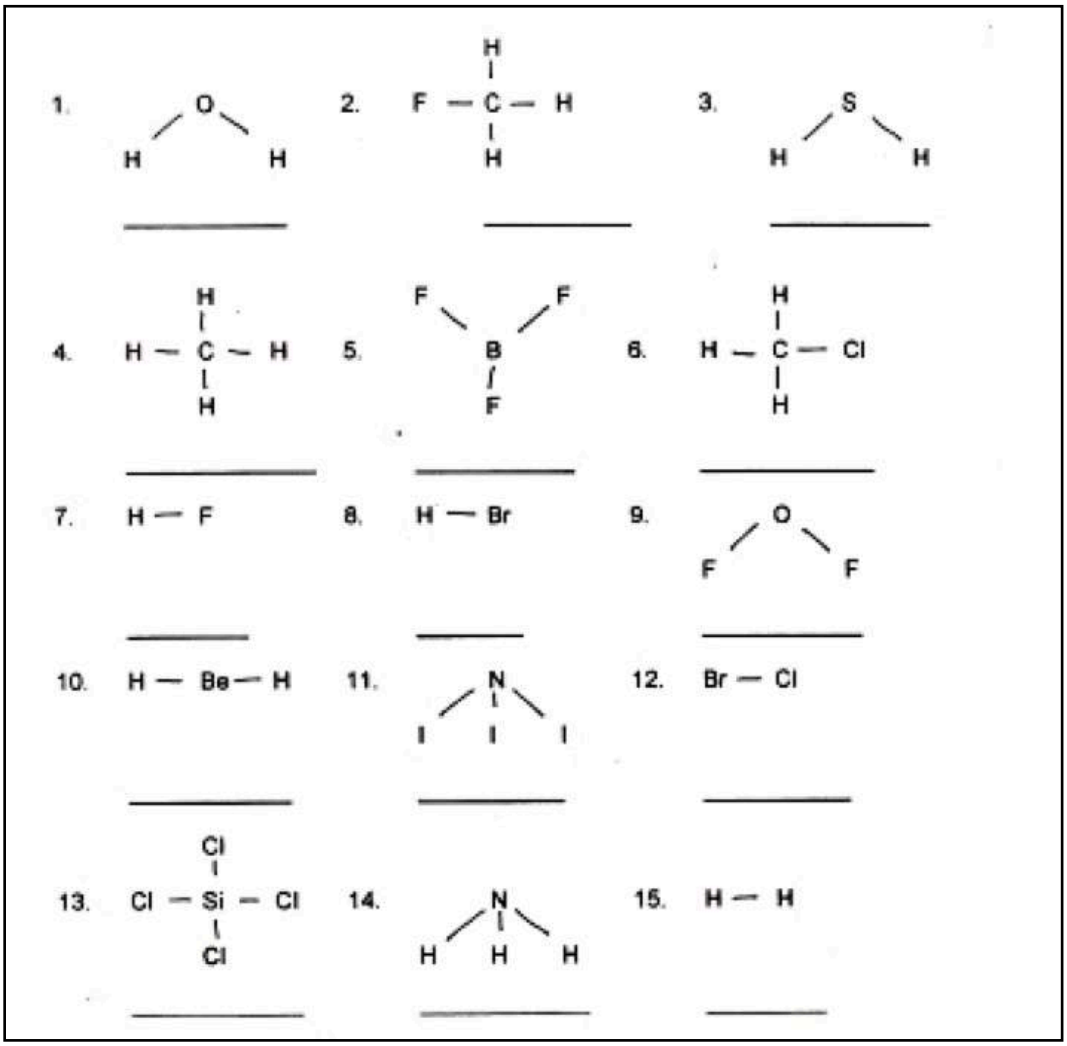
Model 1 – Lewis Structures



Lone pair = ••

H_2CO 3 electron domains (3 bonding, 0 nonbonding)	**3-D Molecular Shape**
BeF_2 2 electron domains (2 bonding, 0 nonbonding)	
CH_4 4 electron domains (4 bonding, 0 nonbonding)	
NH_3 4 electron domains (3 bonding, 1 nonbonding)	
H_2O 4 electron domains (2 bonding, 2 nonbonding)	
CO_2 2 electron domains (2 bonding, 0 nonbonding)	

For each molecule below, identify which are polar molecules and which are nonpolar molecules.



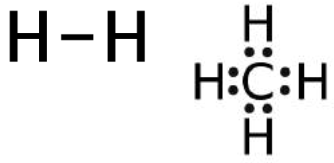
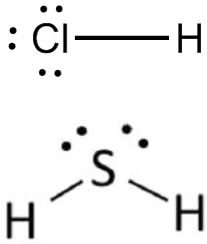
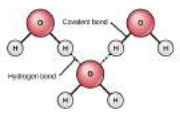
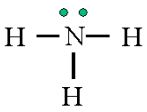
Draw Lewis Dot Diagrams, determine molecular polarity and bond type for each of the following substances:

N_2	HF	H_2O

CO_2	H_2S	NH_3
CH_4	CH_3Cl	O_2
PO_4^{-3}	ClO_3^{-1}	NH_4^{+1}

Video Lesson 9.4: Intermolecular Forces

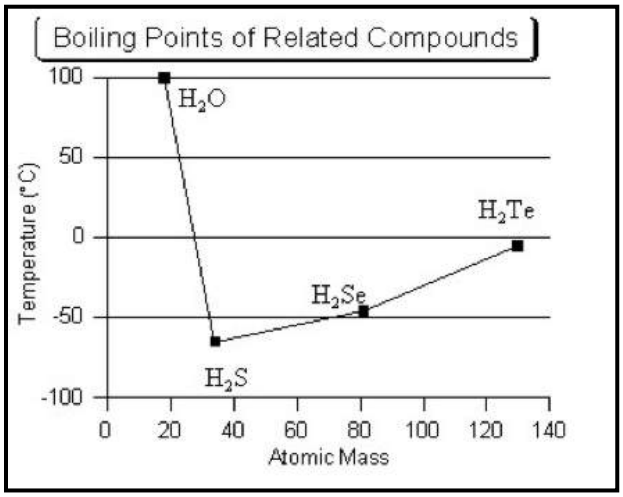
Intermolecular Forces of Attraction Summary

Van der Waal's Attraction	Dipole- Dipole Attraction	Hydrogen Bonds
(weak)	(medium)	(very strong)
• Weak attractions found between nonpolar molecules • Temporary dipole due to asymmetrical distribution of electrons • The electron cloud is always moving • Examples 	• Attraction between polar molecules that occur with oppositely charged regions of the neighboring molecule • Examples 	• FON • Very strong attraction between molecules where the hydrogen atom of one molecule is attracted to the F, O or N atom in another molecule • Responsible for the high BP of water • Examples  

The graph on the right shows the boiling point of compounds of hydrogen and them members of group 16.

1. What is the electronegativity difference in each compound?

2. How can the differences in boiling point be explained?



Intramolecular Bonds

(within a molecule)

TYPES of SUBSTANCES

IONIC BONDS

- Bonds formed by the transfer of electrons
- Generally have an EN difference of more than 1.7
- Generally compounds of metals & nonmetals
- Hard, crystalline solide w/ high M.P. Poor conductors as solids, good conductors in aqueous soln.
- Ex: $NaCl$, CuF_2, Na_2SO_4

COVALENT BONDS

Polar

- Formed by unequal sharing of electrons
- EN diff. btwn 0 & 1.7

Nonpolar

- Formed by equal sharing of electrons
- EN diff. of 0

- Both kinds of covalent bonds are found in nonmetallic elements & compounds composed of nonmetals.
- When properties of substances w/covalent bonds are studied they can be divided into 2 grps

Molecular Substances

At room temp. are gases, liquids or solids w/ low M.P.
Poor conductors always ex: H_2O, H_2, HCl, CH_4, CH_3Cl

METALLIC BONDS

- Bonds formed by the extreme mobility of electrons.
- A metal can be pictured as a collection of positive ions in a ***sea of mobile electrons***
- Solids at room temp. except Hg.
- Good conductors as solids & liquids.
- Have luster
- Ex: Cu, Fe, Hg, Na

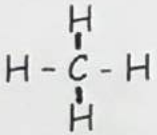
Intermolecular Forces of Attraction

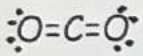
When one looks at the attraction that exists **between** molecules, they see that there are 3 different types:

Van der Waal's Attraction

(weak)

- Weak attractions found btwn nonpolar molecules or noble gases
- Temporary dipole due to asymmetrical distribution of the electrons.
- The electron cloud is always moving, (+) & (-) areas of the molecule
- Ex: H - H



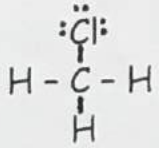


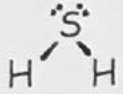
Dipole-Diploe Attraction

(dipoles are polar molecules)

(medium)

- Attraction btwn polar molecules that occur w/ oppositely charged regions of the neighboring molecules
- Ex: H - Cl



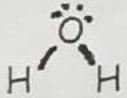


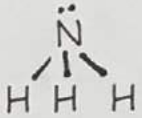
Hydrogen Bonds

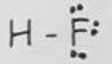
F.O.N.

(very strong)

- Very strong attraction btwn molecules where the Hydrogen atom of one molecule is attracted to the F, O or N atom in another molecule
- Responsible for the high B.P. of H_2O
- Ex:







Name: **Intermolecular Forces**

1. The boiling points, at standard pressure, of four compounds are given in the table below.

Boiling Points of Four Compounds

Compound	Boiling Point (°C)
H_2O	100.0
H_2S	−59.6
H_2Se	−41.3
H_2Te	−2.0

Which type of attraction can be used to explain the unusually high boiling point of H_2O?

1) ionic bonding
2) hydrogen bonding
3) polar covalent bonding
4) nonpolar covalent bonding

2. Hydrogen bonding is a type of

1) strong covalent bond
2) weak ionic bond
3) strong intermolecular force
4) weak intermolecular force

3. In which liquid is hydrogen bonding strongest?

1) $HF(\ell)$
2) $H_2(\ell)$
3) $CH_4(\ell)$
4) $NH_3(\ell)$

4. Which characteristic of the compound C_5H_{12} causes it to have a higher normal boiling point than C_2H_6?

1) The distance between molecules of C_5H_{12} is greater.
2) The force of attraction between molecules of C_5H_{12} is greater.
3) C_5H_{12} has a larger number of ionic bonds.
4) C_5H_{12} has a larger number of double bonds.

5. Which type of attraction results from the formation of weak momentary dipoles?

1) ionic
2) metallic
3) molecule-ion
4) van der Waals forces

6. Which statement explains why Br_2 is a liquid at STP and I_2 is a solid at STP?

1) Molecules of Br_2 are polar, and molecules of I_2 are nonpolar.
2) Molecules of I_2 are polar, and molecules of Br_2 are nonpolar.
3) Molecules of Br_2 have stronger intermolecular forces than molecules of I_2 .
4) Molecules of I_2 have stronger intermolecular forces than molecules of Br_2 .

7. Which statement explains why H_2O has a higher boiling point than N_2?

1) H_2O has greater molar mass than N_2.
2) H_2O has less molar mass than N_2.
3) H_2O has stronger intermolecular forces then N_2.
4) H_2O has weaker intermolecular forces than N_2.

8. The primary forces of attraction between water molecules in $H_2O(\ell)$ are

1) ionic bonds
2) hydrogen bonds
3) molecule-ion attractions
4) van der Waals forces

Name: **Bonding Review**

____ 1. Given the balanced equation representing a reaction:

$Cl_2 \rightarrow Cl + Cl$

What occurs during this reaction?

1) A bond is broken as energy is absorbed.
2) A bond is broken as energy is released.
3) A bond is formed as energy is absorbed.
4) A bond is formed as energy is released.

____ 2. Which statement describes what occurs as two atoms of bromine combine to become a molecule of bromine?

1) Energy is absorbed as a bond is formed.
2) Energy is absorbed as a bond is broken.
3) Energy is released as a bond is formed.
4) Energy is released as a bond is broken.

____ 3. Which of these elements has an atom with the most stable outer electron configuration?

1) Ne 2) Cl 3) Ca 4) Na

____ 4. When a sodium atom reacts with a chlorine atom to form a compound, the electron configurations of the ions forming the compound are the same as those in which noble gas atoms?

1) krypton and neon
2) krypton and argon
3) neon and helium
4) neon and argon

____ 5. Which element has an atom with the greatest attraction for electrons in a chemical bond?

1) As 2) Bi 3) N 4) P

____ 6. Based on electronegativity values, which type of elements tends to have the greatest attraction for electrons in a bond?

1) metals
2) metalloids
3) nonmetals
4) noble gases

____ 7. Which term indicates how strongly an atom attracts the electrons in a chemical bond?

1) alkalinity
2) atomic mass
3) electronegativity
4) activation energy

____ 8. Which bond is *least* polar?

1) As–Cl
2) Bi–Cl
3) P–Cl
4) N–Cl

____ 9. Given the electron dot diagram:

The electrons in the bond between hydrogen and fluorine are more strongly attracted to the atom of

1) hydrogen, which has the higher electronegativity
2) fluorine, which has the higher electronegativity
3) hydrogen, which has the lower electronegativity
4) fluorine, which has the lower electronegativity

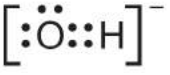
____ 10. An ionic compound is formed when there is a reaction between the elements

1) strontium and chlorine
2) hydrogen and chlorine
3) nitrogen and oxygen
4) sulfur and oxygen

____ 11. Which formula represents an ionic compound?

1) H_2
2) CH_4
3) CH_3OH
4) NH_4Cl

____ 12. Which Lewis electron-dot diagram correctly represents a hydroxide ion?

1) 
3) 
2) 
4) 

13. Which type of bond results when one or more valence electrons are transferred from one atom to another?

1) a hydrogen bond
2) an ionic bond
3) a nonpolar covalent bond
4) a polar covalent bond

14. Based on bond type, which compound has the highest melting point?

1) CH_3OH
2) C_6H_{14}
3) $CaCl_2$
4) CCl_4

15. Which substance is an electrolyte?

1) CH_3OH
2) $C_6H_{12}O_6$
3) H_2O
4) KOH

16. A solid substance was tested in the laboratory. The test results are listed below.
- dissolves in water
- is an electrolyte
- melts at a high temperature

Based on these results, the solid substance could be

1) Cu
2) $CuBr_2$
3) C
4) $C_6H_{12}O_6$

17. Which compound has both ionic and covalent bonding?

1) $CaCO_3$
2) CH_2C1_2
3) CH_3OH
4) $C_6H_{12}O_6$

18. Which element is composed of molecules that each contain a multiple covalent bond?

1) chlorine
2) fluorine
3) hydrogen
4) nitrogen

19. As a bond between a hydrogen atom and a sulfur atom is formed, electrons are

1) shared to form an ionic bond
2) shared to form a covalent bond
3) transferred to form an ionic bond
4) transferred to form a covalent bond

20. What is the total number of electrons shared in the bonds between the two carbon atoms in a the molecule shown below?

1) 6 2) 2 3) 3 4) 8

21. Which formula represents a molecular compound?

1) Kr
2) LiOH
3) N_2O_4
4) NaI

22. In which material are the particles arranged in a regular geometric pattern?

1) $CO_2(g)$
2) NaCl(aq)
3) $H_2O(\ell)$
4) $C_{12}H_{22}O_{11}(s)$

23. What is the maximum number of covalent bonds that a carbon atom can form?

1) 1 2) 2 3) 3 4) 4

24. Which type of bond is found between atoms of solid cobalt?

1) nonpolar covalent
2) polar covalent
3) metallic
4) ionic

25. A solid substance is an excellent conductor of electricity. The chemical bonds in this substance are most likely

1) ionic, because the valence electrons are shared between atoms
2) ionic, because the valence electrons are mobile
3) metallic, because the valence electrons are stationary
4) metallic, because the valence electrons are mobile

26. Which substance contains metallic bonds?

1) $Hg(\ell)$
2) $H_2O(\ell)$
3) NaCl(s)
4) $C_6H_{12}O_6(s)$

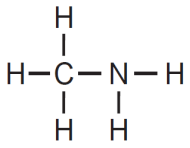
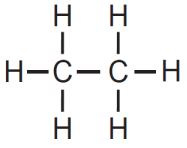
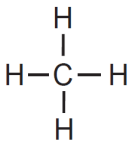
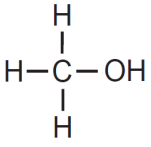
___ 27. A chemist performs the same tests on two homogeneous white crystalline solids, *A* and *B*. The results are shown in the table below.

	Solid A	Solid B
Melting Point	High, 801°C	Low, decomposes at 186°C
Solubility in H_2O (grams per 100.0 g H_2O at 0°C)	35.7	3.2
Electrical Conductivity (in aqueous solution)	Good conductor	Nonconductor

The results of these tests suggest that

1) both solids contain only ionic bonds
2) both solids contain only covalent bonds
3) solid *A* contains only covalent bonds and solid *B* contains only ionic bonds
4) solid *A* contains only ionic bonds and solid *B* contains only covalent bonds

___ 28. Which formula represents a molecule having a nonpolar covalent bond?

1) 
3) 
2) 
4) 

___ 29. The chemical bond between which two atoms is most polar?

1) C–N
2) H–H
3) S–Cl
4) Si–O

___ 30. Which compound has hydrogen bonding between its molecules?

1) CH_4
2) CaH_2
3) KH
4) NH_3

___ 31. Which formula represents a nonpolar molecule containing polar covalent bonds?

1) H_2O
2) CCl_4
3) NH_3
4) H_2

___ 32. Which formula represents a polar molecule?

1) H_2
2) H_2O
3) CO_2
4) CCL_4

___ 33. Which formula represents a nonpolar molecule?

1) HCl
2) H_2O
3) NH_3
4) CH_4

___ 34. At STP, fluorine is a gas and bromine is a liquid because, compared to fluorine, bromine has

1) stronger covalent bonds
2) stronger intermolecular forces
3) weaker covalent bonds
4) weaker intermolecular forces

___ 35. The four single bonds of a carbon atom in CH_4 are directed toward the corners of a

1) square
2) tetrahedron
3) rectangle
4) parallelogram

Base your answers to questions **36** and **37** on the information below.

Physical Properties of CF_4 and NH_3 at Standard Pressure

Compound	Melting Point (°C)	Boiling Point (°C)	Solubility in Water at 20.0°C
CF_4	−183.6	−127.8	insoluble
NH_3	−77.7	−33.3	soluble

____ 36. In the space *in your answer booklet*, draw a Lewis electron-dot diagram for CF_4.

____ 37. State evidence that indicates NH_3 has stronger intermolecular forces than CF_4.

Base your answers to questions **38** and **39** on the information below.

In 1864, the Solvay process was developed to make soda ash. One step in the process is represented by the balanced equation below.

$$NaCl + NH_3 + CO_2 + H2O \rightarrow NaHCO_3 + NH_4Cl$$

____ 38. In the space draw a Lewis electron-dot diagram for the reactant containing nitrogen in the equation.

____ 39. Explain, in terms of electronegativity difference, why the bond between hydrogen and oxygen in a water molecule is more polar than the bond between hydrogen and nitrogen in an ammonia molecule.

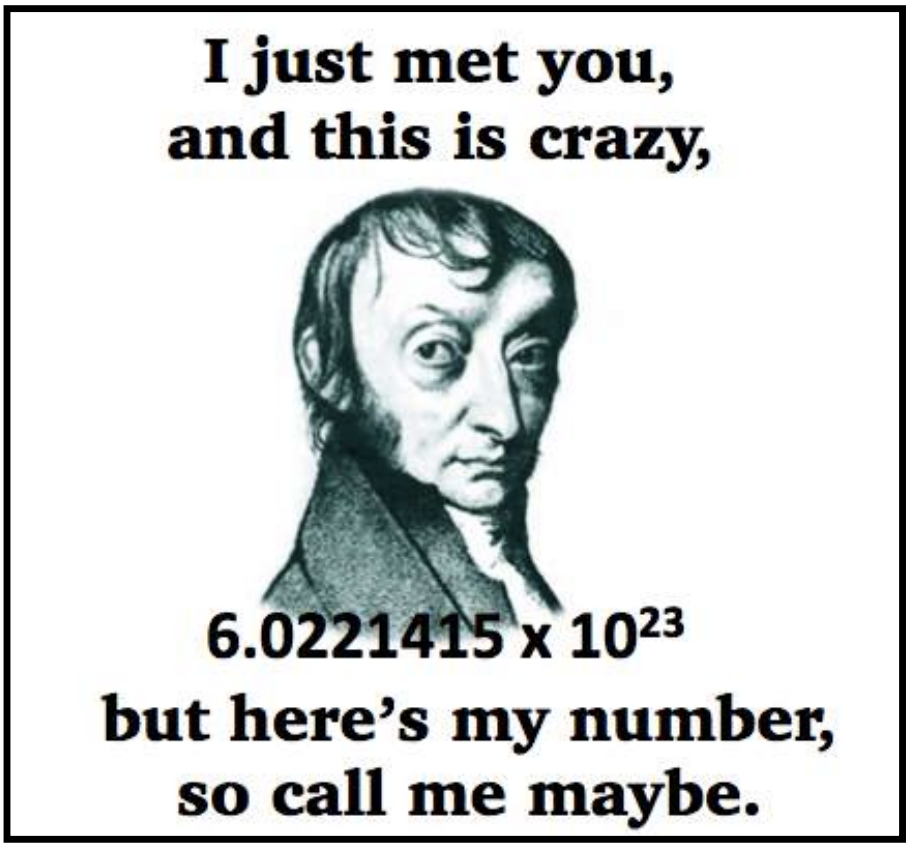
40. Draw a Lewis electron-dot diagram for a molecule of phosphorus trichloride, PCl_3

Name:

Regents Chemistry

Practice Packet

Chapter 10: Chemical Calculations



Chapter 10 : Chemical Calculations Vocabulary

1. **Mole** – a quantity of 6.02×10^{23} units of a substance; the amount of a substance equal to the sum of the atomic masses in grams; Avogadro's number
2. **Formula mass (FM)** – the sum of the atomic masses of a substance in a.m.u.
3. **Gram formula mass (GFM)** – the sum of the atomic masses of a substance in grams
4. **Coefficient** – the integer that appears in front of an element, molecule, or compound indicating the number of moles present
5. **Subscript** – the integer to the lower right of an element which indicates the number of atoms present in the compound
6. **Species** – the individual products and reactants in a chemical reaction
7. **Law of conservation of mass** – in any chemical reaction, mass can neither be created nor destroyed; the mass of the reactants must be equal to the mass of the products
8. **Law of conservation of energy** – in any chemical reaction, energy can neither be created nor destroyed; the energy of the reactants must be equal to the energy of the products
9. **Balanced equation** – a chemical equation in which the number of moles of each element on the reactants side is equal to the number of moles of each element on the products side
10. **Synthesis reaction** – a chemical reaction in which two or more substances combine to form a compound
    Ex: $A + B \rightarrow AB$
11. **Decomposition reaction** – a chemical reaction in which a compound is broken down into simpler substance
    Ex: $AB \rightarrow A + B$
12. **Single-replacement reaction -** a chemical reaction in which a metal replaces a metal OR a nonmetal replaces a nonmetal within a compound
    Ex: $A + BC \rightarrow AC + B$
13. **Double-replacement reaction -** a chemical reaction in which a metal replaces a metal AND a nonmetal replaces a nonmetal within two compounds; two compounds "trade" elements
    Ex: $AB + XY \rightarrow AY + XB$
14. **Molecular formula** – formula for a compound which provides the number and identity of the atoms of each element present
    Ex: $C_6H_{12}O_6$
15. **Empirical formula** – formula for a compound which provides the simplest ratio of the elements present
    Ex: The empirical formula for the molecule $C_6H_{12}O_6$ is CH_2O
16. **Percent mass** – **% composition by mass** = $\frac{\text{mass of part}}{\text{mass of whole}}$ **x 100**

Gram Formula Mass

Video Lesson 10.1

Objectives

- Describe how to calculate formula mass and gram formula mass
- Distinguish between the atomic mass of an element and its molar mass.

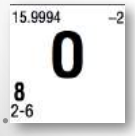
Stoichiometry

- The study of the quantitative relationships in chemical reactions.

IN OTHER WORDS … HOW MUCH?

Quick Review

- What are the units we use for the mass of atoms
  - ANS: Atomic mass units (amu) or (u)
- What is the mass of one atom of oxygen?
  - ANS: 15.9994 amu or u



Gram Formula Mass

- Formula mass
  - The mass of an atom, molecule or compound in **ATOMIC MASS UNITS (amu) or (u)**
  - EX. Formula mass of a hydrogen atom is 1.00794 amu
- Gram Formula Mass
  - The mass of one MOLE of an atom, molecule or compound in **GRAMS (g/mol)**
  - EX. GFM of hydrogen is 1.00794 g/mol

Example 1: Calculating Formula Mass

What is the formula mass of K_2CO_3?

K = 39.1 x 2 = 78.2
C = 12.0 x 1 = 12.0
O = 16.0 x 3 = + 48.0

138.2 amu

Example 2: Calculating Gram Formula Mass

What is the formula mass of K_2CO_3?

K = 39.1 x 2 = 78.2
C = 12.0 x 1 = 12.0
O = 16.0 x 3 = + 48.0
138.2 g/mol

Example 3: Calculating Gram Formula Mass

What is the gram formula mass of NH_3?

N = 1 x 14.0 = 14.0
H = 3 x 1.0 = + 3.0
17.0 g/mol

Example 4: Calculating Gram Formula Mass

- What is the gram formula mass of $Mg(OH)_2$

Mg = 1 x 24.3 = 24.3
O = 2 x 16.0 = 32.0
H = 2 x 1.0 = 2.0
+
58.3 g/mol

Percent Composition

Chemistry 200
Video Lesson 10.2

Objective:

How do we determine the difference between a Hydrate and Anhydrous?

How do we calculate the percent composition by mass for a hydrate or anhydrous using the percent composition formula?

Hydrate -

H_2O molecule(s) attached to an ionic crystal.

We calculate the gram formula mass as a separate unit.

ex: $CaSO_4 \bullet 2H_2O$

Anhydrous - ionic substances without H_2O

Percent Composition

Imagine a big bucket w/ 100 pieces of fruit in it. If the bucket has 25 apples, 30 oranges, 20 pears & 25 bananas, what percentage of the bucket consists of pears?

$$\frac{20 \text{ pears}}{100 \text{ pieces}} \times 100 = 20\%$$

Percent Composition can also be applied to the masses of elements in a compound using the following formula:

$$\% \text{ composition by mass} = \frac{\text{mass of part}}{\textbf{mass of whole}} \times 100$$

ex: What % of oxygen is in $KClO_3$?

GFM → K 1 atom 1 x 39.1 g/mol = 39.1
Cl 1 atom 1 x 35.5 g/mol = 35.5
O 3 atoms 3 x 16.0 g/mol = 48.0
122.6 g/mol

$$\% \text{ Oxygen} = \frac{48.0}{122.6} \times 100 = \textbf{39.2\%}$$

What is the % of water in $Na_2Cr_2O_7 \bullet 10H_2O$

Na = 23.0 g/mole x 2 atoms --> 46.0g/mol
Cr = 52.0 g/mole x 2 atoms --> 104.0g/mol
O = 16.0 g/mole x 7 atoms --> 112.0g/mol
H_2O = 18.0 g/mole x 10 atoms → 180.0g/mol
442.0 g/mol

$$\% \text{ of } H_2O = \frac{180.0 \text{g/mol}}{442.0 \text{ g/mol}} \times 100 = \textbf{40.7 \% } \mathbf{H_2O}$$

The Mole

Video Lesson 10.3

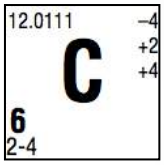
Objectives

- Define Avogadro's number as it relates to a mole of a substance.
- Describe how the mass of a mole of a compound is calculated.

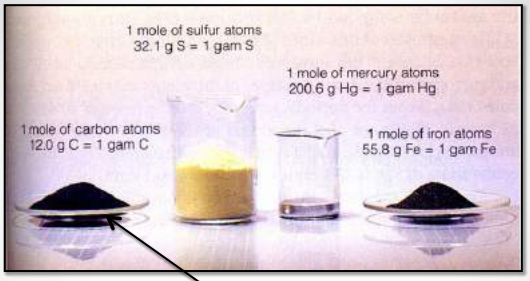
THE MOLE CONCEPT

- A mole is a chemistry unit of measure

12 bagels or 1 dozen



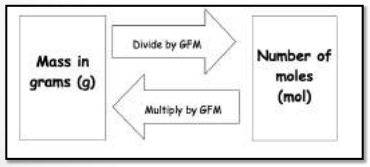
6.02 x 10^{23} atoms of C or 1 mole



Or 6.02 x 10^{23} atoms/particles of Carbon

USING MOLES

Mole Calculations	$\text{number of moles} = \frac{\text{given mass (g)}}{\text{gram-formula mass}}$



MOLE CALCULATION EXAMPLE #1

How many moles are in 39.0 grams of LiF?

Step 1:

Li = 1 x 7.0 = 7.0

F = 1 x 19.0 = 19.0

26.0 g/mol

Step 2: Solve for Moles

$\text{\# of moles} = \frac{\text{given mass (g)}}{\text{GFM (g/mol)}}$ Moles = $\frac{39.0\ g}{26.0\ g/mol}$

Moles = 1.50 mol

CONVERTING MOLES TO GRAMS EXAMPLE #2

What is the mass of 4.50 moles of KOH?

K = 1 x 39.1 = 39.1
O = 1 x 16.0 = 16.0
H = 1 x 1.0 = +1.0
56.1 g/ mol

$$\text{number of moles} = \frac{\text{given mass}}{\text{gram-formula mass}}$$

$$4.50 \text{ mol} = \frac{g}{56.1 \text{ g/mol}} \qquad g = 252.5g$$

Mole to Mole Ratios

Chemistry 200
Video Lesson 10.4

Objective:

How do we determine a new mole ratio in a reaction when the molar amount of a reactant or product is changed?

20 Cookies =

1cup-flour, 4 eggs, 2 tsp-salt, 3oz-butter, 12oz-choc chips

30 Cookies =

1.5cup-flour, 6 eggs, 3 tsp-salt, 4.5oz-butter, 18oz-choc chips

10 Cookies =

.5 cup-flour, 2 eggs, 1 tsp-salt, 1.5oz-butter, 6oz-choc chips

Mole to Mole ratios

In problems involving chemical reactions, the relative amounts of reactants & products are represented by the coefficients. Coefficients represent both the basic unit & mole ratios in balanced equations. The ratio is like a recipe for the reaction. The ratio tells how many moles of each species are necessary for the balanced reaction

We can use the **Mole Ratios** btwn elements & molecules in a chemical equation to make predictions about how they will react in nature by applying the following procedure:

****Ratios must be moles to moles,**
NOT grams to grams or grams to moles**

1. Put boxes around the 2 substances that are referred to in the question
2. Set up a proportion using the mole ratios of the 2 substances:
   a. Place the numbers in the chemical equation on the bottom part of the ratio
   b. Place the numbers in the word problem on the top of the ratio

Combustion of butane

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

3.55 moles of C_4H_{10} will yield how many moles of CO_2?

$$\frac{3.55 \text{ moles } C_4H_{10}}{2 \text{ moles } C_4H_{10}} = \frac{x \text{ moles } CO_2}{8 \text{ moles } CO_2}$$

Solve for x --> cross multiply

$2(x) = 8(3.55)$ $\qquad x = \frac{28.4}{2}$

$\boxed{x = 14.2 \text{ moles of } CO_2}$

$4Al + 3O_2 \rightarrow 2Al_2O_3$

What information do we know about the above equation?

Qualitative- Aluminum, Oxygen & Aluminum Oxide

Quantitative- 4 moles of Aluminum, 3 moles of Oxygen & 2 moles of Aluminum Oxide

Empirical & Molecular Formulas

Chemistry 200
Video Lesson 10.5

Objective:

How do we determine the molecular formula for a substance if given the empirical formula and gram formula mass?

Empirical & Molecular Formulas

Empirical Formula – The lowest whole number ratio of atoms in a compound. For example, C_6H_6 would have an empirical formula of CH. $C_6H_{12}O_6$ would have an empirical formula of CH_2O. An empirical formula can also be the same as the molecular formula, as in the case of H_2O.

Molecular Formula – A molecular formula represents the true formula of a substance. Unlike ionic formulas that need to be simplified, a molecular compound like dinitrogen tetroxide (N_2O_4) would not be simplified because that is the molecular formula.

Other examples would include N_2O, CO_2, C_3H_6, etc. If we were to represent these examples as empirical formulas, C_3H_6 could be simplified while the others are already empirical formulas.

Determining an empirical formula from a molecular formula

To convert a molecular formula into an empirical formula:

- Find the largest common multiple of all the subscripts in a formula
- Divide each subscript by that common multiple to create an empirical formula.

$$\frac{C_6H_9O_3}{3} = \mathbf{C_2H_3O}$$

$$\frac{C_6H_{14}}{2} = \mathbf{C_3H_7}$$

Determining a molecular formula from an empirical formula

- Converting an empirical formula into a molecular formula is a little trickier since we do not know the common multiple.
- Consider the empirical formula CH_2. This empirical formula can yield many molecular formulas including CH_2, C_2H_4, C_3H_6, C_4H_8 etc.

To convert an empirical formula into a molecular formula, use the following steps:

1. Determine the molecular mass of the empirical formula.
2. Divide the molecular mass **(GFM)** of the molecular formula by the molecular mass of the empirical formula. This represents the common multiple.
3. Multiply each subscript in the empirical formula by this common multiple to obtain the molecular formula.

Determine the molecular formula of a substance that has a molecular mass of 84.0 g/mole & an empirical formula of CH_2

Molecular mass of CH_2 (empirical formula)

C --> 1 atom x 12.0 g/mole = 12.0 g/mole

H --> 2 atoms x 1.0 g/mole = 2.0 g/mole

14.0 g/mole

$$\frac{\text{Molecular mass of molecular formula}}{\text{Molecular mass of empirical formula}}$$

$$\frac{84.0 \text{ g/mole}}{14.0 \text{ g/mole}} = 6$$

Molecular formula = $CH_2 \times 6 =$ **C_6H_{12}**

Sketch Notes

Video 10.1: Gram Formula Mass

Chemical Formula	Gram Formula Mass
1. $MgBr_2$	
2. KCl	
3. $FeCl_2$	
4. CrF_2	
5. Al_2S_3	
6. PbO	
7. TiI_4	
8. Mg_3P_2	
9. $SnCl_2$	
10. $HgCl_2$	
11. H_2SO_4	

Name:

Gram Formula Mass

1. The sum of the atomic masses of the atoms in one molecule of $C_3H_6Br_2$ is called the

   1) formula mass
   2) isotopic mass
   3) percent abundance
   4) percent composition

2. The gram-formula mass of NO_2 is defined as the mass of

   1) one mole of NO_2
   2) one molecule of NO_2
   3) two moles of NO
   4) two molecules of NO

3. What is the gram-formula mass of $Ca_3(PO_4)_2$?

   1) 248 g/mol
   2) 263 g/mol
   3) 279 g/mol
   4) 310. g/mol

4. The gram formula mass of NH_4Cl is

   1) 22.4 g/mole
   2) 28.0 g/mole
   3) 53.5 g/mole
   4) 95.5 g/mole

5. The gram-formula mass of $(NH_4)_2CO_3$ is

   1) 46.0 g
   2) 64.0 g
   3) 78.0 g
   4) 96.0 g

6. What is the gram-formula mass of $Fe(NO_3)_3$?

   1) 146 g/mol
   2) 194 g/mol
   3) 214 g/mol
   4) 242 g/mol

7. What is the gram-formula mass of $(NH_4)_3PO_4$?

   1) 112 g/mol
   2) 121 g/mol
   3) 149 g/mol
   4) 242 g/mol

8. A 1.0-mole sample of krypton gas has a mass of

   1) 19 g 2) 36 g 3) 39 g 4) 84 g

9. The molar mass of $Ba(OH)_2$ is

   1) 154.3 g
   2) 155.3 g
   3) 171.3 g
   4) 308.6 g

10. What is the gram formula mass of $Ca_3(PO_4)_2$?

    1) 196 g
    2) 214 g
    3) 245 g
    4) 310. g

Video 10.2: Percent Composition
Complete the following calculations. Show all work and units. Round GFM to nearest tenth.

1. What is the percent by mass of magnesium in $Mg_3(PO_4)_2$?

2. What percentage of Na_2SO_4 is oxygen by mass?

3. What is the percent of nitrogen by mass in ammonium nitrate?

4. What is the percent by mass of water in $Al(ClO_3)_3 \cdot 6H_2O$?

5. Calculate the % of hydrogen in nitroglycerine {$\boldsymbol{C_3H_5N_3O_9}$}.

6. What percentage of $CaSO_4 \cdot 2H_2O$ is water by mass?

7. Calculate the percent composition of oxygen in glycerol {$\boldsymbol{C_3H_5(OH)_3}$}.

8. Calculate the percent aluminum in aluminum oxide.

9. Calculate the percent of phosphorus in sodium phosphate?

10. Calculate the percent by mass of water in $CuNO_3 \bullet 5H_2O$?

Chapter 10: Chemical Calculations

Video 10.3

Moles and Molar Mass: Find the gram formula mass of the following: (Show all work)

1. CO_2
2. FeS
3. NaCl
4. $Al_2(CO_3)_3$
5. H_2SO_4
6. $Al_2(SO_3)_3$
7. Fe_2O_3
8. $Ca(OH)_2$
9. NH_3
10. H_2O_2
11. $NaHCO_3$
12. $C_6H_{12}O_6$
13. MgO
14. $SrSO_4 \cdot 3H_2O$

Moles: Find the number of moles in the following measurements: (Show your work)

Using the mole formula convert the following from grams to moles. An example is provided. Round to the nearest hundredth.

Example: Convert 15.0 g of $CaCl_2$ to moles.

Step 1: Calculate gram formula mass	Step 2: Plug into mole equation moles = grams/GFM
Ca = 1 x 40.08 = 40.08	moles = 15.0 g/110.98 g
Cl = 2 x 35.45 = 70.90	moles = 0.140 moles
110.98 grams/mol	

1. 900. grams $C_6H_{12}O_6$
2. 24.5 grams H_2SO_4
3. 192 grams SiO_2
4. 450. grams of $ZnCl_2$
5. 22 grams of CO_2
6. 20. grams of Fe_2O_3
7. 3.40 grams of H_2O_2
8. 840. grams of NaHCO

Now solve for the mass given the moles. (Show your work)

Using the mole formula convert the following from moles to grams. An example is provided. Round to the nearest hundredth.

Example: Convert 2.5 moles of $CaCl_2$ to grams

Step 1: Calculate gram formula mass

Ca = 1 x 40.08 = 40.08

Cl = 2 x 35.45 = 70.90

110.98 grams/mol

Step 2: Plug into mole equation

moles = grams/GFM

2.5 moles = g/110.98 g

grams = 277.5 grams

1. 2.00 moles of $C_6H_{12}O_6$
2. 5.00 moles of $SrSO_4 \cdot H_2O$
3. 0.250 moles of CH_4
4. 0.100 moles of NH_3
5. 12.0 moles of SiO_2
6. 0.330 moles of FeS
7. 1.50 moles of MgO
8. 0.500 moles of $ZnCl_2$

Moles

1. Which sample contains a mole of atoms?

   1) 23 g Na 2) 24 g C
   3) 42 g Kr 4) 78 g K

2. What is the total mass of 2.0 moles of $H_2(g)$?

   1) 1.0 g 2) 2.0 g 3) 3.0 g 4) 4.0 g

3. What is the mass in grams of 2.0 moles of NO_2?

   1) 92 2) 60. 3) 46 4) 30.

4. What is the total number of moles in 80.0 grams of C_2H_5Cl (gram-formula mass = 64.5 grams/mole)?

5. The number of moles of molecules in a 12.0-gram sample of Cl_2 is

   1) $\frac{12.0}{35.5}$ mole 2) $\frac{12.0}{71.0}$ mole
   3) 12.0 moles 4) 12.0×35.5 moles

6. What is the total mass of oxygen in 1.00 mole of $Al_2(CrO_4)_3$?

   1) 192 g 2) 112 g
   3) 64.0 g 4) 48.0 g

7. What is the total mass in grams of 0.75 mole of SO_2?

   1) 16 g 2) 24 g 3) 32 g 4) 48 g

8. The total number of moles represented by 20 grams of $CaCO_3$ is

   1) 1 2) 2 3) 0.1 4) 0.2

9. What is the mass in grams of 1.00 mole of O_2 gas?

   1) 11.2 2) 16.0 3) 22.4 4) 32.0

10. What is the total mass of iron in 1.0 mole of Fe_2O_3?

    1) 160 g 2) 112 g
    3) 72 g 4) 56 g

Video 10.4: Mole to Mole Ratios

Determine the molar amount using mole to mole ratios for the following reactions

1. Given the reaction: $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

   What is the maximum number of moles of H_2O that can be produced when 2.0 moles of NH_3 are completely reacted?

2. Given the reaction: $2KClO_3 \rightarrow 2KCl + 3O_2$

   What is the total number of moles of potassium chlorate needed to produce 6.0 moles of O_2?

3. Given the reaction: $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$

   How many moles of oxygen are required to react completely with 1.0 moles of C_2H_2?

4. Given the reaction: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

   How many moles of carbon dioxide are produced when 0.5 moles of Fe_2O_3 are completely reacted?

   Convert your answer to grams of CO_2

5. Given the reaction: $3MgCl_2 + 2Al \rightarrow 3Mg + 2AlCl_3$
   How many moles of aluminum chloride are produced when 8 moles of magnesium chloride are completely reacted with aluminum?

6. Consider the combustion of methane. How many moles of carbon dioxide are obtained when 20 moles of methane are completely burned.

7. Given the reaction: $2H_2 + O_2 \rightarrow 2H_2O$
   How many moles of H_2O can be produced if 20 moles of hydrogen completely react with oxygen?

Law of Conservation of Mass Worksheet

Since there is a conservation of mass in all balanced chemical reactions, these problems are also quite easy. Consider the following balanced chemical reaction **$CaCO_3 \rightarrow CaO + CO_2$**. In this reaction the mass of $CaCO_3$ **must** equal the combined mass of CaO and CO_2. If you know the mass of the reactants you automatically know the mass of the products and vice-versa.

Using the reaction above, determine the mass of CaO produced if 200 grams of $CaCO_3$ decomposed and produced 88 grams of CO_2.

$CaCO_3 \rightarrow CaO + CO_2$
200g = Xg + 88g → 200g – 88g = 112g CaO produced

Complete the following problems

1) Hydrogen and oxygen react chemically to form water. How much water would form if 14.8 grams of hydrogen reacted with 34.8 grams of oxygen? ($H_2 + O_2 \rightarrow H_2O$)

2) When ammonium nitrate (NH_4NO_3) explodes, the products are nitrogen, oxygen, and water. When 40 grams of ammonium nitrate explode, 14 grams of nitrogen and 8 grams of oxygen form. How many grams of water form? ($NH_4NO_3 \rightarrow N_2 + O_2 + H_2O$**)**

3) 40 g of calcium reacts with 71 g of chlorine to produce ____ g of calcium chloride.

4) ____ g of potassium reacts with 16 g of oxygen to produce 94 g of potassium oxide.

5) 14 g of lithium reaction with ____ g sulfur to produce 46 g of lithium sulfide.

6) 24 g of magnesium reacts with 38 g of fluorine to produce ____ g magnesium fluoride.

7) 65.5 g copper reacts with ____ g oxygen to produce 81 g copper (I) oxide.

8) 88 g of strontium reacts with 160 g bromine to produce ____ g strontium bromide.

9) 46 g of sodium reacts with ____ g chlorine to produce 117 g sodium chloride.

10) ___ g iron reacts with 71 g chlorine to produce 129 g of iron (II) chloride.

11) 137 g of barium reacts with _____ g iodine to produce 391 g barium iodide.

12)____ g hydrogen reacts with 32 g of oxygen to produce 34 g of hydrogen peroxide.

13)Why do we balance chemical reactions?

Practice:

1. Given the balanced equation representing a reaction:

$$CaO_{(s)} + CO_{2(g)} \rightarrow CaCO_{3\,(s)} + \text{heat}$$

What is the total mass of CaO that reacts with 75 grams of $CO_{2\,(g)}$ to produce 200. Grams of $CaCO_{3(s)}$?

2. Given the balanced equation representing a reaction:

$$2H_2 + O_2 \rightarrow 2H_2O$$

What is the total mass of water formed when 4 grams of hydrogen reacts completely with 32 grams of oxygen?

Video 10.5: Empirical and Molecular Formulas

Determining an empirical formula from a molecular formula

To convert a molecular formula into an empirical formula, we need to find the largest common multiple of all the subscripts in a formula. Then simply divide each subscript by that common multiple to create an empirical formula.

Molecular Formula	Largest Common Multiple	Empirical Formula
$C_6H_9O_3$	3	C_2H_3O
C_6H_{14}		
$C_2H_4O_2$		
H_2O_2		
$C_{12}H_{22}$		
$C_{12}H_{24}$		
N_2O_3		
C_3H_6		

Determining a molecular formula from an empirical formula

1. Determine the molecular mass of the empirical formula.
2. Divide the molecular mass of the molecular formula by the molecular mass of the empirical formula. This represents the common multiple.
3. Multiply each subscript in the empirical formula by this common multiple to obtain the molecular formula.

Practice

1. A substance w/ a molecular mass of 172.0 g/mole is determined to have an empirical formula of C_3H_7. Determine the molecular formula of this substance.

2. Determine the molecular formula of a substance w/ an empirical formula of $C_2H_3S_2$ and a molecular mass of 456.0 g/mole.

3. Determine the molecular formula of a substance w/ an empirical formula of C_2H_4N and a molecular mass of 294.0 g/mole.

4. Determine the molecular formula of a compound with an empirical formula of NH_2 and a molecular mass of 32.0 g/mole.

5. Organic gas has the empirical formula CH_3 and a molecular mass of 165.0 g/mole. Determine the molecular formula.

6. Vitamin C has an empirical formula of $C_3H_4O_3$ and a molecular mass of 264.0 g/mole. Determine the molecular formula.

7. Ibuprofen, a common headache remedy, has an empirical formula of C_7H_9O and a molar mass of 545.0 g/mole. Determine the molecular formula.

8. Oxalic acid has the empirical formula HCO_2 and a molar mass of 90.0 g/mole. Determine the molecular formula.

Name:

Chemical Calculations Review

1. Which type of formula represents the simplest whole-number ratio of atoms of the elements in a compound?

   1) molecular formula
   2) condensed formula
   3) empirical formula
   4) structural formula

2. Given two formulas representing the same compound:

   Formula A CH_3
   Formula B C_2H_6
   Which statement describes these formulas?

   1) Formulas *A* and *B are both empirical.*
   2) Formulas *A* and *B are both molecular.*
   3) Formula *A is empirical, and formula B is molecular.*
   4) Formula *A is molecular, and formula B is empirical.*

3. What is the empirical formula of a compound that has a carbon-to-hydrogen ratio of 2 to 6?

   1) CH_3 2) C_2H_6 3) C_3H 4) C_6H_2

4. Which pair consists of a molecular formula and its corresponding empirical formula?

   1) C_2 H_2 and CH_3 CH_3
   2) C_6 H_6 and C_2H_2
   3) P_4O_{10} and P_2O_5
   4) SO_2 and SO_3

5. Which formula is both a molecular and an empirical formula?

   1) $C_6H_{12}O_6$ 3) C_3H_8O
   2) $C_2H_4O_2$ 4) C_4H_8

6. What is the gram-formula mass of $(NH_4)_3PO_4$?

   1) 112 g/mol 3) 149 g/mol
   2) 121 g/mol 4) 242 g/mol

7. A 1.0-mole sample of krypton gas has a mass of

   1) 19 g 2) 36 g 3) 39 g 4) 84 g

8. The gram-formula mass of H_2O is defined as the mass of

   1) one mole of H_2O
   2) one molecule of H_2O
   3) two moles of H_2O
   4) two molecules of H_2O

9. What is the gram formula mass of Li_2SO_4?

   1) 54 g 2) 55 g 3) 110 g 4) 206 g

10. What is the total number of oxygen atoms in the formula $MgSO_4 \cdot 7\ H_2O$? [The • represents seven units of H_2O attached to one unit of $MgSO_4$.]

    1) 11 2) 7 3) 5 4) 4

11. The gram-formula mass of a compound is 48 grams. The mass of 1.0 mole of this compound is

    1) 1.0 g 2) 4.8 g 3) 48 g 4) 480 g

12. Which sample contains a mole of atoms?

    1) 23 g Na 3) 42 g Kr
    2) 24 g C 4) 78 g K

13. A compound has a molar mass of 90. grams per mole and the empirical formula CH_2O. What is the molecular formula of this compound?

    1) CH_2O 3) $C_3H_6O_3$
    2) $C_2H_4O_2$ 4) $C_4H_8O_4$

14. A substance has an empirical formula of CH_2 and a molar mass of 56 grams per mole. The molecular formula for this compound is

    1) CH_2 2) C_4H_6 3) C_4H_8 4) C_8H_4

15. Which quantity can be calculated for a solid compound, given only the formula of the compound and the Periodic Table of the Elements?

    1) the density of the compound
    2) the heat of fusion of the compound
    3) the melting point of each element in the compound
    4) the percent composition by mass of each element in the compound

16. Which compound has the highest precent composition by mass of strontium?

    1) $SrCl_2$ 2) SrI_2 3) SrO 4) SrS

17. What is the percent composition by mass of sulfur in the compound $MgSO_4$ (gram-formula mass = 120. grams per mole)?

    1) 20% 2) 27% 3) 46% 4) 53%

18. What is the percent composition by mass of nitrogen in NH_4NO_3 (gram-formula mass = 80.0 grams/mole)?

    1) 17.5% 3) 52.5%
    2) 35.0% 4) 60.0%

19. During all chemical reactions, mass, energy, and charge are

    1) absorbed 3) formed
    2) conserved 4) released

20. The coefficients in a balanced chemical equation represent

    1) the mass ratios of the substances in the reaction
    2) the mole ratios of the substances in the reaction
    3) the total number of electrons in the reaction
    4) the total number of elements in the reaction

21. Given the balanced equation representing a reaction:

    $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

    What is the *minimum* number of moles of O_2 that are needed to completely react with 16 moles of NH_3?

    1) 16 mol
    2) 20. mol
    3) 64 mol
    4) 80. mol

22. Given the balanced equation representing a reaction:

    $$Al_2(SO_4)_3 + 6NaOH \rightarrow 2Al(OH)_3 + 3Na_2SO_4$$

    The mole ratio of $NaOH$ to $Al(OH)_3$ is

    1) 1:1 2) 1:3 3) 3:1 4) 3:7

23. Given the balanced equation representing a reaction:

    $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

    What is the total number of moles of $O_2(g)$ required for the complete combustion of 1.5 moles of $C_3H_8(g)$?

    1) .30 mol
    2) 1.5 mol
    3) 4.5 mol
    4) 7.5 mol

24. Given the balanced equation representing a reaction:

    $CaO(s) + CO_2(g) \rightarrow CaCO_3(s) + heat$

    What is the total mass of CaO(s) that reacts completely with 88 grams of $CO_2(g)$ to produce 200. grams of $CaCO_3(s)$?

    1) 56 g 2) 88 g 3) 112 g 4) 288 g

25. Given the balanced equation representing a reaction:

    $2H_2 + O_2 \rightarrow 2H_2O$

    What is the total mass of water formed when 8 grams of hydrogen reacts completely with 64 grams of oxygen?

    1) 18 g 2) 36 g 3) 56 g 4) 72 g

26. Given the reaction:

    $2\ H_2 + O_2 \rightarrow\ 2\ H_2O$

    The total number of grams of O_2 needed to produce 54 grams of water is

    1) 36 2) 48 3) 61 4) 75

27. Base your answer to the following question on the information below and on your knowledge of chemistry.

Many breads are made by adding yeast to dough, causing the dough to rise. Yeast is a type of microorganism that produces the catalyst zymase, which converts glucose, $C_6H_{12}O_6$, to ethanol and carbon dioxide gas. The balanced equation for this reaction is shown below.

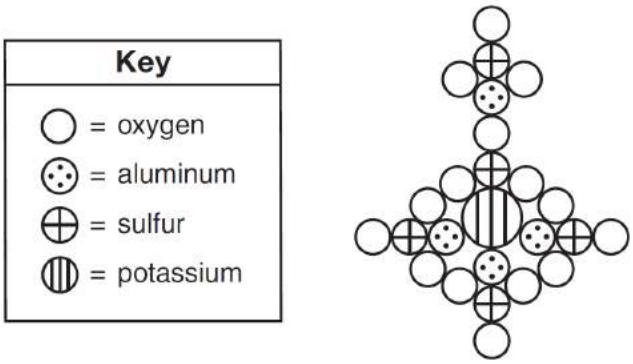
$$C_6H_{12}O_6(aq) \xrightarrow{zymase} 2C_2H_5OH(aq) + 2CO_2(g)$$

Determine the total mass of ethanol produced when 270. grams of glucose reacts completely to form ethanol and 132 grams of carbon dioxide.

28. Base your answer to the following question on the information below.

John Dalton, an early scientist, sketched the structure of compounds using his own symbols for the elements known at the time. Dalton's symbols for four elements and his drawing of potassium aluminum sulfate are represented by the diagram below.

Dalton's Drawing for Potassium Aluminum Sulfate



Today, it is known that the chemical formula for potassium aluminum sulfate is $KAl(SO_4)_2 \bullet 12H_2O$. It is a hydrated compound because water molecules are included within its crystal structure. There are 12 moles of H_2O for every 1 mole of $KAl(SO_4)_2$. The compound contains two different positive ions. The gram-formula mass of $KAl(SO_4)_2 \bullet 12H_2O$ is 474 grams per mole.

Show a numerical setup for calculating the percent composition by mass of water in $KAl(SO_4)_2 \bullet 12H_2O$.

29. Base your answer to the following question on the information below.

Vitamin C, also known as ascorbic acid, is water soluble and cannot be produced by the human body. Each day, a person's diet should include a source of vitamin C, such as orange juice. Ascorbic acid has a molecular formula of $C_6H_8O_6$ and a gram-formula mass of 176 grams per mole.

Show a numerical setup for calculating the percent composition by mass of oxygen in ascorbic acid.

Base your answers to questions **30** and **31** on the information below.

Hydrogen peroxide, H_2O_2, is a water-soluble compound. The concentration of an aqueous hydrogen peroxide solution that is 3% by mass H_2O_2 is used as an antiseptic. When the solution is poured on a small cut in the skin, H_2O_2 reacts according to the balanced equation below.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

30. Determine the gram-formula mass of H_2O_2.

31. Calculate the total mass of H_2O_2 in 20.0 grams of an aqueous H_2O_2 solution that is used as an antiseptic. Your response must include *both* a numerical setup and the calculated result.

32. Base your answer to the following question on the information below.

The Solvay process is a multistep industrial process used to produce washing soda, $Na_2CO_3(s)$. In the last step of the Solvay process, $NaHCO_3(s)$ is heated to 300°C, producing washing soda, water, and carbon dioxide. This reaction is represented by the balanced equation below.

$$2NaHCO_3(s) + \text{heat} \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Determine the total mass of washing soda produced if 3360. kilograms of $NaHCO_3$ reacts completely to produce 360. kilograms of H_2O and 880. kilograms of CO_2.

Free Response Review

1. Consider the reaction below that represents the combustion of benzene, C_2H_6

$$2\ C_2H_6 + 15\ O_2 \rightarrow 12\ CO_2 + 6H_2O$$

   a. How many moles of benzene are in 22.3 gram sample of benzene?

   b. What is the mass of a 1.79 mole sample of O_2?

   c. A 4.4 mole sample of O_2 will react to produce how many moles of CO_2?

   d. How many moles of benzene are needed to react completely with 3.47 moles of O_2?

2. Putrescine, is a compound produced in decaying animals and creates the characteristic odor associated with rotten meat. The empirical formula for putrescine is C_2H_6N.
   a. If putrescine has a molecular mass of 88.0 g/mol, determine the molecular formula of putrescine.

   b. How many moles of putrescine are in a 37.5 gram sample of putrescine?

   c. What is the percent of nitrogen in putrescine?

3. Answer the following questions about the following hydrated salt:

   a. Calculate the gram formula mass of $CaSO_3 \bullet 4H_2O$

   b. Determine the moles of $CaSO_3 \bullet 4H_2O$ in a 15.6 gram sample.

   c. Calculate the percent of water in $CaSO_3 \bullet 4H_2O$

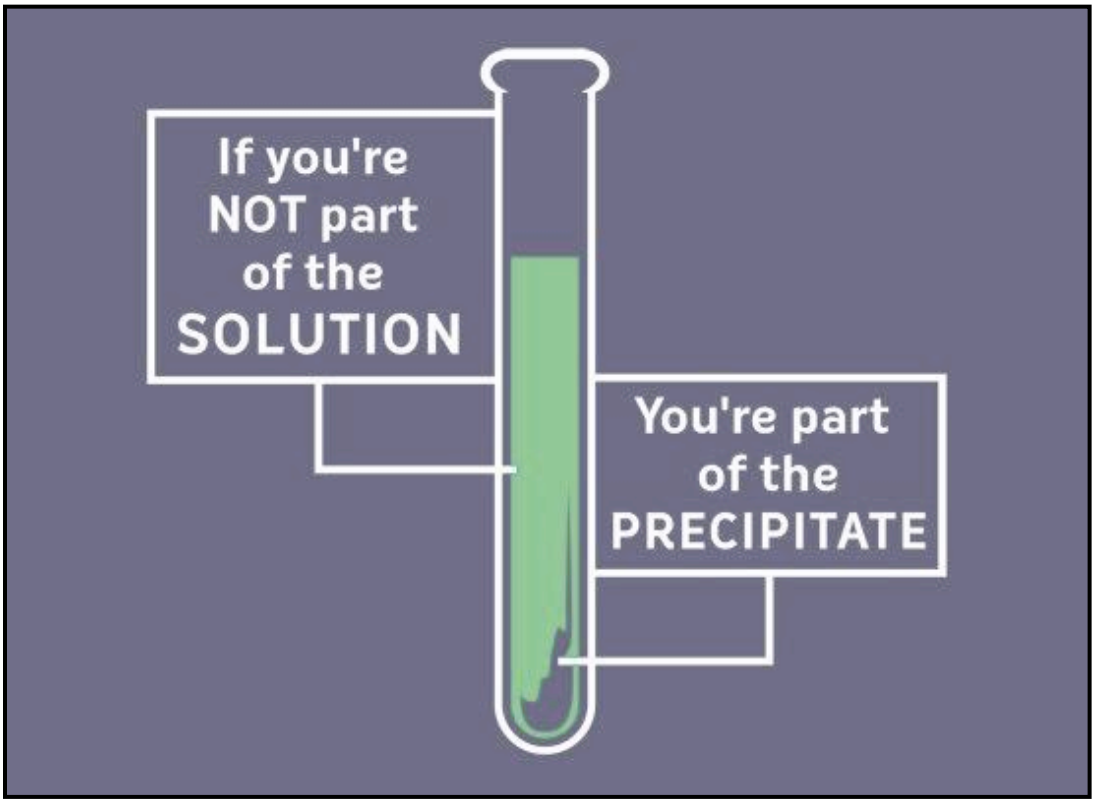
4. In an experiment, 4.56 grams of sodium completely reacts with sulfur, producing 7.89 grams of sodium sulfide. Determine the total mass of sulfur consumed.

Name:

Regents Chemistry

Practice Packet

Chapter 11: Solutions



Regents Chemistry **Name** _____

VOCAB – Solutions **Period** ____

Chapter 7

1. **Alloy** – a homogenous mixture/solution containing at least one metal. Ex: brass, steel, bronze
2. **Aqueous** – a homogenous mixture/solution in which a solute is dissolved in water.
3. **Boiling Point** – the temperature at which a liquid undergoes a phase change from liquid to gas; the temperature at which the vapor pressure of a liquid is equal to the atmospheric pressure.
4. **Boiling Point Elevation** – the boiling point of a solution is higher than the boiling point of the pure solvent (colligative property)
5. **Colloid** – a heterogeneous mixture composed of tiny particles suspended in another material. The particles are larger than the particles in a solution but smaller than particles in a suspension. Ex: milk, blood
6. **Concentrated** – Having a relatively large amount of substance present in a unit amount of mixture. For example, a 12 M HCl solution is more concentrated than an 0.001 M HCl solution.
7. **Concentration** – A measure of the amount of solute present in a unit amount of mixture. (Ex: ppm = parts per million, molarity = moles solute/liter solution); the process of increasing the amount of substance in a given amount of mixture.
8. **Dilute** – having a relatively low concentration of solute in a mixture.
9. **Freezing Point Depression** – the freezing point/melting point of a solution is lower than the freezing point/melting point of the pure solvent (colligative property)
10. **Heterogeneous** – A sample of matter consisting of more than one pure substance and more than one phase
11. **Homogeneous** – A sample of matter consisting of more than one pure substance with properties that do not vary within the sample
12. **Insoluble** – Refers to a substance that does not dissolve in a solvent to any significant degree
13. **Miscible** – Two liquids are considered "miscible" or mixable if shaking them together results in a single liquid phase with no visible separation
14. **Mixture** – two or more pure substance PHYSICALLY combined; a combination of two or more pure substances that can be separated by physical means
15. **Molarity** – a measure of concentration; M = moles of solute/liters of solution
16. **Parts Per Million** – a measure of concentration; ppm = parts of solute/million parts of solution
17. **Percent Composition (by mass or volume)** – % comp = (part/whole) x 100
18. **Precipitate** – An insoluble substance that has been formed from a chemical reaction between substances dissolved in a solution
19. **Saturated** – a solution that has reached equilibrium; a solution which can not dissolve any more solute
20. **Solubility** – a measure of the concentration of a substance in a saturated solution; a measure of how much of a substance can dissolve in a given amount of solvent

21. **Soluble** – capable of being dissolved in a solvent
22. **Solution** – a homogenous mixture
23. **Solute** – A substance dissolved in a solvent to make a solution
24. **Solvent** – The most abundant component in a solution
25. **Supersaturated** – a solution in which the concentration of solute is higher than the solubility; more solute is dissolved than should be under a given set of conditions
26. **Suspension** – A heterogenous mixture in which relatively large particles are suspended in a liquid
27. **Tyndall Effect** – Light passing through a colloid is scattered by suspended particles (the light beam becomes clearly visible)
28. **Unsaturated** – A solution with a concentration lower than its equilibrium solubility; a solution in which more solute can be dissolved

What are Solutions?

Chemistry 200
Video Lesson 11.1

Objective:

How do we recognize the parts of a solution, its characteristics and the difference between dissolving and dissociation?

Solution Chemistry

Solution - a homogeneous mixture of substances

- contains atoms, ions or molecules of one substance spread evenly throughout a 2nd substance

Homogeneous mixture

- a mixture whose composition is consistent throughout

Parts of a Solution:

1. Solute - substance being dissolved, present in a smaller amount
2. Solvent - substance that dissolves the solute, present in a greater amount
   (H_2O is a common solvent, Alcohol in H_2O)

(aq) = an H_2O solution

Tyndall Effect

- light scatters in a colloid but will not scatter in a true solution. This effect is used to determine whether a mixture is a true solution or a colloid.

Colloids are common materials with one material evenly distributed within another material on a very tiny scale.

Ex: milk, fog, jelly, styrofoam and whipped cream

Solution Characteristics

1. Solutions are homogeneous mixtures
2. Solutions are clear & will not disperse light
3. Solutions can have color
4. Solutions will not settle after standing
5. Solutions will pass through a filter

Dissolving vs. Dissociation

Dissolving - a general term used when a substance is placed into a solvent (H_2O) and broken down into smaller pieces; the substance may seem to not exist anymore.

Dissociation ***(ionization)*** - an ionic molecule separates into 2 or more ions; usually by dissolving an ionic compound in H_2O → no new substances formed

$$NaCl_{(s)} \xrightarrow{H_2O} Na^{+1}_{(aq)} + Cl^{-1}_{(aq)}$$

$$Al_2(SO_4)_{3(s)} \xrightarrow{H_2O} 2\ Al^{+3}_{(aq)} + 3\ SO_4^{-2}_{(aq)}$$

What is an electrolyte?

- Compounds that conduct electricity in solution due to mobile charged ions. They must dissociate!!
- This type of solution will cause a bulb to glow brightly.
- If the compound does not conduct in solution, it is called a ***nonelectrolyte***.

Solutions

Solute	***Solvent***	***Solution***	***Examples***
Gas	***Gas***	***Gas***	Air
Solid	***Liquid***	***Liquid***	Salt H_2O, Sugar H_2O
Liquid	***Liquid***	***Liquid***	Alcohol in H_2O
Gas	***Liquid***	***Liquid***	Carbonated H_2O Aquarium--> $O_{2(g)}$
Solid	***Solid***	***Solid***	"Gold & Silver" Jewelry Steel, each is an Alloy

Factors that Affect Solubility & Table G

Video Lesson 11.2

Objectives

- Identify the factors that determine the rate at which a solute dissolves.
- Identify the factors that determine the mass of solute that will dissolve in a given mass of solvent.

Solubility

- How much the substance can be dissolved in a given quantity of the solvent

Factors that affect solubility

Temperature: Higher the temp the greater the solubility
Gases require a lower temp

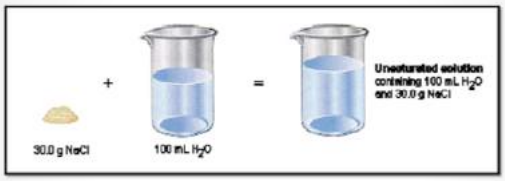
Pressure: Higher the pressure the greater the solubility (gases only!)

Nature of solute & solvent: polar solute will dissolve in polar solvent

Different Types of Solutions

- Unsaturated
  - Can dissolve more solute
- Saturated
  - contain the maximum solute that will dissolve at a specified temperature
  - At dynamic equilibrium.
    - Rate of dissolving is equal to rate of precipitation
- Supersaturated
  - contains more than the saturated amount of solute.
    - Dissolve solute at elevated temps

Unsaturated Solutions



Saturated Solution

Supersaturated Solution

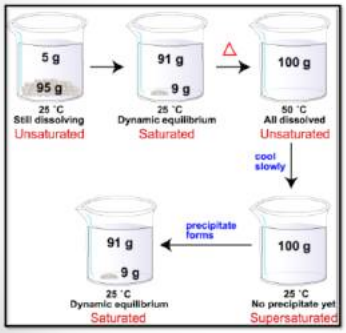
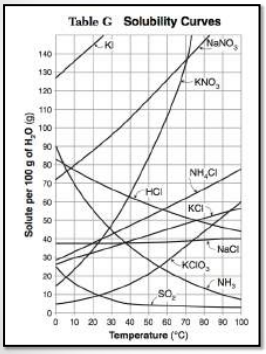
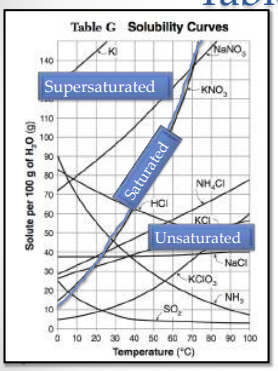


Table G

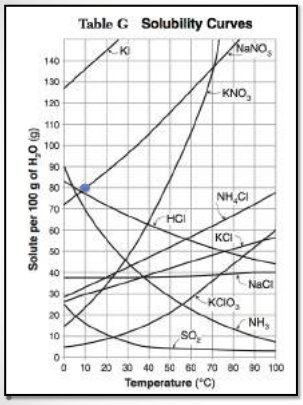


- Shows how temperature affects the solubility of 10 different substances.
- How much of the substance can dissolve in 100 grams of water.

Table G

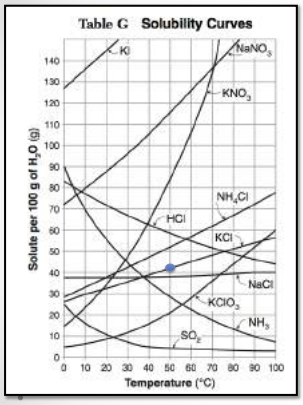


On the line = saturated
Below line = unsaturated
Above line = supersaturated



Which substance forms an unsaturated solution when 80 grams of the substance is added in 100 g of H_2O at 10°C?

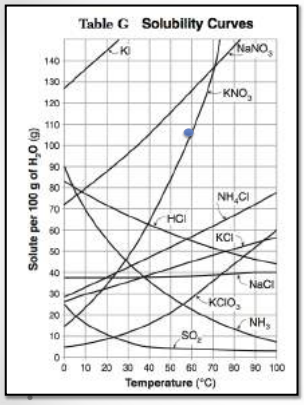
1. KI
2. KNO_3
3. $NaNO_3$
4. NaCl



What is the maximum number of grams of KCl that will dissolve in a 200 grams of water at 50°C to produce a saturated solution?

$$\frac{41\text{ g}}{100\text{ g of water}} = \frac{x}{200\text{ g of water}}$$

82 grams



A supersaturated solution of KNO_3 is created by adding 120 g of KNO_3 to 100 g of water at 60°C. If the solution is disrupted, how many grams of solid KNO_3 will "fall out" (will not dissolve) of the solution?

120 g – 109 g = 11 g

Table F

Video Lesson 11.3

Objectives

- Identify ionic solutes that will dissolve in water using Table F.
- Write double replacement reactions involving solutions.

Solubility Factors

- "Like Dissolves Like"

	Nonpolar Solvent	Polar Solvent
Nonpolar Solute	SOLUBLE	INSOLUBLE
Polar Solute	INSOLUBLE	SOLUBLE
***Ionic Solute (charges)**	INSOLUBLE	SOLUBLE

Solubility of Ionic Solids in Water

- Dissociation of Ions (ions separate)
- $NaCl\ (aq) \rightarrow Na^+ + Cl^-$

Ionic Solids & Solubility in Water

Table F
Solubility Guidelines for Aqueous Solutions

Ions That Form Soluble Compounds	Exceptions	Ions That Form Insoluble Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)		carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
ammonium (NH_4^+)		chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
nitrate (NO_3^-)		phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
acetate ($C_2H_3O_2^-$ or CH_3COO^-)		sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydrogen carbonate (HCO_3^-)		hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)
chlorate (ClO_3^-)			
perchlorate (ClO_4^-)			
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}		
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}		

Insoluble

soluble

Soluble

Is K_2CO_3 soluble or insoluble in water?

Reactions & Table F

- Will a precipitate form when silver nitrate nitrate is mixed with potassium chloride? Identify the precipitate.
- **Step 1**: Write the chemical formulas and *double replacement* equation. (Be sure to balance charge)

$$\overset{+1}{Ag}\overset{-1}{NO_3}_{(aq)} + \overset{+1}{K}\overset{-1}{Cl}_{(aq)} \rightarrow \overset{+1}{Ag}\overset{-1}{Cl} + \overset{+1}{K}\overset{-1}{NO_3}$$

Reactions & Table F

Step 2: Check the solubility's of the products, (on table F) if insoluble you found the precipitate!

$$AgNO_{3\,(aq)} + KCl_{\,(aq)} \rightarrow AgCl_{(s)} + KNO_{3\,(aq)}$$

precipitate

Silver nitrate and sodium chromate solutions are mixed together. Will a precipitate form? If so, what is the name of the precipitate?

Step 1: Write the equation for the double replacement reaction.

$$2\,\overset{+1}{Ag}\overset{-1}{NO_3} + \overset{+1}{Na_2}\overset{-2}{CrO_4} \rightarrow \overset{+1}{Ag_2}\overset{-2}{CrO_4} + 2\,\overset{+1}{Na}\overset{-1}{NO_3}$$

Step 2: Check solubility of both products. If any are insoluble according to Table F you have found the precipitate.

Ag_2CrO_4

Concentrations of Solutions

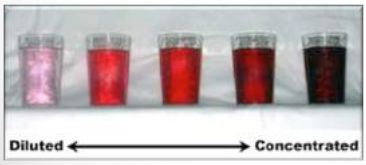
Video Lesson 11.4

Objectives

- Solve problems involving the molarity of a solution.
- Define percent by volume and percent by mass solutions

Concentration

- The amount of solute dissolved in a given solvent
- Qualitative!
- Concentrated = contains large amounts of solute
- Dilute = small amounts of solute



Dilute = $\frac{\text{solute}}{\text{SOLVENT}}$

Concentrated = $\frac{\text{SOLUTE}}{\text{solvent}}$

Molarity

- Quantitative
- The concentration of a solution may be expressed in Molarity

Table T

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Example #1

- If 2.50 moles of concentrated HCl are diluted with water to a volume of 1.79L, what is the molarity of the solution?

$$\text{Molarity} = \frac{2.50 \text{ moles}}{1.79 \text{ L}}$$

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Molarity = 1.40 M

Molar

Example #2

- How many moles of solute are needed to prepare 2.5 liters of a 0.60 M solution?

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$0.60 = \frac{x}{2.5 \text{ L}}$$

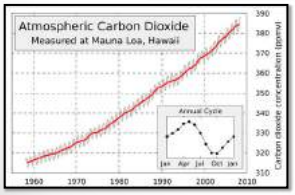
x = (0.60)(2.5L)
x = 1.5 moles of solute

Parts per Million

- Concentrations that are very dilute solutions are expressed in ppm

Table T

$$\text{parts per million} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1000000$$



Example #3

- What is the concentration of a solution in parts per million that contains 0.194 g of $C_6H_{12}O_6$ in 256 g of solution?

$$\text{parts per million} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 1000000$$

$$\text{ppm} = \frac{0.194 \text{ g}}{256 \text{ g}} \times 1{,}000{,}000$$

Ans. = 757.8 ppm

Percent by Mass

- The concentration of a solute dissolved in a solution can be found by using percent by mass.

$$\% \text{ mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Example #4

- What is the percent by mass of a solution that contains 65.3 g of NaOH dissolved in enough water to produce 178 g of solution?

$$\% \text{ mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

$$\% \text{ mass} = \frac{65.3 \text{ g}}{178 \text{g}} \times 100$$

% mass = 36.7%

Electrolytes and Colligative Properties of Solutions

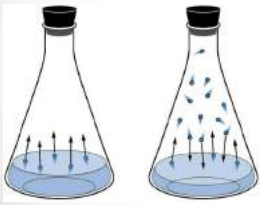
Chemistry 200
Video Lesson 11.5

Objective:

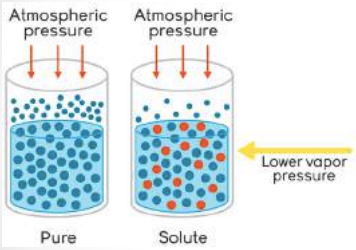
How are physical properties such as boiling point, freezing point and vapor pressure effected by the number of solute particles in solution or solute concentration?

Colligative Properties

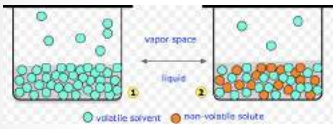
1. Vapor Pressure
   - the pressure of a vapor over its liquid in a closed container or at equilibrium.



- the greater the number of solute particles in a solvent, the lower the vapor pressure will be

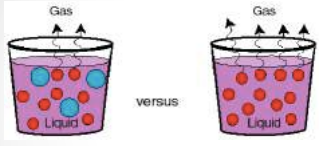


2. Freezing Point Depression
   - the freezing point of a solvent will decrease ***(get colder)*** when a solute is added to it.



Solute molecules block solvent molecules from organizing. More energy must be released to solidify, therefore the freezing point or temperature is lower.

3. Boiling Point Elevation
   - the boiling point of a solvent will increase ***(get hotter)*** when solute is added to it



Solute particles decrease vapor pressure. More energy must be absorbed to allow vapor pressure to equal atm. pressure. This causes the boiling point or temperature to increase.

Counting Particles

A. The influence of solute particles depends only on the number of particles.

B. Molecular and ionic compounds will produce *different* numbers of particles per mole of substance.

1 mole of a molecular solid → 1 mole of particles

$C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$

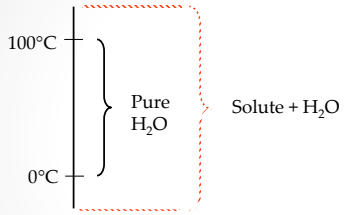
1 mole of an ionic solid → 2 moles of particles

$NaCl_{(s)} \rightarrow Na^{+1}_{(aq)} + Cl^{-1}_{(aq)}$

1 mole of an ionic solid → 3 moles of particles

$MgCl_{2(s)} \rightarrow Mg^{+2}_{(aq)} + 2\ Cl^{-1}_{(aq)}$

The more moles of particles produced, the more effect it has on F.P. and B.P.



Sketch Notes

Sketch Notes

Name _____

Solubility of Gases

Guiding Question: How does a change in temperature affect the solubility of gases?

Figure 1.

Pre-Demonstration Questions

1. What substance makes Coca-Cola bubbly?

Demonstration/ Phenomenon: Record your observations as the soda is placed into different temperatures of water.

	Temperature in °C	Observations
Cold Water		
Room Temperature		
Hot Water		

Model: Model what you think is going on in the three beakers. **Be sure to include labels, particles and/or energy flow arrows.**

Video 11.1: What are solutions?

Answer the following questions

1. _____ In an aqueous solution of potassium chloride, the solute is
   1) K 2) Cl 3) KCl 4) H_2O

2. _____ Which sample of matter is classified as a solution?
   1) H_2O (s) 2) H_2O (l) 3) CO_2(g) 4) CO_2 (aq)

3. _____ Which formula represents a homogeneous mixture?
   1) H_2O (l) 2) HCl (aq) 3) NaH (s) 4) H_2S (g)

4. _____ In a true solution, the dissolved particles
   1. are visible to the eye
   2. will settle out on standing
   3. are always solids
   4. cannot be removed by filtration

5. _____ Which of the following is a solution:
   1. table salt 2. steel 3. water 4. mercury

6. ___ Give an example of a solution for each of the following:

	Solute	Solvent	
1.	gas	gas	→
2.	solid	liquid	→
3.	liquid	liquid	→
4.	gas	liquid	→
5.	solid	solid	→

7. How is dissociation different than dissolving? What type of substances dissociate?

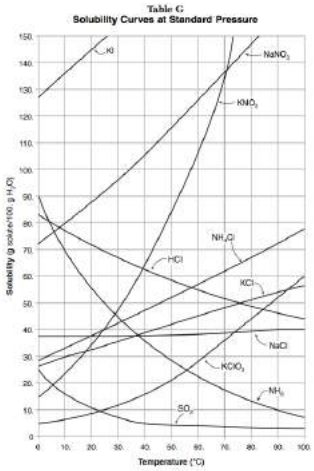
8. Which of the following is considered an electrolyte?
   1. H_2O 2. CH_3OH 3. $MgCl_2$ 4. CCl_4

9. On the line at the right, write the number of the definition that best matches each term.

a) solution _____	1) capable of being dissolved
b) solute _____	2) solution with water as the solvent
c) solvent _____	3) substance that is dissolved in a solution
d) soluble _____	4) solid solution containing two or more metals
e) aqueous solution ____	5) homogeneous mixture of two or more substances in a single physical state
f) alloy _____	6) substance that does the dissolving in a solution

Video 11.2: Factors that Affect Solubility & Table G

The solubility of solid solutes generally increases as temperature increases, while the solubility of gaseous solutes generally decreases as temperature increases. A solution that holds as much solute as can dissolve at a given temperature is saturated. A solution that can dissolve more solute at a given temperature is unsaturated. A solution that holds more solute than can dissolve at a given temperature is supersaturated. The amount of solute that is needed to form a saturated solution at various temperatures can be graphed. This is what is shown in Table G. The values in Table G are based on solute dissolved in 100 g of water. Since water has a density of 1 g/mL, the graph can be based on 100 mL of water. A 200 mL sample of water would be able to dissolve twice as much at each temperature.



Answer the questions below based on Table G.

1. The compound which is most soluble at 20°C _____
2. The compound which is least soluble at 10°C _____
3. The compound which is least soluble at 80°C _____
4. How many grams of potassium nitrate are needed to saturate 100 ml of water at 70°C?

5. Write the 3 formulas for the compounds that have an inverse relationship between temperature and solubility. _____, _____ & _____

6. Which 2 salts have the same solubility at 71°C? _____ & _____

For each question an amount of solute is given and a temperature is stated. If all the solute could be dissolved in 100g of water at the stated temperature, would the resulting solution be ***unsaturated, saturated or supersaturated?***

7. 70 g of KCl at 60°C _____
8. 90 g of KNO_3 at 70°C _____
9. 110 g of $NaNO_3$ at 45°C _____
10. 5 g of $KClO_3$ at 10°C _____
11. 60 g of NH_3 at 80°C _____

For each question a solute and temperature are given. Tell how many grams of each solute must be added to 100g of water to form a saturated solution.

12. $NaNO_3$ at 40°C _____
13. $KClO_3$ at 80°C _____
14. KNO_3 at 45°C _____
15. KCl at 50°C _____
16. NaCl at 90°C _____
17. If 50 g of $KClO_3$ are added to 100 g of water at 10°C, how many grams do not dissolve?

11.1 - 11.2 Practice

1. Which compound becomes *less* soluble in water as the temperature of the solution is increased?

   1) HCl 2) KCl
   3) NaCl 4) NH_4Cl

2. An unsaturated aqueous solution of NH_3 is at 90°C in 100. grams of water. According to Reference Table *G*, how many grams of NH_3 could this unsaturated solution contain?

   1) 5 g 2) 10. g 3) 15 g 4) 20. g

3. According to your Reference Tables, which substance forms an unsaturated solution when 80 grams of the substance is dissolved in 100 grams of H_2O at 10°C?

   1) KI 2) KNO_3
   3) $NaNO_3$ 4) NaCl

4. The solubility of KCl(s) in water depends on the

   1) pressure on the solution
   2) rate of stirring
   3) size of the KCl sample
   4) temperature of the water

5. Under which conditions of temperature and pressure is a gas most soluble in water?

   1) high temperature and low pressure
   2) high temperature and high pressure
   3) low temperature and low pressure
   4) low temperature and high pressure

6. At room temperature, the solubility of which solute in water would be most affected by a change in pressure?

   1) methanol 2) sugar
   3) carbon dioxide 4) sodium nitrate

7. What is the total mass of KNO_3 that must be dissolved in 50. grams of H_2O at 60.°C to make a saturated solution?

   1) 32 g 2) 53 g 3) 64 g 4) 106 g

8. What is the mass of NH_4Cl that must dissolve in 200. grams of water at 50.°C to make a saturated solution?

   1) 26 g 2) 42 g 3) 84 g 4) 104 g

9. An unsaturated solution is formed when 80. grams of a salt is dissolved in 100. grams of water at 40.°C. This salt could be

   1) KCl 2) KNO_3
   3) NaCl 4) $NaNO_3$

10. A solution contains 35 grams of KNO_3 dissolved in 100 grams of water at 40°C. How much *more* KNO_3 would have to be added to make it a saturated solution?

    1) 29 g 2) 24 g 3) 12 g 4) 4g

18. What mass of NH_4Cl would be needed to form a saturated solution if the NH_4Cl was dissolved in 200 g of water at 50°C?

19. What mass of NaCl would be needed to form a saturated solution at 30°C if the NaCl was dissolved in 35 g of water?

20. At what temperature would you need 100g of water to dissolve 70g NH_4Cl?

21. A supersaturated solution of $NaNO_3$ is created by adding 120 g of $NaNO_3$ to 100 g of water at 20°C. If the solution is disrupted, how many grams of solid $NaNO_3$ will "fall out" (will not dissolve) of the solution?

Video 11.3: Double Replacement Reactions & Table F

1. Which barium salt is insoluble in water?
    1. $BaCO_3$
    2. $BaCl_2$
    3. $Ba(ClO_3)_2$
    4. $Ba(NO_3)_2$

2. Which compound is insoluble in water?
    1. Calcium bromide
    2. Potassium bromide
    3. Silver bromide
    4. Sodium bromide

3. Based on reference table F, which salt is least soluble?
    1. $FeCO_3$
    2. Na_2CO_3
    3. $BaCl_2$
    4. $CaCl_2$

Table F
Solubility Guidelines for Aqueous Solutions

Ions That Form *Soluble* Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, or Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, or Pb^{2+}

Ions That Form *Insoluble* Compounds*	Exceptions
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)

*compounds having very low solubility in H_2O

11.3 Table F Practice

1. Which compound is insoluble in water?

   1) calcium bromide
   2) potassium bromide
   3) silver bromide
   4) sodium bromide

2. Which barium salt is *insoluble* in water?

   1) $BaCO_3$ 2) $BaCl_2$
   3) $Ba(ClO_3)_2$ 4) $Ba(NO_3)_2$

3. Which ion, when combined with chloride ions, Cl^-, forms an insoluble substance in water?

   1) Fe^{2+} 2) Mg^{2+} 3) Pb^{2+} 4) Zn^{2+}

4. Based on Reference Table *F*, which of these saturated solutions has the lowest concentration of dissolved ions?

   1) NaCl(aq) 2) $MgCl_2$(aq)
   3) $NiCl_2$(aq) 4) AgCl(aq)

5. According to Reference Table F, which substance is most soluble?

   1) AgI 2) $CaSO_4$
   3) $PbCl_2$ 4) $(NH_4)_2CO_3$

6. The attraction between water molecules and an Na^+ ion or a Cl^- ion occurs because water molecules are

   1) linear 2) symmetrical
   3) polar 4) nonpolar

7. According to Reference Table F, which compound is most soluble in water?

   1) $BaCO_3$ 2) $BaSO_4$
   3) $ZnCO_3$ 4) $ZnSO_4$

8. Which compound is most soluble in water?

   1) silver sulfate 2) silver chloride
   3) silver nitrate 4) silver hydroxide

9. Which compound is insoluble in water?

   1) KOH 2) NH_4Cl
   3) Na_3PO_4 4) $PbSO_4$

10. According to Table *F*, which ions combine with chloride ions to form an insoluble compound?

    1) Fe^{2+} ion 2) Ca^{2+} ions
    3) Li^+ ions 4) Ag^+ ions

Table F can also be used to write balanced equations for the following double replacement reactions. Complete, ***balance and circle the precipitate in the following reactions.***

NOTE:

- A soluble compound is followed by the symbol (aq)
- The formation of an insoluble compound (precipitate) is indicated with a (s)
- If no precipitate forms, no reaction occurred between the ions dissolved in solution.

1. _____ $Li_3PO_{4\,(aq)}$ + _____$NH_4Br_{(aq)}$ →

2. _____ $FeCl_{2(aq)}$ + _____ $KOH_{\,(aq)}$ →

3. _____ $MgCl_{2\,(aq)}$ + _____ $CsCl_{\,(aq)}$ →

4. _____ $Al(ClO_3)_{3\,(aq)}$ + _____ $MgSO_{4\,(aq)}$ →

5. _____ $Na_2CO_{3\,(aq)}$ + _____ $Ba(ClO_3)_{2\,(aq)}$ →

6. _____ $Ca(OH)_{2(aq)}$ + _____ $Al(NO_3)_{3\,(aq)}$ -->

7. _____ $K_3PO_{4\,(aq)}$ + _____$Cs_2S_{(aq)}$ →

8. _____ $CaCrO_{4\,(aq)}$ + _____ $(NH_4)_2CO_{3\,(aq)}$ →

11.4 Molarity Practice:

Directions: *Solve the following problems. Include the equation used and show all work. Please state the answer to the correct number of significant figures and box all answers with proper units.*

1. What is the molarity of a solution that contains 0.40 moles of KBr in 0.50 L of solution?

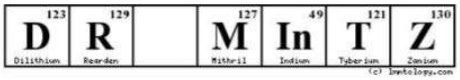
2. If you have 5.0 moles of NaCl in a 2.0 L solution, what is the molarity of the solution?

3. If you have 60. moles of HCl, what should the total volume of the solution be to make a 10. M solution of $HCl_{(aq)}$.

4. Which solution is most concentrated?
    1. 5 M HCl
    2. 0.09 M HCl
    3. 3 M HCl
    4. 23 M HCl

5. Which solution is the most dilute?
    1. 5 M HCl
    2. 0.09 M HCl
    3. 3 M HCl
    4. 23 M HCl

6. What is the molarity of a solution that contains 65.1 g of NH_4Cl in 3.50 L of solution?



Video 11.4: Concentration

Perform the following calculations and express your answers using the proper units and the correct amount of significant figures. Be sure to show work for credit.

1. What is the percent by mass of a solution that contains 47.8 g of KOH dissolved in enough water to produce 293 g of solution?

2. What is the percent by mass of a solution that contains 5.61 g of H_2SO_4 dissolved in 142 g of H_2O

3. How many grans of KNO_3 must be dissolved in 325 g of solution, if the solution contains 14% by mass of KNO_3

4. What is the concentration of a solution in parts per million that contains 0.118 g of $C_{12}H_{22}O_{11}$ dissolved in 237 g of solution?

5. What is the concentration of a solution in parts per million that contains 0.472 g of Li_3PO_4 dissolved in 598 g of H_2O

6. How many grams of MgF_2 must be dissolved in water to produce 306 g of a 14.5 ppm solution

Name: **Concentration (Molarity & ppm)**

___ 1. The molarity of an aqueous solution of NaCl is defined as the

1) grams of NaCl per liter of water
2) grams of NaCl per liter of solution
3) moles of NaCl per liter of water
4) moles of NaCl per liter of solution

___ 2. What is the molarity of 1.5 liters of an aqueous solution that contains 52 grams of lithium fluoride, LiF, (gram-formula mass = 26 grams/mole)?

1) 1.3 M
2) 2.0 M
3) 3.0 M
4) 0.75 M

___ 3. A 3.0 M HCl(aq) solution contains a total of

1) 3.0 grams of HCl per liter of water
2) 3.0 grams of HCl per mole of solution
3) 3.0 moles of HCl per liter of solution
4) 3.0 moles of HCl per mole of water

___ 4. How many total moles of KNO_3 must be dissolved in water to make 1.5 liters of a 2.0 M solution?

1) 0.50 mol
2) 2.0 mol
3) 3.0 mol
4) 1.3 mol

___ 5. What is the total number of grams of NaI(s) needed to make 1.0 liter of a 0.010 M solution?

1) 0.015
2) 0.15
3) 1.5
4) 15

___ 6. Which unit can be used to express the concentration of a solution?

1) L/s 2) J/g 3) ppm 4) kPa

___ 7. What is the total mass of solute in 1000. grams of a solution having a concentration of 5 parts per million?

1) 0.005 g
2) 0.05g
3) 0.5 g
4) 5g

___ 8. What is the concentration of $O_2(g)$, in parts per million, in a solution that contains 0.008 gram of $O_2(g)$ dissolved in 1000. grams of $H_2O(l)$?

1) 0.8 ppm
2) 8 ppm
3) 80 ppm
4) 800 ppm

___ 9. What is the concentration of a solution, in parts per million, if 0.02 gram of Na_3PO_4 is dissolved in 1000 grams of water?

1) 20 ppm
2) 2 ppm
3) 0.2 ppm
4) 0.02 ppm

___ 10. What is the molarity of a solution containing 20 grams of NaOH in 500 milliliters of solution?

1) 1 M
2) 2 M
3) 0.04 M
4) 0.5 M

Video 11.5: Colligative Properties

Complete the following reactions and balance

1. How many particles are formed when 1 molecule of each of the following substances is dissolved in 1000g of H_2O?

   a) $CaBr_{2(s)} \rightarrow$

   b) $C_6H_{12}O_{6(s)} \rightarrow$

   c) $Li_3PO_{4(s)} \rightarrow$

   d) $NH_4NO_{3(s)} \rightarrow$

   e) $CH_3OH_{(l)} \rightarrow$

2. Which of the compound(s) above cause(s) the freezing point of water to decrease the most? Why?

3. Which of the compound(s) above cause(s) the boiling point of water to increase the least? Why?

4. Explain why do we put salt down on the roads when it snows instead of sugar?

Name: **Colligative Properties**

___ 1. Which solution has the lowest freezing point?

1) 10. g of KI dissolved in 100. g of water
2) 20. g of KI dissolved in 200. g of water
3) 30. g of KI dissolved in 100. g of water
4) 40. g of KI dissolved in 200. g of water

___ 2. Compared to a 2.0 M aqueous solution of NaCl at 1 atmosphere, a 3.0 M aqueous solution of NaCl at 1 atmosphere has a

1) lower boiling point and a higher freezing point
2) lower boiling point and a lower freezing point
3) higher boiling point and a higher freezing point
4) higher boiling point and a lower freezing point

___ 3. Compared to a 0.1 M aqueous solution of NaCl, a 0.8 M aqueous solution of NaCl has a

1) higher boiling point and a higher freezing point
2) higher boiling point and a lower freezing point
3) lower boiling point and a higher freezing point
4) lower boiling point and a lower freezing point

___ 4. At standard pressure when NaCl is added to water, the solution will have a

1) higher freezing point and a lower boiling point than water
2) higher freezing point and a higher boiling point than water
3) lower freezing point and a higher boiling point than water
4) lower freezing point and a lower boiling point than water

___ 5. As a solute is added to a solvent, what happens to the freezing point and the boiling point of the solution?

1) The freezing point decreases and the boiling point decreases.
2) The freezing point decreases and the boiling point increases.
3) The freezing point increases and the boiling point decreases.
4) The freezing point increases and the boiling point increases.

___ 6. Which sample, when dissolved in 1.0 liter of water, produces a solution with the *lowest* freezing point?

1) 0.1 mol of C_2H_5OH
2) 0.1 mol of LiBr
3) 0.2 mol of $C_6H_{12}O_6$
4) 0.2 mol of $CaCl_2$

___ 7. Which aqueous solution has the *lowest* freezing point?

1) 1.0 M $C_6H_{12}O_6$
2) 1.0 M C_2H_5OH
3) 1.0 M CH_3COOH
4) 1.0 M NaCl

___ 8. Which solution has the highest boiling point?

1) 1.0 M KNO_3
2) 2.0 M KNO_3
3) 1.0 M $Ca(NO_3)_2$
4) 2.0 M $Ca(NO_3)_2$

Name: **Solutions Review**

1. Which barium salt is *insoluble* in water?

   1) $BaCO_3$ 3) $Ba(ClO_4)_2$
   2) $BaCl_2$ 4) $Ba(NO_3)_2$

2. Based on Reference Table *F*, which of these saturated solutions has the lowest concentration of dissolved ions?

   1) NaCl(aq) 3) $NiCl_2$(aq)
   2) $MgCl_2$(aq) 4) AgCl(aq)

3. According to Reference Table F, which substance is most soluble?

   1) AgI 3) $PbCl_2$
   2) $CaSO_4$ 4) $(NH_4)_2CO_3$

4. Which compound becomes *less* soluble in water as the temperature of the solution is increased?

   1) HCl 3) NaCl
   2) KCl 4) NH_4Cl

5. According to Table *F* which compound is soluble in water?

   1) barium phosphate
   2) calcium sulfate
   3) silver iodide
   4) sodium perchlorate

6. Which compound is *least* soluble in water at 60. °C?

   1) $KClO_3$ 3) NaCl
   2) KNO_3 4) NH_4Cl

7. According to your Reference Tables, which substance forms an unsaturated solution when 80 grams of the substance is dissolved in 100 grams of H_2O at 10°C?

   1) KI 3) $NaNO_3$
   2) KNO_3 4) NaCl

8. The solubility of $KClO_3$(s) in water increases as the

   1) temperature of the solution increases
   2) temperature of the solution decreases
   3) pressure on the solution increases
   4) pressure on the solution decreases

9. According to Reference Table *G*, how many grams of KNO_3 would be needed to saturate 200 grams of water at 70ºC?

   1) 43 g 2) 86 g 3) 134 g 4) 268 g

10. Based on Reference Table *G*, what is the maximum number of grams of KCl(s) that will dissolve in 200 grams of water at 50°C to produce a saturated solution?

    1) 38 g 2) 42 g 3) 58 g 4) 84 g

11. The solubility of KCl(s) in water depends on the

    1) pressure on the solution
    2) rate of stirring
    3) size of the KCl sample
    4) temperature of the water

12. Under which conditions of temperature and pressure is a gas most soluble in water?

    1) high temperature and low pressure
    2) high temperature and high pressure
    3) low temperature and low pressure
    4) low temperature and high pressure

13. A student prepares four aqueous solutions, each with a different solute. The mass of each dissolved solute is shown in the table below.

Mass of Dissolved Solute for Four Aqueous Solutions

Solution Number	Solute	Mass of Dissolved Solute (per 100. g of H_2O at 20.°C)
1	KI	120. g
2	$NaNO_3$	88 g
3	KCl	25 g
4	$KClO_3$	5 g

Which solution is saturated?

1) 1 2) 2 3) 3 4) 4

14. What is the total mass of KNO_3 that must be dissolved in 50. grams of H_2O at 60.°C to make a saturated solution?

    1) 32 g 2) 53 g 3) 64 g 4) 106 g

15. When 5 grams of KCl are dissolved in 50. grams of water at 25°C, the resulting mixture can be described as

    1) heterogeneous and unsaturated
    2) heterogeneous and supersaturated
    3) homogeneous and unsaturated
    4) homogeneous and supersaturated

16. A solution contains 35 grams of KNO_3 dissolved in 100 grams of water at 40°C. How much *more* KNO_3 would have to be added to make it a saturated solution?

    1) 29 g 2) 24 g 3) 12 g 4) 4g

17. A solution is formed by dissolving 45 grams of NH_4Cl in 100 grams of H_2O at 70°C. Which statement correctly describes this solution?

    1) NH_4Cl is the solute, and the solution is saturated.
    2) NH_4Cl is the solute, and the solution is unsaturated.
    3) NH_4Cl is the solvent, and the solution is saturated.
    4) NH_4Cl is the solvent, and the solution is unsaturated.

18. The molarity of an aqueous solution of NaCl is defined as the

    1) grams of NaCl per liter of water
    2) grams of NaCl per liter of solution
    3) moles of NaCl per liter of water
    4) moles of NaCl per liter of solution

19. What is the molarity of 1.5 liters of an aqueous solution that contains 52 grams of lithium fluoride, LiF, (gram-formula mass = 26 grams/mole)?

    1) 1.3 M
    2) 2.0 M
    3) 3.0 M
    4) 0.75 M

20. Which phrase describes the molarity of a solution?

    1) liters of solute per mole of solution
    2) liters of solution per mole of solution
    3) moles of solute per liter of solution
    4) moles of solution per liter of solution

21. A 3.0 M HCl(aq) solution contains a total of

    1) 3.0 grams of HCl per liter of water
    2) 3.0 grams of HCl per mole of solution
    3) 3.0 moles of HCl per liter of solution
    4) 3.0 moles of HCl per mole of water

22. How many total moles of KNO_3 must be dissolved in water to make 1.5 liters of a 2.0 M solution?

    1) 0.50 mol
    2) 2.0 mol
    3) 3.0 mol
    4) 1.3 mol

23. What is the molarity of a solution of NaOH if 2 liters of the solution contains 4 moles of NaOH?

    1) 0.5 M 2) 2 M 3) 8 M 4) 80 M

24. What is the total number of grams of NaI(s) needed to make 1.0 liter of a 0.010 M solution?

    1) 0.015 2) 0.15 3) 1.5 4) 15

25. Which unit can be used to express the concentration of a solution?

    1) L/s 2) J/g 3) ppm 4) kPa

26. What is the total mass of solute in 1000. grams of a solution having a concentration of 5 parts per million?

    1) 0.005 g
    2) 0.05g
    3) 0.5 g
    4) 5g

27. What is the concentration of $O_2(g)$, in parts per million, in a solution that contains 0.008 gram of $O_2(g)$ dissolved in 1000. grams of $H_2O(l)$?

    1) 0.8 ppm
    2) 8 ppm
    3) 80 ppm
    4) 800 ppm

28. What is the concentration of a solution, in parts per million, if 0.02 gram of Na_3PO_4 is dissolved in 1000 grams of water?

    1) 20 ppm
    2) 2 ppm
    3) 0.2 ppm
    4) 0.02 ppm

29. How do the boiling point and freezing point of a solution of water and calcium chloride at standard pressure compare to the boiling point and freezing point of water at standard pressure?

    1) Both the freezing point and boiling point of the solution are higher.
    2) Both the freezing point and boiling point of the solution are lower.
    3) The freezing point of the solution is higher and the boiling point of the solution is lower.
    4) The freezing point of the solution is lower and the boiling point of the solution is higher.

30. Which aqueous solution of KI freezes at the lowest temperature?

    1) 1 mol of KI in 500. g of water
    2) 2 mol of KI in 500. g of water
    3) 1 mol of KI in 1000. g of water
    4) 2 mol of KI in 1000. g of water

31. Which solution has the lowest freezing point?

    1) 10. g of KI dissolved in 100. g of water
    2) 20. g of KI dissolved in 200. g of water
    3) 30. g of KI dissolved in 100. g of water
    4) 40. g of KI dissolved in 200. g of water

32. Compared to a 0.1 M aqueous solution of NaCl, a 0.8 M aqueous solution of NaCl has a

    1) higher boiling point and a higher freezing point
    2) higher boiling point and a lower freezing point
    3) lower boiling point and a higher freezing point
    4) lower boiling point and a lower freezing point

Base your answers to questions **33** and **34** on the information below and on your knowledge of chemistry.

Seawater contains dissolved salts in the form of ions. Some of the ions found in seawater are $Ca^{2+}, Mg^{2+}, K^{+}, Na^{+}, Cl^{-}, HCO_3^{-}$, and SO_4^{2-}

An investigation was conducted to determine the concentration of dissolved salts in seawater at one location. A 300.-gram sample of the seawater was placed in an open container. After a week, all the water had evaporated and 10. grams of solid salts remained in the container.

33. At standard pressure, compare the freezing point of seawater to the freezing point of distilled water.

34. Determine the concentration, expressed as percent by mass, of the dissolved salts in the original sample of seawater.

---

35. Base your answer to the following question on the information below.

A total of 1.4 moles of sodium nitrate is dissolved in enough water to make 2.0 liters of an aqueous solution. The gram-formula mass of sodium nitrate is 85 grams per mole.

Determine the molarity of the solution.

36. A scientist makes a solution that contains 44.0 grams of hydrogen chloride gas, HCl(g), in 200. grams of water, $H_2O(\ell)$, at 20. °C. This process is represented by the balanced equation below.

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

Based on Reference Table *G*, identify, in terms of saturation, the type of solution made by the scientist.

Base your answers to questions **37** through **40** on the information below.

In a laboratory, a student makes a solution by completely dissolving 80.0 grams of $KNO_3(s)$ in 100.0 grams of hot water. The resulting solution has a temperature of 60.°C. The room temperature in the laboratory is 22°C.

37. Describe a laboratory procedure that can be used to recover the solid solute from the aqueous solution.

38. Describe the direction of heat flow between the solution made by the student and the air in the laboratory.

39. Compare the boiling point of the solution at standard pressure to the boiling point of water at standard pressure.

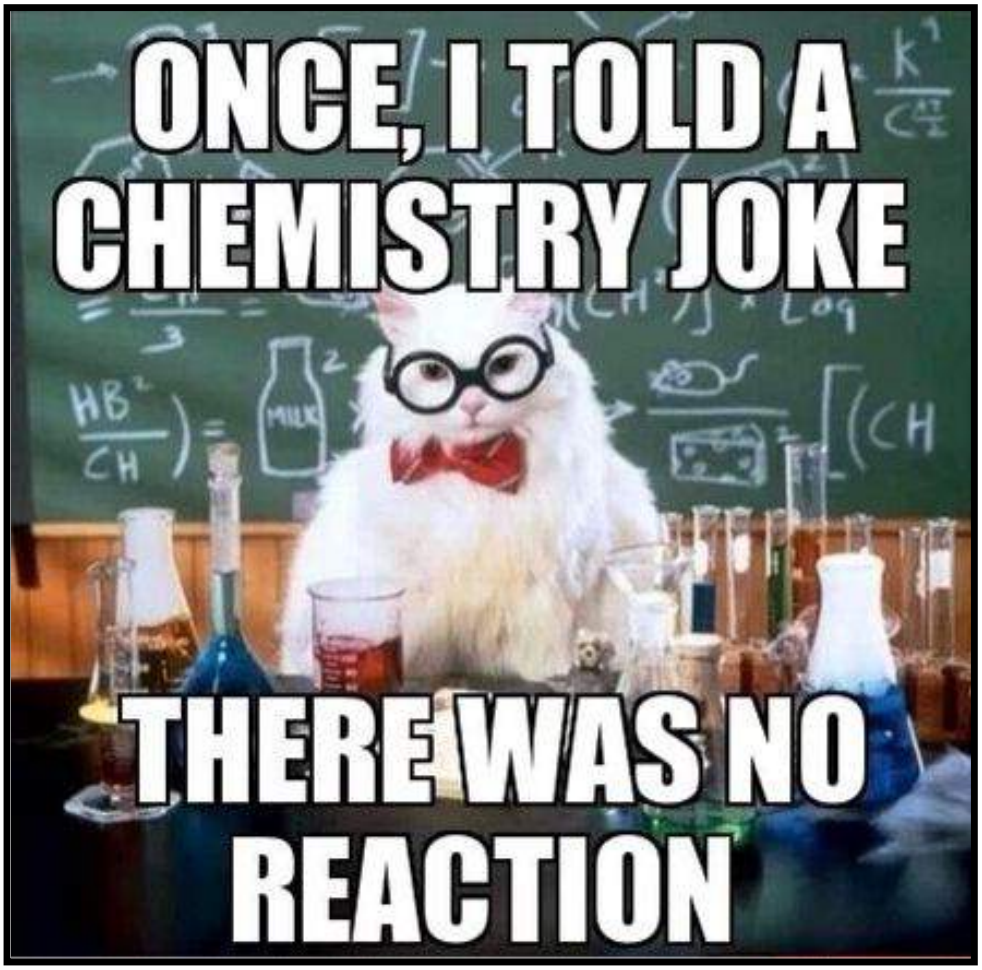
40. Classify, in terms of saturation, the type of solution made by the student.

Name:

Regents Chemistry

Practice Packet

Chapter 12: Kinetics & Equilibrium



Regents Chemistry **Name _____**

VOCAB – Kinetics & Equilibrium **Period ____**

1. Reaction Rate – the speed at which reactants are converted into products in a chemical reaction.
2. Collision Theory – in order for a chemical reaction/effective collision to occur, particles must collide with proper energy AND proper alignment.
3. Reaction Mechanism – the specific set of steps/reactions involved in an overall chemical reaction

(#4-9 are defined in terms of their GENERAL EFFECT on the rate chemical of a reaction)

4. Nature of Reactants – reactions involving ionic substances tend to have faster rates than reactions involving covalent substances.
5. Concentration – an increase in concentration of reactants will increase the rate of a chemical reaction
6. Surface Area – an increase in the surface area of reactants will increase the rate of a chemical reaction
7. Pressure – an increase in pressure will increase the rate of a chemical reaction (only for reactions involving GASES!)
8. Catalyst – a substance that is neither a reactant nor a product, but functions to speed up the rate of a chemical reaction by lowering activation energy/providing a shorter or "alternate" pathway
9. Temperature – an increase in temperature will increase the rate of a chemical reaction
10. Equilibrium – when two opposing processes are occurring at equal rates
11. Physical Equilibrium – when two opposing physical processes are occurring at equal rates; ex: phase equilibrium, solution equilibrium (saturation)
12. Phase Equilibrium – when the processes of freezing and melting or evaporating and condensing are occurring at equal rates
13. Solution Equilibrium – when the processes of dissolving and precipitating are occurring at equal rates; when a solution has reached its saturation point
14. Chemical Equilibrium – in a chemical reaction, when the forward and reverse reactions are occurring at equal rates
15. Le Chatelier's Principle – predicts that when a stress is applied to an equilibrium mixture, the equilibrium will shift to relieve the stress (stresses include temperature, pressure, concentration)
16. Enthalpy – the heat energy absorbed or released during a chemical reaction
17. Entropy – a measure of the randomness or chaos associated with a chemical reaction
18. Potential Energy Diagrams – used to illustrated the energy lost or gained (the reaction pathway) for a given chemical reaction
19. Endothermic Reactions – chemical reactions that consume or require energy; chemical reactions in which energy is a reactant

20. Exothermic Reactions – chemical reactions that produce or release energy; chemical reactions in which energy is a product
21. Activated Complex – an intermediate structure formed in the conversion of reactants to products. The activated complex is the structure at the maximum energy point along the reaction path
22. Activation Energy – The minimum energy required to convert reactants into products; the difference between the energies of the activated complex and the reactants

Factors that affect the rate of chemical reactions

Chemistry 200
Video Lesson 12.1

Objective:

How can different factors change the rate of a chemical reaction?

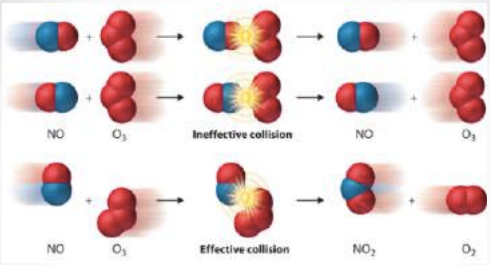
Kinetics

- the branch of chemistry dealing w/ rates of chem reaction

Collision Theory

- reactions occur when reactant particles collide. Anything that increases the number of collisions will increase the rate of the reaction.

Reactions only occur if the colliding particles positioning (orientation) is correct & they have enough kinetic energy to support the rxn. If either or both do not occur, no products are formed & no reaction has not occurred



Factors Affecting Rate of Rxn

1. Nature of Reactant

- covalent or molecular substances react slower than ionic substances because more bonds are present
- if more energy is needed to break bonds, less energy is available for reaction

2. Concentration

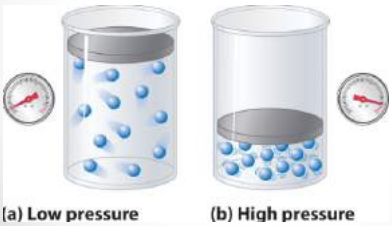
- rate of reaction will be faster if the concentration of one or more of the reactants is increased

3. Surface Area

- the greater the surface area exposed, the more possibility of collisions, therefore an increased rate of reaction

4. Pressure (little to no effect on solids or liquids, HUGE effect on gases)

- increasing pressure on a gas is like increasing concentration, resulting in an increased rate of reaction. The same mass of gas is squeezed into a smaller volume.



5. Presence of a Catalyst

- a catalyst is a substance that increases the rate of a reaction by creating an alternate pathway. This uses less energy & does not alter the product. Enzymes are a common example.

**** catalyst remains unchanged****

6. Temperature (measure of avg. Kinetic Energy)

- an increase results in increased K.E., therefore more mvmt & collisions of particles --> increase rate of reaction

Potential Energy Diagrams

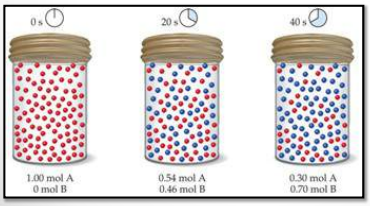
Chemistry 200
Video Lesson 12.2

Objectives

- Describe how to express the rate and heat of reaction using potential energy diagrams.

Rate of a Chemical Reaction

- The rate of a chemical change or the reaction rate is usually expressed as the amount of reactant changing per unit time.
- The more effective collisions between reacting particles the faster the rate of change from reactants to products.



Potential Energy Diagram

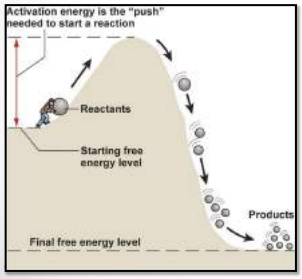
- Chemical Bonds have stored energy (AKA potential energy)
- Potential Energy Diagrams used to illustrate how energy changes during a chemical reaction

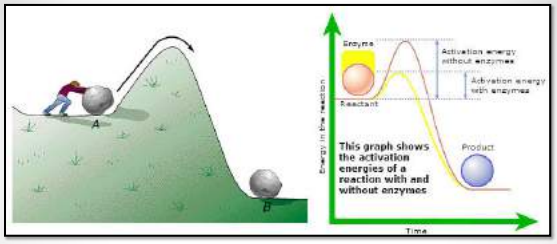
A + B → AB + energy

AB + energy → A + B

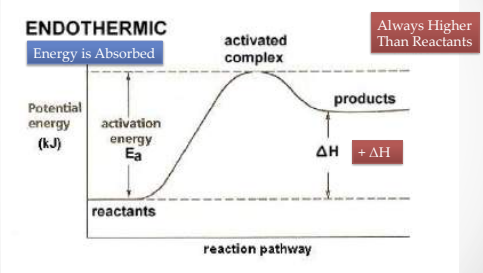
Activation Energy

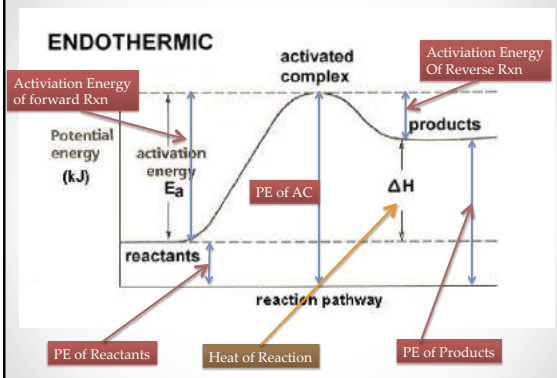
- The minimum energy that colliding particles must have in order to react is called the activation energy.



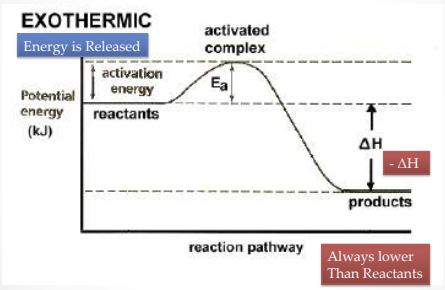


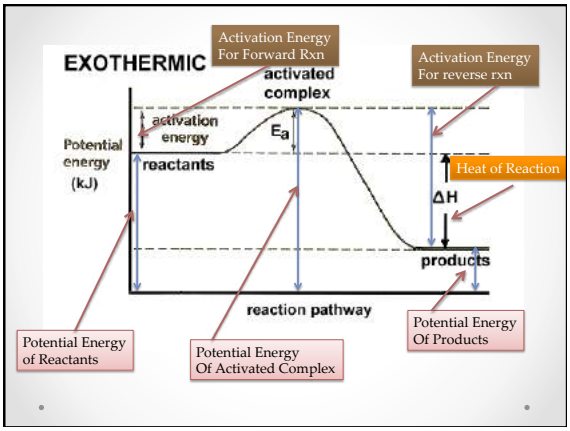
Potential Energy Diagram Endothermic

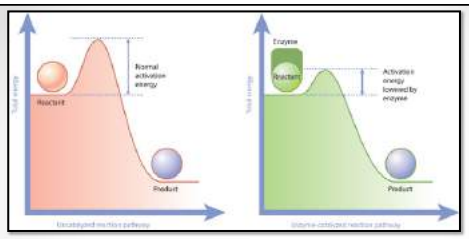


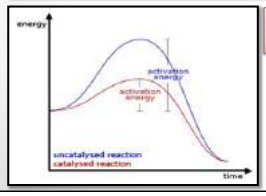


Potential Energy Diagram Exothermic









A Catalyst (enzyme): Lowers the amount of activation energy

Enthalpy & Entropy

Chemistry 200
Video Lesson 12.3

Objective:

How do we determine the heat of reaction, if particle disorder is increasing or decreasing and how these two factors influence the spontaneity of the reaction?

Enthalpy vs. Entropy

Enthalpy (Heat of reaction)

$\Delta H = H_P - H_R$ (*change in P.E. btwn reactants & products*) **kJ**

- in nature the tendency for rxns is to move towards a lower energy state or lower enthalpy **(-ΔH)**
- reactions tend to move in an exothermic direction **(-ΔH)** rather than endothermic **(+ΔH)** **
- This is due to activation energy being less for exothermic reactions vs. endothermic reactions**

Table I
Heats of Reaction at 101.3 kPa and 298 K

Reaction	ΔH (kJ)*
$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$	−890.4
$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(\ell)$	−2219.2
$2C_4H_{10}(\ell) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(\ell)$	−10 943
$2CH_3OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(\ell)$	−1452
$C_2H_5OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(\ell)$	−1367
$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(\ell)$	−2804
$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$	−566.0
$C(s) + O_2(g) \longrightarrow CO_2(g)$	−393.5
$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$	−3351
$N_2(g) + O_2(g) \longrightarrow 2NO(g)$	+182.6
$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$	+66.4
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$	−483.6
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(\ell)$	−571.6
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$	−91.8
$2C(s) + 3H_2(g) \longrightarrow C_2H_6(g)$	−84.0
$2C(s) + 2H_2(g) \longrightarrow C_2H_4(g)$	+52.4
$2C(s) + H_2(g) \longrightarrow C_2H_2(g)$	+227.4
$H_2(g) + I_2(g) \longrightarrow 2HI(g)$	+52.0
$KNO_3(s) \xrightarrow{H_2O} K^+(aq) + NO_3^-(aq)$	+34.90
$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$	−44.51
$NH_4Cl(s) \xrightarrow{H_2O} NH_4^+(aq) + Cl^-(aq)$	+14.78
$NH_4NO_3(s) \xrightarrow{H_2O} NH_4^+(aq) + NO_3^-(aq)$	+26.69
$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$	+3.88
$LiBr(s) \xrightarrow{H_2O} Li^+(aq) + Br^-(aq)$	−48.83
$H^+(aq) + OH^-(aq) \longrightarrow H_2O(\ell)$	−55.8

*The ΔH values are based on molar quantities represented in the equations. A minus sign indicates an exothermic reaction.

- -ΔH = exothermic *(look at the bottom)*
- Coefficients represent molar quantities.
- If molar quantities change, ΔH also changes proportionally.
- If the reaction in question is opposite the reaction on Table I, ΔH is reversed.

Entropy (ΔS)

- the more disorder a system gets, the more this system moves from a state of:
  low entropy (less disorder) ---> high entropy (great disorder)
  -ΔS +ΔS

ex: Phase Δ's

(s) <----------> (l) <------------> (g) <----> (aq)

less disorder low entropy	more disorder higher entropy	max. disorder highest entropy

Determining Change in Entropy (applied in this order)

Rule one (phase rule) – a reaction that involves making more disorganized phases of matter has an increase in entropy

$H_2O_{(s)} \rightarrow H_2O_{(l)}$	Entropy increases +ΔS
$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$	Entropy increases +ΔS
$NaCl_{(s)} \rightarrow NaCl_{(aq)}$	Entropy increases +ΔS

Rule two (particle rule) – a rxn that involves creating more particles than you began w/ has an increase in entropy.

$2CO_{2(g)} \rightarrow 2CO_{(g)} + O_{2(g)}$	Entropy increases +ΔS
$3\ A_{(g)} + B_{(g)} \rightarrow 2\ C_{(g)} + D_{(g)}$	Entropy decreases -ΔS
$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$	Entropy remains the same ΔS = 0

Spontaneity **Has nothing to do w/ rate of reaction**

Will a reaction occur on its own?

Yes --> It's spontaneous

No --> It's not spontaneous, but can be forced to happen, just not by itself

Most Ideal Situation for a spontaneous rxn

low enthalpy & **high entropy**

-ΔH exothermic **+ΔS** high disorder

ΔH Enthalpy	*ΔS Entropy*	***Spontaneous?***
↓ **(-ΔH)** Exothermic rxn	↑ **(+ΔS)**	Always
↑ **(+ΔH)** Endothermic rxn	↓ **(-ΔS)**	Never
↓ **(-ΔH)** Exothermic rxn	↓ **(-ΔS)**	Sometimes at low temps $H_2O_{(l)} \rightarrow H_2O_{(s)}$ + Energy
↑ **(+ΔH)** Endothermic rxn	↑ **(+ΔS)**	Sometimes at high temps $H_2O_{(l)}$ + Energy $\rightarrow H_2O_{(g)}$

Physical & Chemical Equilibrium

Chemistry 200
Video Lesson 12.4

Objective:

How do we determine the difference between physical & chemical equilibrium?

Equilibrium

Relationship btwn Forward & Reverse Rxns

Equilibrium

- when both the forward & reverse rxns occur at the same rate & time indicated w/ a ⟷ or ⇌

  ex: A + B ⇌ C + D

- Concentrations of reactant & products of a system are at equilibrium when concentrations remain constant *(unchanged)*.
- Concentrations of reactant & product **do not** have to be equal to each other & usually are not.

Physical Equilibrium

- equilibrium occurring during a physical process like dissolving or a phase change

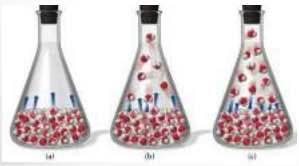
2 Types

A. **Phase Equilibrium** exists btwn the phases of the same substanc

$H_2O_{(s)} \rightleftharpoons H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$

1. freezing/melting point of a solid
2. boiling/condensation point of a liquid

} phase Δ



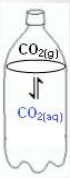
B. **Solution Equilibrium** occurs in saturated solutions where the dissolved & undissolved solutes (solids or gases) are at equilibrium with each other.

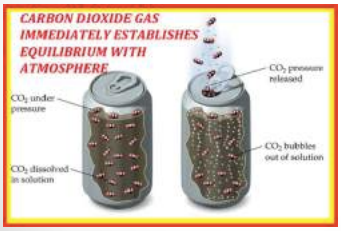
1. Solid in a liquid
   - solids dissolve in a liquid until the solid starts settling on the bottom. At this point, saturation has been reached & the system is at equilibrium.
     Rate of dissolving = Rate of recrystallization

   ex: $C_{12}H_{22}O_{11(s)} \rightleftharpoons C_{12}H_{22}O_{11(aq)}$

2. Gas in a liquid ***(closed system)***
   - equilibrium reached when a gas is dissolved in a liquid & undissolved gas remains

   ex: Closed can/bottle of soda $CO_{2(g)} \rightleftharpoons CO_{2(aq)}$



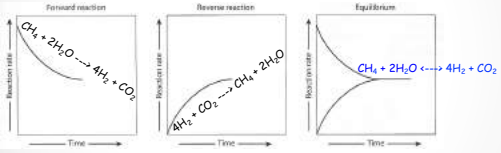


Chemical Equilibrium

- when the forward & reverse chemical reactions occur simulataneously at the same rates.

Observable changes (color, temp., pressure) no longer occur

ex: $CH_{4(g)} + 2H_2O_{(g)} \longleftrightarrow 4H_{2(g)} + CO_{2(g)}$



Le Chatelier's Principle

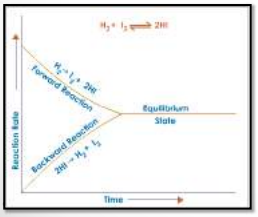
Chemistry 200
Video Lesson 12.5

Objectives

- Describe how the amounts of reactants and products change in a chemical system at equilibrium.
- Identify three stresses that can change the equilibrium position of a chemical system.

Chemical Equilibrium

- Rate of forward reaction equal rate of reverse reaction.
  - Rate at which products are made is equal to the rate in which reactants are made.



Le Chatielers Principle

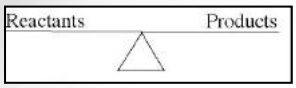
- Explains how a system in equilibrium will respond to STRESS
  - STRESS is any change in:
    - Concentration
    - Pressure
    - Temperature

- When a stress is applied, the system or reaction will shift in order to relieve that stress and reach a new equilibrium
- SHIFT: an increase in the RATE of EITHER the forward or reverse reaction.
- Shift to right (towards the products)

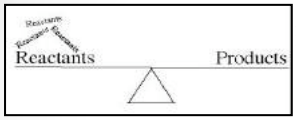
Reactants ⇀ Products

- Shift to left (towards the reverse reaction)

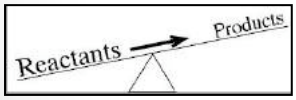
Reactants ↽ Products



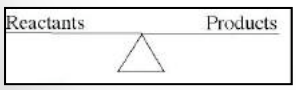
Equilibrium
Rate Forward = Rate Reverse



Add more reactant
Increase concentration of reactants



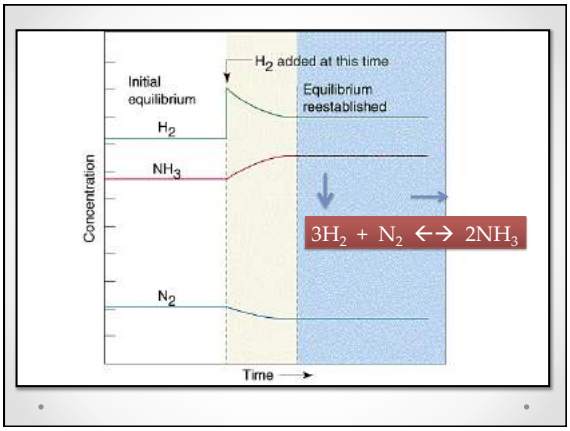
Produce more products
Rate of forward reaction increases
Shift left



Return to equilibrium
Rate of forward =rate of reverse

Concentration

- Concentration
  - When the concentration of a reactant or product is increased: the reactions shifts away from the increase
  - When the concentration of a reactant or product is decreased the reaction will shift toward the side that has experienced the decrease
    - Replaces what was taken!



Temperature

- Temperature
  - Involves increasing or decreasing "Heat"

NOTE: HEAT/ENERGY/J/KJ will either be a reactant or a product

A + B ⇆ C + D + HEAT

A + B + energy ⇆ C + D

- When temp is increased reaction shifts away from the side containing heat
- When temp is decreased the reaction shifts towards the side containing heat

Shift Right Or reverse

Add heat

$4NH_3(g) + 5O_2(g) \leftrightarrows 4NO(g) + 6H_2O(g) + HEAT$

If we added heat, which concentrations will decrease?

NO and H_2O

If we added heat, which concentrations will increase?

NH_3 and O_2

Pressure

- Pressure: Affects GASES ONLY!
  - Increase in Pressure
    - Reaction shifts to the side with the LEAST # of gas molecules (moles)
  - Decrease in Pressure
    - Reaction shifts to the side with the GREATEST # of gas molecules (moles)

★Note: if rxn contains no gas molecule or the rxn has same # of mols on reactant and product sides NO SHIFT★

$CO_{2\ (g)} \leftrightarrows CO_{2\ (aq)}$

If we increase the pressure, the concentration of which species will increase?

$CO_{2\ (aq)}$

If we decrease the pressure, the concentration of which species will increase?

$CO_{2\ (g)}$

Sketch Notes

Video 12.1: Collision Theory

1. Indicate whether each of the following would increase or decrease the rate of reaction.
    1. adding heat
    2. adding a catalyst
    3. diluting a solution
    4. removing an enzyme
    5. lowering the temperature
    6. decreasing the surface area
    7. increasing the concentration of a solution
    8. breaking a reactant down into smaller pieces

2. Identify which situation would have a higher reaction rate. Then state the factor that affected the rate of reaction in each situation.

	Situation X	**Situation Y**	**Higher reaction rate (X or Y)**	**Factor affecting the rate of reaction**
A	1 g of sugar cubes	1 g of sugar granules		
B	H_2O at 50°C	H_2O at 0°C		
C	1M $NaCl_{(aq)}$	5M $NaCl_{(aq)}$		
D	Protein synthesis w/ enzymes	Protein synthesis w/o enzymes		
E	3M $CO_{2(aq)}$ @ 1atm	3M $CO_{2(aq)}$ @ 2atm		

3. _____ Which of the following are true about how temperature affects the rate of reaction?
    I. heating causes the particles of the reactants to move quickly
    II. lowering the temperature will raise the energy of the particles
    III. increasing the temperature results in more collisions between the particles

    1. I and II only
    2. I and III only
    3. II and III only
    4. I, II and I

4. _____ Increasing which of the following will increase the frequency of collisions?
    I. temperature
    II. surface area
    III. concentration

    1. I and II only
    2. I and III only
    3. II and III only
    4. I, II and II

5. _____ Which of the following will lower the rate of reaction?
    1. adding an enzyme to the reaction
    2. decreasing the temperature from 40°C to 10°C
    3. breaking a chunk of calcium up into smaller pieces
    4. increasing the amount of solute dissolved in a solution

12.1 Collision Theory & Reaction Rate

1. As the temperature of a chemical reaction in the gas phase is increased, the rate of the reaction increases because

   1) fewer particle collisions occur
   2) more effective particle collisions occur
   3) the required activation energy increases
   4) the concentration of the reactants increases

2. A chemical reaction between iron atoms and oxygen molecules can only occur if

   1) the particles are heated
   2) the atmospheric pressure decreases
   3) there is a catalyst present
   4) there are effective collisions between the particle

3. Why can an increase in temperature lead to more effective collisions between reactant particles and an increase in the rate of a chemical reaction?

   1) The activation energy of the reaction increases.
   2) The activation energy of the reaction decreases.
   3) The number of molecules with sufficient energy to react increases.
   4) The number of molecules with sufficient energy to react decreases.

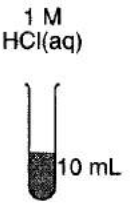
4. A reaction is most likely to occur when reactant particles collide with

   1) proper energy, only
   2) proper orientation, only
   3) both proper energy and proper orientation
   4) neither proper energy nor proper orientation

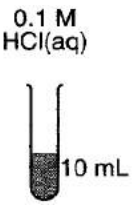
5. Increasing the temperature increases the rate of a reaction by

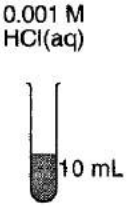
   1) lowering the activation energy
   2) increasing the activation energy
   3) lowering the frequency of effective collisions between reacting molecules
   4) increasing the frequency of effective collisions between reacting molecules

6. Each of four test tubes contains a different concentration of HCl(aq) at 25°C. A 1-gram cube of Zn is added to each test tube. In which test tube is the reaction occurring at the fastest rate?

   1) 

   3) 

   2) 

   4) 

7. A 5.0-gram sample of zinc and a 50.-milliliter sample of hydrochloric acid are used in a chemical reaction. Which combination of these samples has the fastest reaction rate?

   1) a zinc strip and 1.0 M HCl(aq)
   2) a zinc strip and 3.0 M HCl(aq)
   3) zinc powder and 1.0 M HCl(aq)
   4) zinc powder and 3.0 M HCl(aq)

8. For a given chemical reaction, the addition of a catalyst provides a different reaction pathway that

   1) decreases the reaction rate and has a higher activation energy
   2) decreases the reaction rate and has a lower activation energy
   3) increases the reaction rate and has a higher activation energy
   4) increases the reaction rate and has a lower activation energy

___ 9. Given the balanced equation representing a reaction:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

This reaction occurs more quickly when powdered iron is used instead of a single piece of iron of the same mass because the powdered iron

1) acts as a better catalyst than the single piece of iron
2) absorbs less energy than the single piece of iron
3) has a greater surface area than the single piece of iron
4) is more metallic than the single piece of iron

___ 10. The activation energy of a chemical reaction can be *decreased* by the addition of

1) a catalyst
2) an indicator
3) electrical energy
4) thermal energy

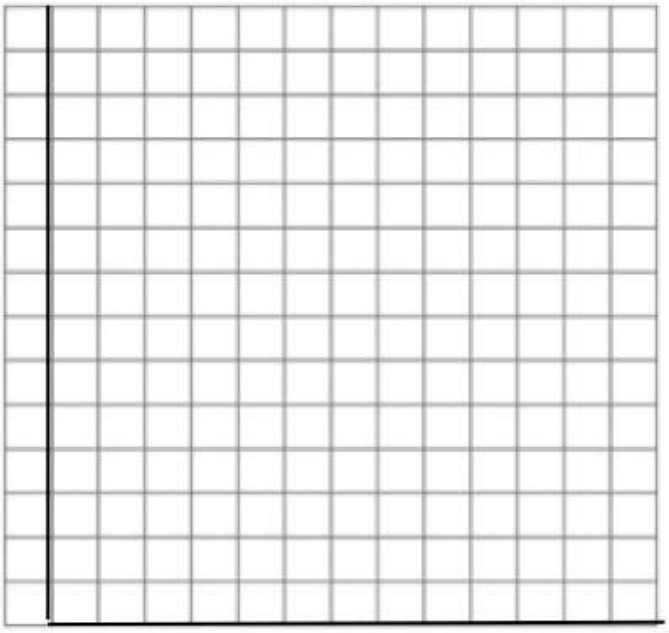
Video 12.2: Potential Energy Diagrams

Potential Energy Diagram #1

1. Construct a potential energy diagram for the reversible reaction below, give the following information:

$$A + B + \text{Energy} \leftrightarrows C + D$$

   1. Potential Energy of Reactants = 20 kJ
   2. Potential Energy of the Activated Complex = 90 kJ
   3. Potential Energy of the Products = 60 kJ



2. Label the Reactants and Products on the diagram
3. Label the following using the numbers listed:
   1. Potential Energy of Reactants
   2. Potential Energy of the Products
   3. Potential Energy of the Activated Complex
   4. Heat of Reaction (H) for the forward reaction Value = _____
   5. Activation energy for the forward reaction Value = _____
   6. Activation Energy of the Reverse reaction Value = _____

4. Is this reaction an Endothermic or Exothermic Reaction? How do you know?

5. Use a dotted line to show the addition of a Catalyst affects the potential energy diagram above.

1. How does the addition of a catalyst affect the following energy values (decreases, increase or remains the same)?

   1. Activation Energy of the reverse reaction:
   2. Potential Energy of the products:
   3. Potential Energy of the activated complex:

2. How do the following factors affect the rate of a chemical reaction (decreases, increase or remains the same)?

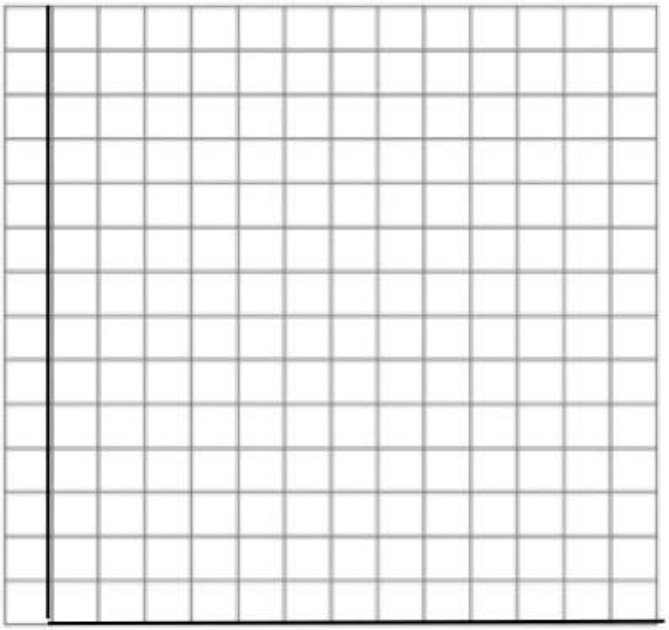
   1. Temperature:
   2. Concentration:
   3. Surface Area:

Potential Energy Diagram #2

6. Construct a potential energy diagram for the reversible reaction below, give the following information:

$$A + B \leftrightarrow C + D + \text{energy}$$

1. Potential Energy of Reactants = 50 kJ
2. Potential Energy of the Activated Complex = 80 kJ
3. Potential Energy of the Products = 30 kJ



7. Label the Reactants and Products on the diagram
8. Label the following using the numbers listed:
   1. Potential Energy of Reactants
   2. Potential Energy of the Products
   3. Potential Energy of the Activated Complex
   4. Heat of Reaction (H) for the forward reaction Value = _____
   5. Activation energy for the forward reaction Value = _____
   6. Activation Energy of the Reverse reaction Value = _____

9. Is this reaction an Endothermic or Exothermic Reaction? How do you know?

10. Use a dotted line to show the addition of a Catalyst affects the potential energy diagram above.

1. How does the addition of a catalyst affect the following energy values (decreases, increase or remains the same)?

1. Activation Energy of the reverse reaction:

2. Potential Energy of the products:

3. Potential Energy of the activated complex:

2. How do the following factors affect the rate of a chemical reaction (decreases, increase or remains the same)?

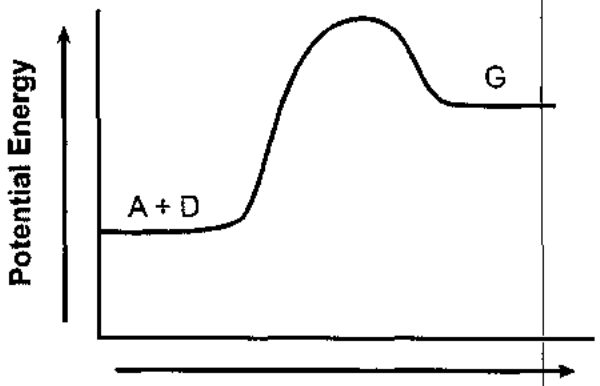
1. Temperature:

2. Concentration:

3. Surface Area:

Name: **Potential Energy Diagram Practice**

1. Given the potential energy diagram and equation representing the reaction between substances *A* and *D*:



$A + D \longrightarrow G$

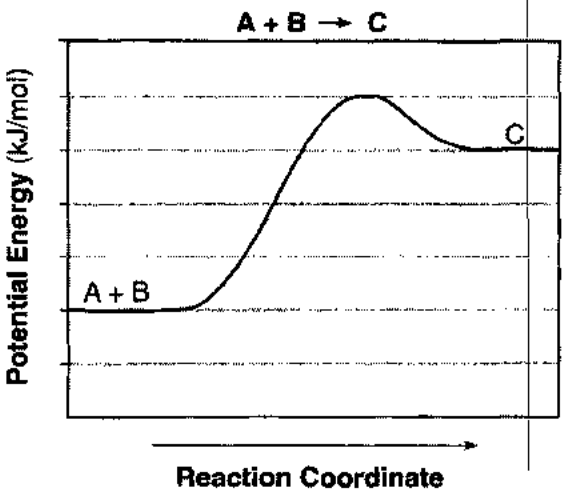
According to Table *I*, substance *G* could be

A) HI(g) B) $H_2O(g)$
C) $CO_2(g)$ D) $C_2H_6(g)$

2. In a chemical reaction, the difference between the potential energy of the products and the potential energy of the reactants is equal to the

A) activation energy
B) entropy of the system
C) heat of fusion
D) heat of reaction

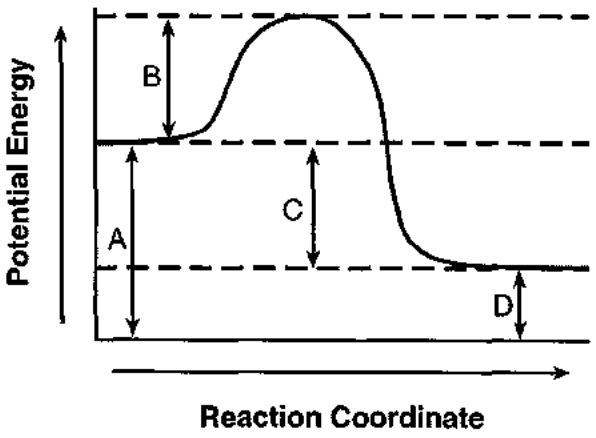
3. Given the equation and potential energy diagram representing a reaction:



If each interval on the axis labeled "Potential Energy (kJ/mol)" represents 10. kJ/mol, what is the heat of reaction?

A) +60. kJ/mol B) +20. kJ/mol
C) +30. kJ/mol D) +40. kJ/mol

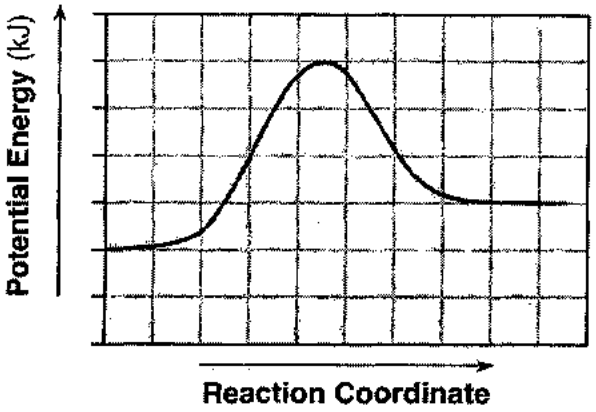
4. Given the potential energy diagram representing a reversible reaction:



The activation energy for the reverse reaction is represented by

A) $A + B$ B) $B + C$ C) $B + D$ D) $C + D$

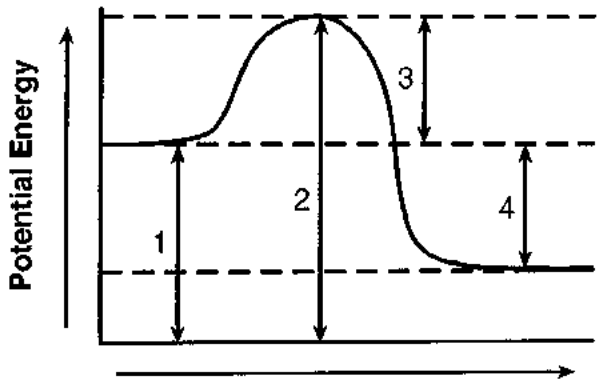
5. The potential energy diagram for a chemical reaction is shown below.



Each interval on the axis labeled "Potential Energy (kJ)" represents 40 kilojoules. What is the heat of reaction?

A) −120kJ B) −40kJ
C) +40kJ D) +160kJ

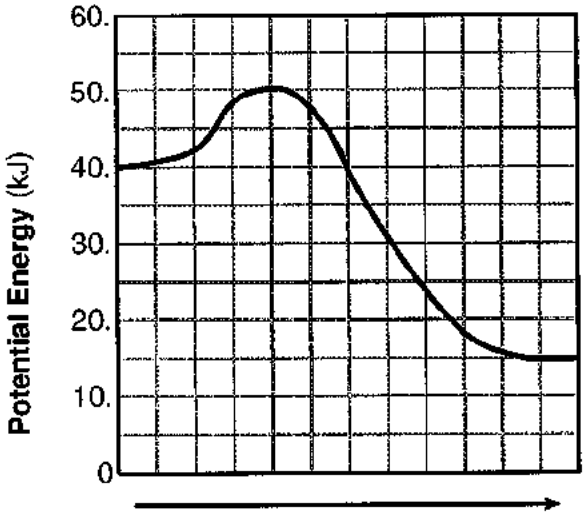
6. Given the potential energy diagram for a reaction:



Which interval on this diagram represents the difference between the potential energy of the products and the potential energy of the reactants?

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

7. Given the potential energy diagram for a chemical reaction:



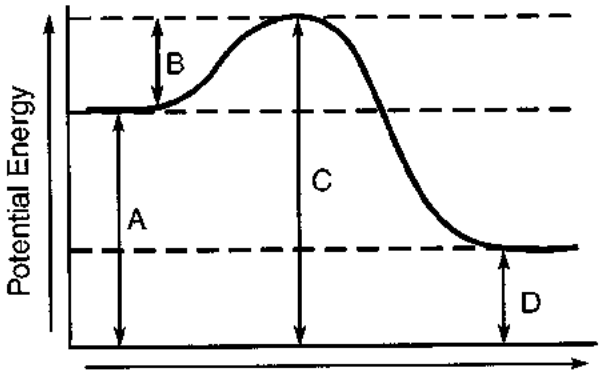
Which statement correctly describes the energy changes that occur in the forward reaction?

A) The activation energy is 10. kJ and the reaction is endothermic.
B) The activation energy is 10. kJ and the reaction is exothermic.
C) The activation energy is 50. kJ and the reaction is endothermic.
D) The activation energy is 50. kJ and the reaction is exothermic.

8. Which information about a chemical reaction is provided by a potential energy diagram?

A) the oxidation states of the reactants and products
B) the average kinetic energy of the reactants and products
C) the change in solubility of the reacting substances
D) the energy released or absorbed during the reaction

9. The potential energy diagram below represents a reaction.



Which arrow represents the activation energy of the forward reaction?

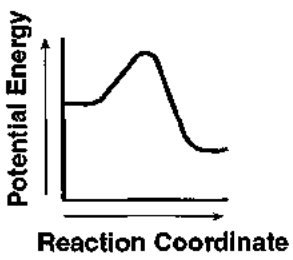
A) *A* B) *B* C) *C* D) *D*

10. Given the reaction:

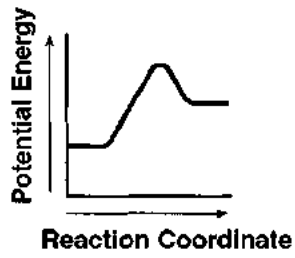
$S(s) + O_2(g) \rightarrow SO_2(g) + \text{energy}$

Which diagram best represents the potential energy changes for this reaction?

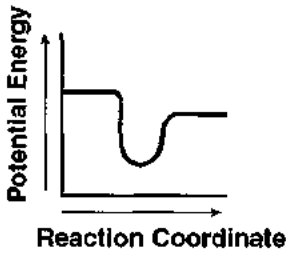
A)



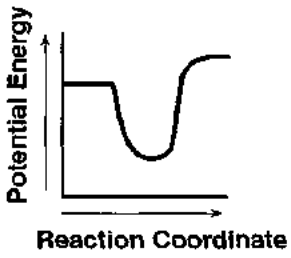
B)



C)

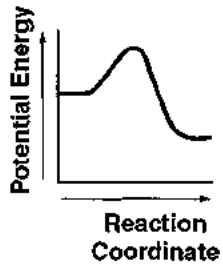


D)

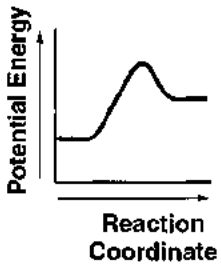


11. According to Table *I*, which potential energy diagram best represents the reaction that forms $H_2O(\ell)$ from its elements?

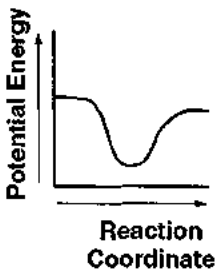
A)



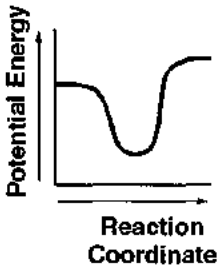
B)



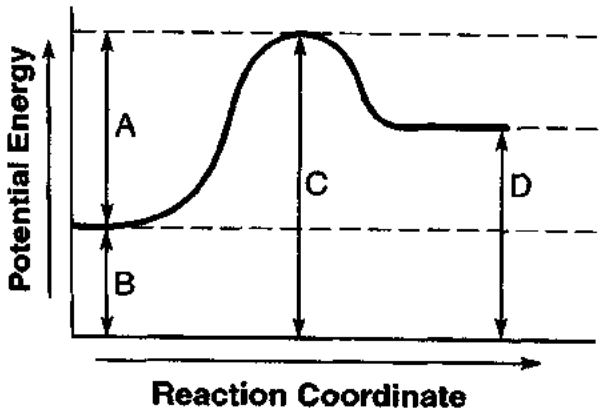
C)



D)



12. Given the potential energy diagram of a chemical reaction:



Which arrow represents the potential energy of the reactants?

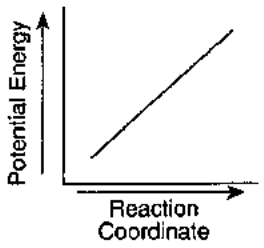
A) *A* B) *B* C) *C* D) *D*

13. The activation energy required for a chemical reaction can be *decreased* by

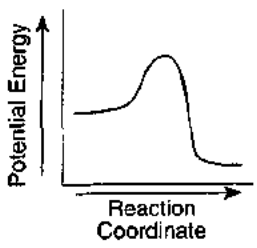
A) increasing the surface area of the reactant
B) increasing the temperature of the reactant
C) adding a catalyst to the reaction
D) adding more reactant

14. When a spark is applied to a mixture of hydrogen and oxygen, the gases react explosively. Which potential energy diagram best represents the reaction?

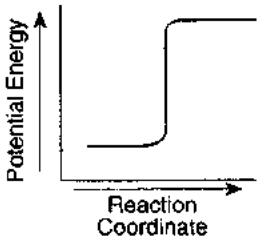
A)



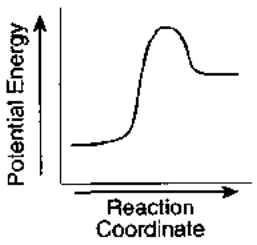
B)



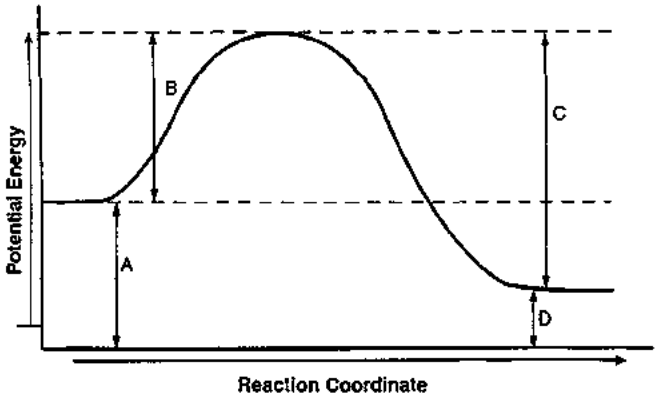
C)



D)



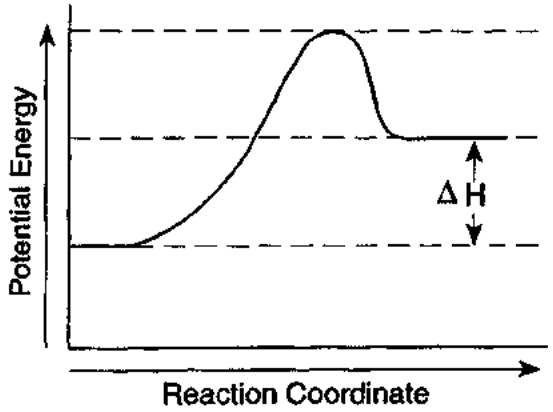
15. A potential energy diagram is shown below.



Which letters represent the activation energy of the forward and reverse reactions, respectively?

A) *A* and *C*
B) *A* and *D*
C) *B* and *C*
D) *B* and *D*

16. The diagram below represents the energy changes that occur during the formation of a certain compound under standard conditions.

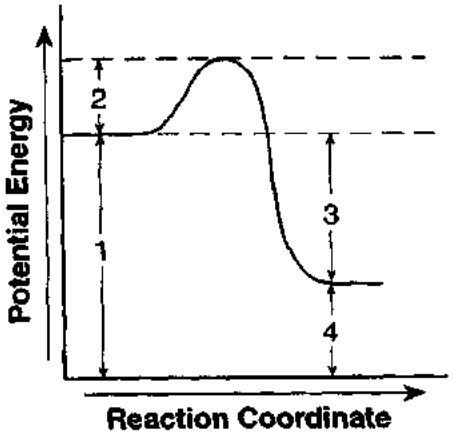


According to Reference Table I, the compound could be

A) $C_2H_6(g)$
B) $CO_2(g)$
C) $HI(g)$
D) $NH_3(g)$

Base your answers to questions **17** and **18** on the potential energy diagram below, which represents the reaction:

$A + B \rightarrow C$ + energy.



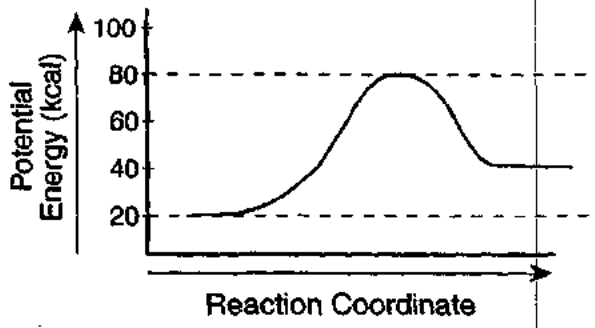
17. Which numbered interval will change with the addition of a catalyst to the system?

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

18. Which statement correctly describes this reaction?

A) It is endothermic and energy is absorbed.
B) It is endothermic and energy is released.
C) It is exothermic and energy is absorbed.
D) It is exothermic and energy is released.

19. A potential energy diagram of a chemical reaction is shown below.



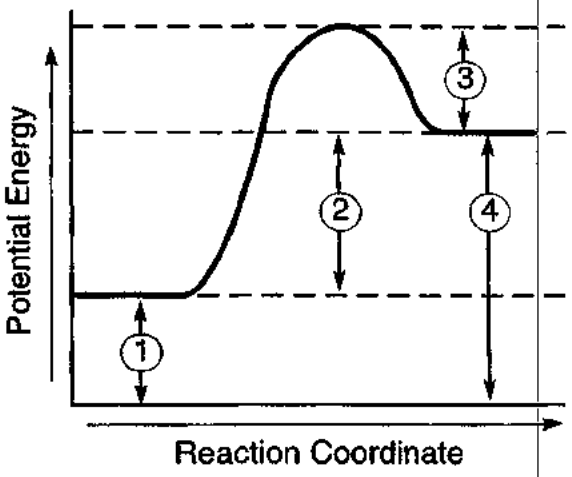
What is the difference between the potential energy of the reactants and the potential energy of the products?

A) 20. kcal B) 40. kcal
C) 60. kcal D) 80. kcal

20. Given the reaction:

$N_2(g) + 2\ O_2(g) \leftrightarrow 2\ NO_2(g)$
$\Delta H = +7.9$ kcal/mole

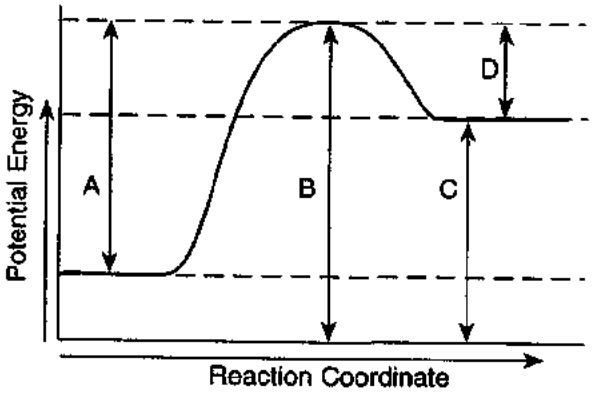
The potential energy diagram of the reaction is shown below.



Which arrow represents the heat of reaction (ΔH) for the reverse reaction?

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

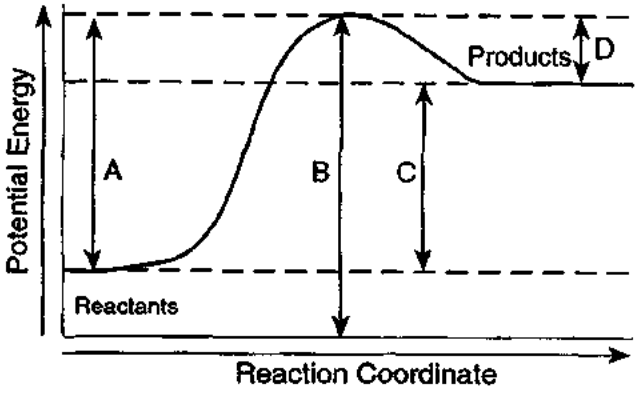
21. Base your answer to the following question on the potential energy diagram of a chemical reaction shown below.



The forward reaction is best described as an

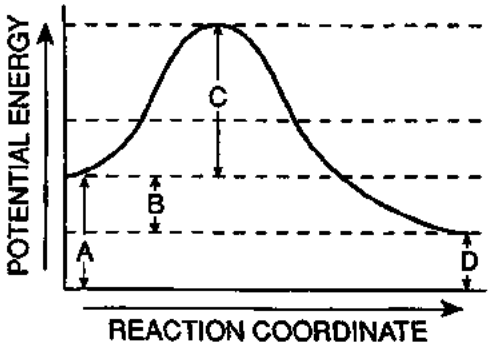
A) exothermic reaction in which energy is released
B) exothermic reaction in which energy is absorbed
C) endothermic reaction in which energy is released
D) endothermic reaction in which energy is absorbed

22. In the diagram below, which letter represents the activation energy for the reverse reaction?



A) *A* B) *B* C) *C* D) *D*

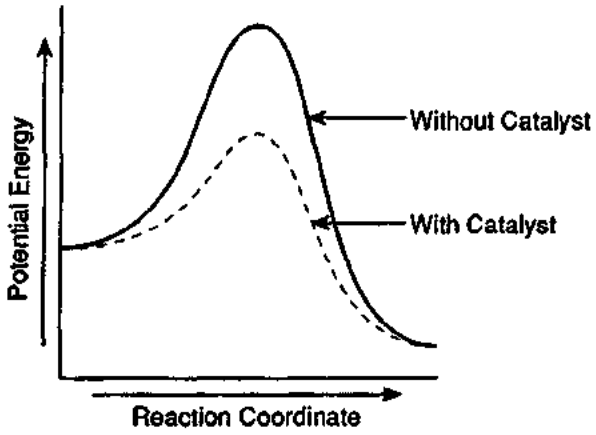
23. The potential energy diagram of a chemical reaction is shown below.



Which letter in the diagram represents the heat of reaction (ΔH)?

A) *A* B) *B* C) *C* D) *D*

24. A potential energy diagram is shown below.

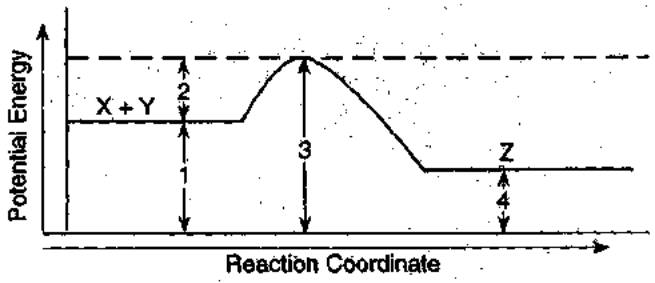


Which reaction would have the lowest activation energy?

A) the forward catalyzed reaction
B) the forward uncatalyzed reaction
C) the reverse catalyzed reaction
D) the reverse uncatalyzed reaction

25. The potential energy diagram below shows the reaction

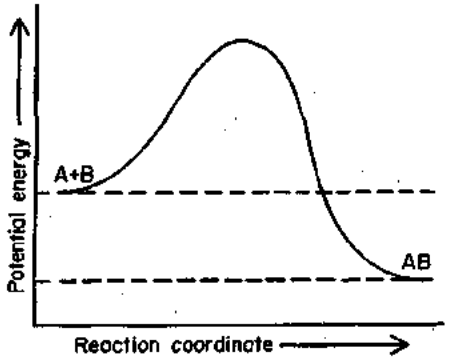
$X + Y \rightarrow Z$.



When a catalyst is added to the reaction, it will change the value of

A) 1 and 2 B) 1 and 3
C) 2 and 3 D) 3 and 4

26. Given the potential energy diagram:

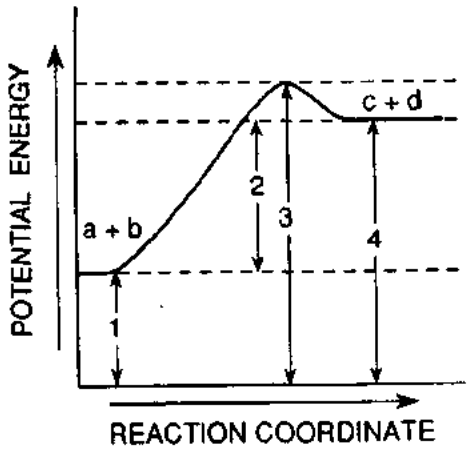


With reference to energy, the reaction $A + B \rightarrow AB$ can best be described as

A) endothermic, having a $+\Delta H$
B) endothermic, having a $-\Delta H$
C) exothermic, having a $+\Delta H$
D) exothermic, having a $-\Delta H$

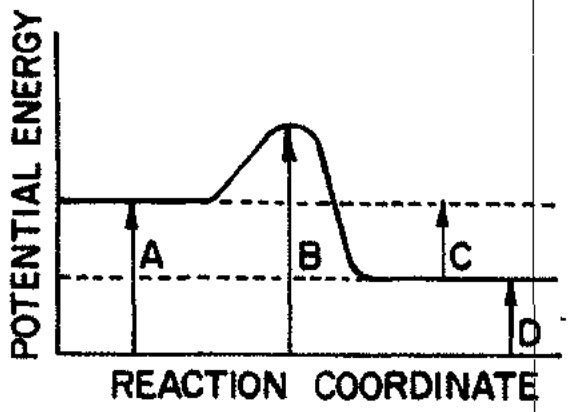
27. Which interval on the potential energy diagram shown below represents the ΔH of the reaction

$a + b \rightarrow c + d$?



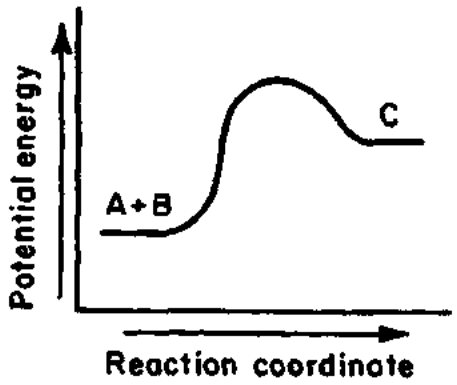
A) 1 B) 2 C) 3 D) 4

28. In the potential energy diagram below, which letter represents the potential energy of the activated complex?



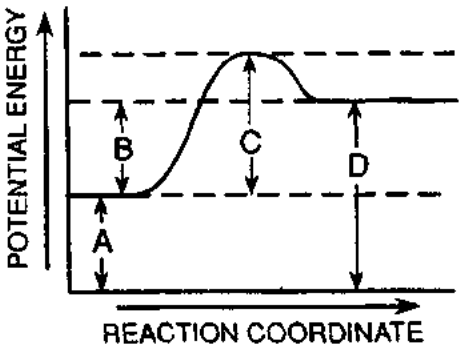
A) *A* B) *B* C) *C* D) *D*

29. According to the potential energy diagram below, what is the reaction $A + B \rightarrow C$?



A) endothermic and ΔH *is positive*
B) endothermic and ΔH *is negative*
C) exothermic and ΔH *is positive*
D) exothermic and ΔH *is negative*

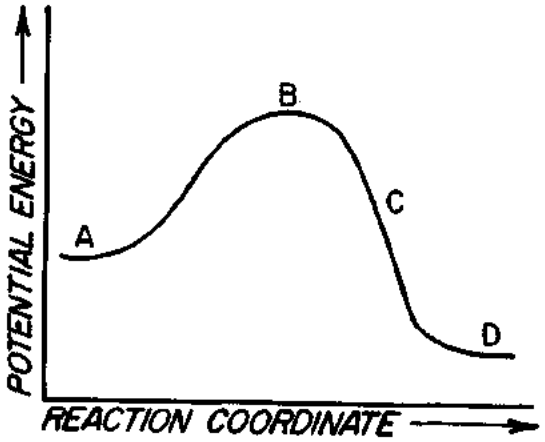
30. Base your answer to the following question on the reaction coordinate shown below:



Which interval represents the heat of reaction?

A) *A* B) *B* C) *C* D) *D*

31. The graph below represents the potential energy changes that occur in a chemical reaction. Which letter represents the activated complex?

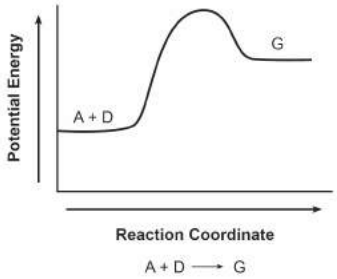


A) *A* B) *B* C) *C* D) *D*

Video 12.3: Enthalpy vs. Entropy
Enthalpy (ΔH) is also known as the heat of reaction.

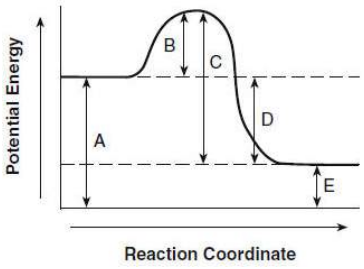
1. _____ According to Table *I*, which equation represents a change resulting in the greatest quantity of energy released?
   1. $2C_{(s)} + 3H_{2(g)} \rightarrow C_2H_{6(g)}$
   2. $2C_{(s)} + 2H_{2(g)} \rightarrow C_2H_{4(g)}$
   3. $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$
   4. $N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$

2. _____ Given the potential energy diagram and equation representing the reaction between substances *A* and *D*: According to Table I, substance G can be:
   1. $HI_{(g)}$
   2. $H_2O_{(g)}$
   3. $CO_{2(g)}$
   4. $C_2H_{6(g)}$



3. _____ In a chemical reaction, the difference between the potential energy of the products and the potential energy of the reactants is equal to the
   1. activation energy
   2. entropy of the system
   3. heat of fusion
   4. heat of reaction

Base your answers to question 4-5 on the P.E. diagram



4. _____ Which letter on represents the heat of reaction
   1. A
   2. B
   3. C
   4. D

5. This diagram represents an (endothermic/exothermic) reaction.

6. Answer the following questions using Table I and the reactions below. Determine the heat of reaction and state whether the reaction is endothermic or exothermic.
   1. $2NH_4NO_{3(s)} \rightarrow 2NH_4^+{}_{(aq)} + 2NO_3^-{}_{(aq)}$
   2. $2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$
   3. The synthesis of $H_2O_{(l)}$ from its elements.
   4. The decomposition of 1 mole of aluminum oxide.

7. _____ According to Table I, when 2.00 moles of $NaOH_{(s)}$ dissolves in water
   1. 44.5 kJ of energy is released and the temperature of the water increases
   2. 44.5 kJ of energy is absorbed and the temperature of the water decreases
   3. 89 kJ of energy is released and the temperature of the water increases
   4. 89 kJ of energy is absorbed and the temperature of the water decreases

Entropy (ΔS) is the degree of randomness in a substance. Determine whether the following reactions show on increase or decrease in entropy.

1. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ _____
2. $H_2O(l) \rightarrow H_2O(s)$ _____
3. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ _____
4. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ _____
5. $KCI(s) \rightarrow KCI(I)$ _____
6. $CO_2(s) \rightarrow CO_2(g)$ _____
7. $C(s) + O_2(g) \rightarrow CO_2(g)$ _____

8. _____ Given the balanced equation: **$I_2(s)$ + energy → $I_2(g)$**
   As a sample of $I_2(s)$ sublimes to $I_2(g)$, the entropy of the sample
   1. increases because the particles are less randomly arranged
   2. increases because the particles are more randomly arranged
   3. decreases because the particles are less randomly arranged
   4. decreases because the particles are more randomly arranged

9. _____ Which of these changes produces the greatest increase in entropy?
   1. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   2. $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$
   3. $H_2O(g) \rightarrow H_2O(l)$
   4. $CO_2(g) \rightarrow CO_2(s)$

10. _____ Which process is accompanied by a *decrease* in entropy?
    1. boiling of water
    2. condensing of water vapor
    3. subliming of iodine
    4. melting of ice

11. _____ Given the balanced equation:

$$KNO_3(s) + 34.89\ kJ \xrightarrow{H_2O} K^+(aq) + NO_3^-(aq)$$

Which statement best describes this process?

1. It is endothermic and entropy increases.
2. It is endothermic and entropy decreases.
3. It is exothermic and entropy increases.
4. It is exothermic and entropy decreases

12. _____ A thermometer is in a beaker of water. Which statement best explains why the thermometer reading initially increases when LiBr(s) is dissolved in the water?

1. The entropy of the LiBr(aq) is greater than the entropy of the water.
2. The entropy of the LiBr(aq) is less than the entropy of the water.
3. The dissolving of the LiBr(s) in water is an endothermic process.
4. The dissolving of the LiBr(s) in water is an exothermic process.

Spontaneous Reactions

Use Table I & your knowledge of enthalpy & entropy to determine if the following reactions are always, sometimes or never spontaneous.

Chemical Reactions	ΔH	ΔS	Spontaneous?
$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$			
$NaOH_{(s)} \rightarrow Na^{+1}_{(aq)} + OH^{-1}_{(aq)}$			
$3CO_{2(g)} + 4H_2O_{(l)} \rightarrow 5O_{2(g)} + C_3H_{8(g)}$			
$2NO_{2(g)} \rightarrow N_{2(g)} + 2O_{2(g)}$			
$Li^{+1}_{(aq)} + Br^{-1}_{(aq)} \rightarrow LiBr_{(s)}$			

Video 12.5. Le Chatelier's Principle

For each of the following, indicate the direction the equilibrium would shift **and** what would happen to the concentrations of each substance in equilibrium.

1. The following equilibrium maybe established with carbon dioxide and steam.

$$CO(g) + H_2O\ (g) <\rightarrow CO_2\ (g) + H_2\ (g) + \text{heat}$$

What would be the effect of each of the following on the equilibrium and concentrations?

	Shift	***Concentrations***
a) The addition of more H_2O?	_____	_____
b) The removal of some H_2?	_____	_____
c) Raising the temperature?	_____	_____
d) Increasing the pressure?	_____	_____
e) Addition of a catalyst?	_____	_____

2. What would be the effect of each of the following on the equilibrium involving the synthesis of methanol?

$$CO(g) + 2H_2\ (g) <\rightarrow CH_3OH\ (g)$$

	Shift	***Concentration***
a. The removal of CH_3OH?	_____	_____
b. An increase in pressure	_____	_____
c. Lowering the concentration of H_2?	_____	_____
d. The addition of a catalyst?	_____	_____

3. A small percentage of nitrogen gas and oxygen gas in the air combine at high temperatures found in automobile engines to produce NO(g), which is an air pollutant.

$$N2(g) + O2(g) + heat <\rightarrow 2NO(g)$$

   a. Higher engine temperatures are used to minimize carbon monoxide production. What effect does higher engine temperatures have on the production of NO(g)? Why?

   b. What effect would high pressures have on the production of NO(g)? Why?

12. 5 Le Chatelier' s Principle

1. Given the equation representing a reaction at equilibrium:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + heat$$

Which change causes the equilibrium to shift to the right?

A) adding a catalyst
B) adding more $O_2(g)$
C) decreasing the pressure
D) increasing the temperature

2. Given the equation representing a reaction at equilibrium:

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

What occurs when the concentration of $H_2(g)$ is increased?

A) The equilibrium shifts to the left, and the concentration of $N_2(g)$ decreases.
B) The equilibrium shifts to the left, and the concentration of $N_2(g)$ increases.
C) The equilibrium shifts to the right, and the concentration of $N_2(g)$ decreases.
D) The equilibrium shifts to the right, and the concentration of $N_2(g)$ increases

3. Given the balanced equation representing a reaction:

$2HCl(aq) + Na_2S_2O_3(aq) \rightarrow S(s) + H_2SO_3(aq) + 2NaCl(aq)$
Decreasing the concentration of $Na_2S_2O_3(aq)$ decreases the rate of reaction because the

A) activation energy decreases
B) activation energy increases
C) frequency of effective collisions decreases
D) frequency of effective collisions increases

4. Given the equation representing a reaction at equilibrium:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) + energy$$

Which change causes the equilibrium to shift to the right?

A) decreasing the concentration of $H_2(g)$
B) decreasing the pressure
C) increasing the concentration of $N_2(g)$
D) increasing the temperature

5. Given the system at equilibrium:
$2\ POCl_3(g) + energy \rightleftharpoons 2\ PCl_3(g) + O_2(g)$
Which changes occur when $O_2(g)$ is added to this system?

A) The equilibrium shifts to the right and the concentration of $PCl_3(g)$ increases.
B) The equilibrium shifts to the right and the concentration of $PCl_3(g)$ decreases.
C) The equilibrium shifts to the left and the concentration of $PCl_3(g)$ increases.
D) The equilibrium shifts to the left and the concentration of $PCl_3(g)$ decreases.

6. Given the reaction at equilibrium:

$A(g) + B(g) \leftrightarrow C(g) + D(g)$
The addition of a catalyst will

A) shift the equilibrium to the right
B) shift the equilibrium to the left
C) increase the rate of forward and reverse reactions equally
D) have no effect on the forward or reverse reactions

7. Given the reaction at equilibrium:

$2\ A(g) + 3\ B(g) \leftrightarrow A_2B_3(g) + \text{heat}$

Which change will not affect the equilibrium concentrations of $A(g)$, $B(g)$, and $A_2B_3(g)$?

A) adding more $A(g)$
B) adding a catalyst
C) increasing the temperature
D) increasing the pressure

8. The addition of a catalyst to a system at equilibrium will increase the rate of

A) the forward reaction, only
B) the reverse reaction, only
C) both the forward and reverse reactions
D) neither the forward nor reverse reaction

9. Given the equation representing a system at equilibrium:

$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + \text{energy}$

Which changes occur when the temperature of this system is *decreased*?

A) The concentration of $H_2(g)$ increases and the concentration of $N_2(g)$ increases.
B) The concentration of $H_2(g)$ decreases and the concentration of $N_2(g)$ increases.
C) The concentration of $H_2(g)$ decreases and the concentration of $NH_3(g)$ decreases.
D) The concentration of $H_2(g)$ decreases and the concentration of $NH_3(g)$ increases.

10. Given the equation representing a reaction at equilibrium:

$H_2(g) + I_2(g) + \text{heat} \leftrightarrow 2HI(g)$
Which change favors the reverse reaction?

A) decreasing the concentration of HI(g)
B) decreasing the temperature
C) increasing the concentration of $I_2(g)$
D) increasing the pressure

11. Given the system at equilibrium:

$N_2O_4(g) + 58.1\ \text{kJ} \leftrightarrow 2\ NO_2(g)$
What will be the result of an increase in temperature at constant pressure?

A) The equilibrium will shift to the left, and the concentration of $NO_2(g)$ will decrease.
B) The equilibrium will shift to the left, and the concentration of $NO_2(g)$ will increase.
C) The equilibrium will shift to the right, and the concentration of $NO_2(g)$ will decrease.
D) The equilibrium will shift to the right, and the concentration of $NO_2(g)$ will increase.

12. Given the equilibrium reaction in a closed system:

$H_2(g) + I_2(g) + \text{heat} \leftrightarrow 2\ HI(g)$
What will be the result of an increase in temperature?

A) The equilibrium will shift to the left and $[H_2]$ will increase.
B) The equilibrium will shift to the left and $[H_2]$ will decrease.
C) The equilibrium will shift to the right and [HI] will increase.
D) The equilibrium will shift to the right and [HI] will decrease.

13. Given the reaction at equilibrium:

$2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g) + \text{heat}$

Which change will shift the equilibrium to the right?

A) increasing the temperature
B) increasing the pressure
C) decreasing the amount of $SO_2(g)$
D) decreasing the amount of $O_2(g)$

14. Ammonia is produced commercially by the Haber reaction:

$$N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g) + \text{heat}$$

The formation of ammonia is favored by

A) an increase in pressure
B) a decrease in pressure
C) removal of $N_2(g)$
D) removal of $H_2(g)$

15. Given the system at equilibrium:

$$H_2(g) + F_2(g) \leftrightarrow 2\ HF(g) + \text{heat}$$

Which change will *not* shift the point of equilibrium?

A) changing the pressure
B) changing the temperature
C) changing the concentration of $H_2(g)$
D) changing the concentration of $HF(g)$

16. Given the reaction:

$$A(g) + B(g) \leftrightarrow AB(g)$$

As the pressure increases at a constant temperature, the rate of the forward reaction will

A) decrease B) increase
C) remain the same

17. Given the reaction at equilibrium:

$$2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g)$$

As the pressure is increased at constant temperature, the number of moles of $SO_3(g)$ produced will

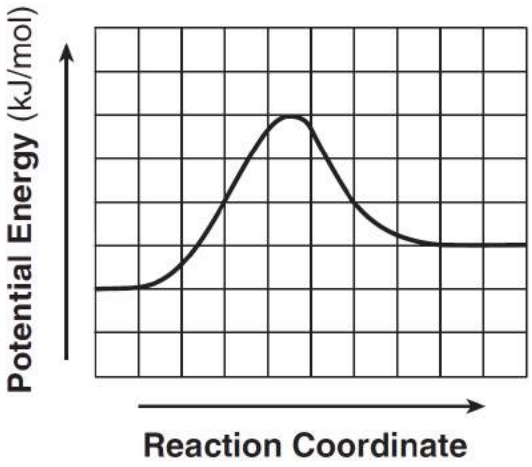
A) decrease B) increase
C) remain the same

18. Which system at equilibrium will be *least* affected by a change in pressure?

A) $3\ H_2(g) + N_2(g) \leftrightarrow 2\ NH_3(g)$
B) $2\ S(s) + 3\ O_2(g) \leftrightarrow 2\ SO_3(g)$
C) $AgCl(s) \leftrightarrow Ag^+(aq) + Cl^-(aq)$
D) $2\ HgO(s) \leftrightarrow 2\ Hg(\ell) + O_2(g)$

Name: **Kinetics & Equilibrium Review**

1. Given the potential energy diagram for a reversible chemical reaction:



Each interval on the axis labeled "Potential Energy (kJ/mol)" represents 10. kilojoules per mole. What is the activation energy of the forward reaction?

1) 10. kJ/mol
2) 30. kJ/mol
3) 40. kJ/mol
4) 60. kJ/mol

2. What is required for a chemical reaction to occur?

1) standard temperature and pressure
2) a catalyst added to the reaction system
3) effective collisions between reactant particles
4) an equal number of moles of reactants and products

3. Why can an increase in temperature lead to more effective collisions between reactant particles and an increase in the rate of a chemical reaction?

1) The activation energy of the reaction increases.
2) The activation energy of the reaction decreases.
3) The number of molecules with sufficient energy to react increases.
4) The number of molecules with sufficient energy to react decreases.

4. Increasing the temperature increases the rate of a reaction by

1) lowering the activation energy
2) increasing the activation energy
3) lowering the frequency of effective collisions between reacting molecules
4) increasing the frequency of effective collisions between reacting molecules

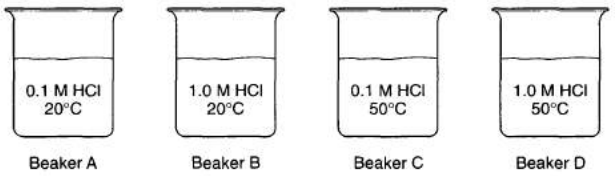
5. Given the reaction:

$$Mg + 2\ H_2O \rightarrow Mg(OH)_2 + H_2$$

At which temperature will the reaction occur at the greatest rate?

1) 25ºC
2) 50ºC
3) 75ºC
4) 100ºC

6. In each of the four beakers shown below, a 2.0-centimeter strip of magnesium ribbon reacts with 100 milliliters of HCl(aq) under the conditions shown.



In which beaker will the reaction occur at the fastest rate?

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

7. Given the reaction:

$$Zn(s) + 2\ HCl(aq) \rightarrow Zn^{2+}(aq) + 2\ Cl^{-}(aq) + H_2(g)$$

If the concentration of HCl(aq) is increased, the frequency of reacting collisions will

1) decrease, producing a decrease in the reaction rate
2) decrease, producing an increase in the reaction rate
3) increase, producing a decrease in the reaction rate
4) increase, producing an increase in the reaction rate

8. Given the balanced equation representing a reaction:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

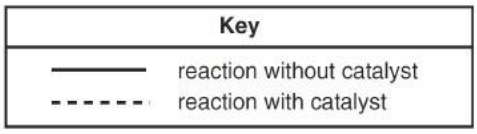
This reaction occurs more quickly when powdered iron is used instead of a single piece of iron of the same mass because the powdered iron

1) acts as a better catalyst than the single piece of iron
2) absorbs less energy than the single piece of iron
3) has a greater surface area than the single piece of iron
4) is more metallic than the single piece of iron

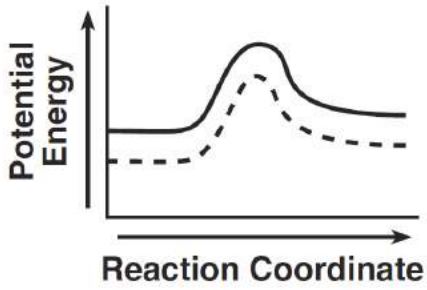
9. At STP, which 4.0-gram zinc sample will react fastest with dilute hydrochloric acid?

1) lump
2) bar
3) powdered
4) sheet metal

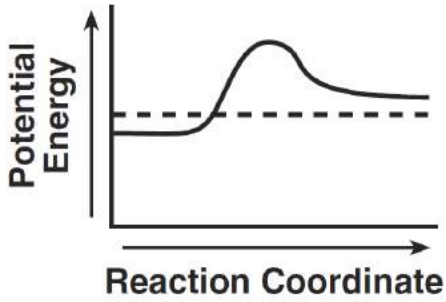
10. Which potential energy diagram represents the change in potential energy that occurs when a catalyst is added to a chemical reaction?



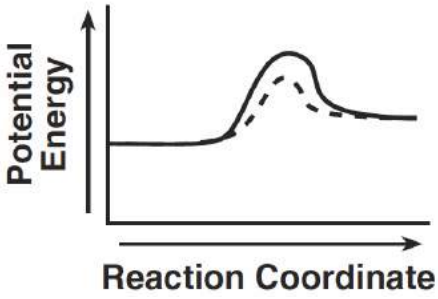
1)



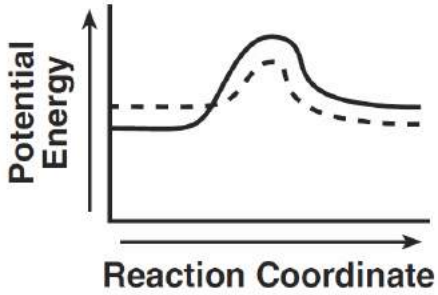
3)



2)



4)



11. The activation energy of a chemical reaction can be *decreased* by the addition of

1) a catalyst
2) an indicator
3) electrical energy
4) thermal energy

12. For a given reaction, adding a catalyst increases the rate of the reaction by

1) providing an alternate reaction pathway that has a higher activation energy
2) providing an alternate reaction pathway that has a lower activation energy
3) using the same reaction pathway and increasing the activation energy
4) using the same reaction pathway and decreasing the activation energy

13. A thermometer is in a beaker of water. Which statement best explains why the thermometer reading initially increases when LiBr(s) is dissolved in the water?

1) The entropy of the LiBr(aq) is greater than the entropy of the water.
2) The entropy of the LiBr(aq) is less than the entropy of the water.
3) The dissolving of the LiBr(s) in water is an endothermic process.
4) The dissolving of the LiBr(s) in water is an exothermic process.

14. Which balanced equation represents an endothermic reaction?

1) $C(s) + O_2(g) \rightarrow CO_2(g)$
2) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$
3) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
4) $N_2(g) + O_2(g) \rightarrow 2NO(g)$

15. For a chemical reaction, the difference between the potential energy of the products and the potential energy of the reactants is equal to the

1) heat of fusion
2) heat of reaction
3) activation energy of the forward reaction
4) activation energy of the reverse reaction

16. Given the balanced equation representing a reaction at 101.3 kPa and 298 K:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) + 91.8$ kJ

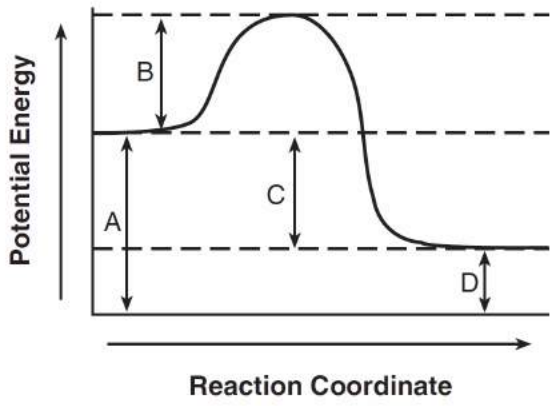
Which statement is true about this reaction?

1) It is exothermic and ΔH *equals –91.8 kJ.*
2) It is exothermic and ΔH *equals +91.8 kJ.*
3) It is endothermic and ΔH *equals –91.8 kJ.*
4) It is endothermic and ΔH *equals +91.8 kJ.*

17. According to Table *I*, which equation representsa change resulting in the greatest quantity of energy released?

1) $2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$
2) $2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$
3) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
4) $N_2(g) + O_2(g) \rightarrow 2NO(g)$

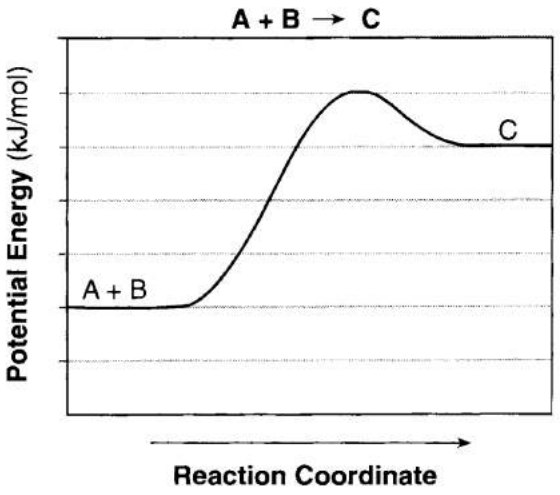
18. Given the potential energy diagram representing a reversible reaction:



The activation energy for the reverse reaction is represented by

1) *A + B*
2) *B + C*
3) *B + D*
4) *C + D*

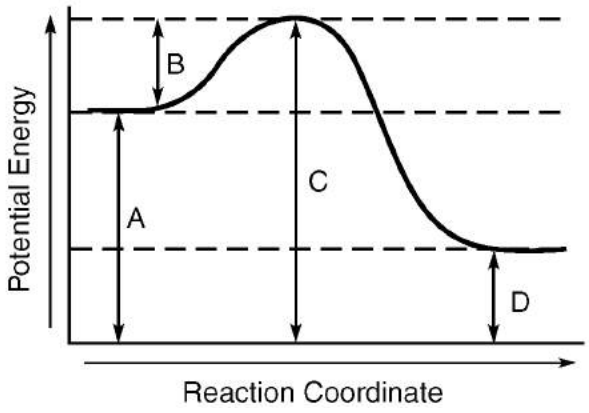
19. Given the equation and potential energy diagram representing a reaction:



If each interval on the axis labeled "Potential Energy (kJ/mol)" represents 10. kJ/mol, what is the heat of reaction?

1) +60. kJ/mol
2) +20. kJ/mol
3) +30. kJ/mol
4) +40. kJ/mol

20. The potential energy diagram below represents a reaction.



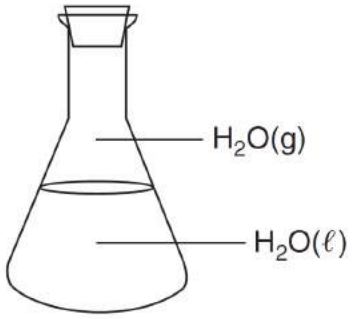
Which arrow represents the activation energy of the forward reaction?

1) *A*
2) *B*
3) *C*
4) *D*

21. Which statement describes a chemical reaction at equilibrium?

1) The products are completely consumed in the reaction.
2) The reactants are completely consumed in the reaction.
3) The concentrations of the products and reactants are equal.
4) The concentrations of the products and reactants are constant.

22. Given the diagram representing a closed system at constant temperature:



Stoppered Flask

Which statement describes this system at equilibrium?

1) The mass of $H_2O(\ell)$ equals the mass of $H_2O(g)$.
2) The volume of $H_2O(\ell)$ equals the volume of $H_2O(g)$.
3) The number of moles of $H_2O(\ell)$ equals the number of moles of $H_2O(g)$.
4) The rate of evaporation of $H_2O(\ell)$ equals the rate of condensation of $H_2O(g)$.

23. Which two factors must be equal when a chemical reaction reaches equilibrium?

1) the concentration of the reactants and the concentration of the products
2) the number of reactant particles and the number of product particles
3) the rate of the forward reaction and the rate of the reverse reaction
4) the mass of the reactants and the mass of the products

24. Some solid KNO_3 remains at the bottom of a stoppered flask containing a saturated $KNO_3(aq)$ solution at 22°C. Which statement explains why the contents of the flask are at equilibrium?

1) The rate of dissolving is equal to the rate of crystallization.
2) The rate of dissolving is greater than the rate of crystallization.
3) The concentration of the solid is equal to the concentration of the solution.
4) The concentration of the solid is greater than the concentration of the solution.

25. Given the equation representing a system at equilibrium:

H_2O (s) $\leftrightarrow$ H_20 (l)

At which temperature does this equilibrium exist at 101.3 kilopascals?

1) 0 K
2) 0°C
3) 32 K
4) 273°C

26. Given the reaction at equilibrium:

$$2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g) + \text{heat}$$

Which change will shift the equilibrium to the right?

1) increasing the temperature
2) increasing the pressure
3) decreasing the amount of $SO_2(g)$
4) decreasing the amount of $O_2(g)$

27. Given the reaction at equilibrium:

$$2\ A(g) + 3\ B(g) \leftrightarrow A_2B_3(g) + \text{heat}$$

Which change will not affect the equilibrium concentrations of $A(g)$, $B(g)$, and $A_2B_3(g)$?

1) adding more $A(g)$
2) adding a catalyst
3) increasing the temperature
4) increasing the pressure

28. Given the equation representing a reaction at equilibrium:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + \text{heat}$$

Which change causes the equilibrium to shift to the right?

1) adding a catalyst
2) adding more $O_2(g)$
3) decreasing the pressure
4) increasing the temperature

29. Given the equation representing a system at equilibrium:

$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + \text{energy}$

Which changes occur when the temperature of this system is *decreased*?

1) The concentration of $H_2(g)$ increases and the concentration of $N_2(g)$ increases.
2) The concentration of $H_2(g)$ decreases and the concentration of $N_2(g)$ increases.
3) The concentration of $H_2(g)$ decreases and the concentration of $NH_3(g)$ decreases.
4) The concentration of $H_2(g)$ decreases and the concentration of $NH_3(g)$ increases.

30. In terms of entropy and energy, systems in nature tend to undergo changes toward

1) lower entropy and lower energy
2) lower entropy and higher energy
3) higher entropy and lower energy
4) higher entropy and higher energy

31. Which equation represents a change that results in an increase in disorder?

1) $I_2(s) \rightarrow I_2(g)$
2) $CO_2(g) \rightarrow CO_2(s)$
3) $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
4) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

Base your answers to questions **32** through **34** on the information below and on your knowledge of chemistry.

Common household bleach is an aqueous solution containing hypochlorite ions. A closed container of bleach is an equilibrium system represented by the equation below.

$$Cl_2(g) + 2OH^-(aq) \rightleftharpoons ClO^-(aq) + Cl^-(aq) + H_2O(\ell)$$

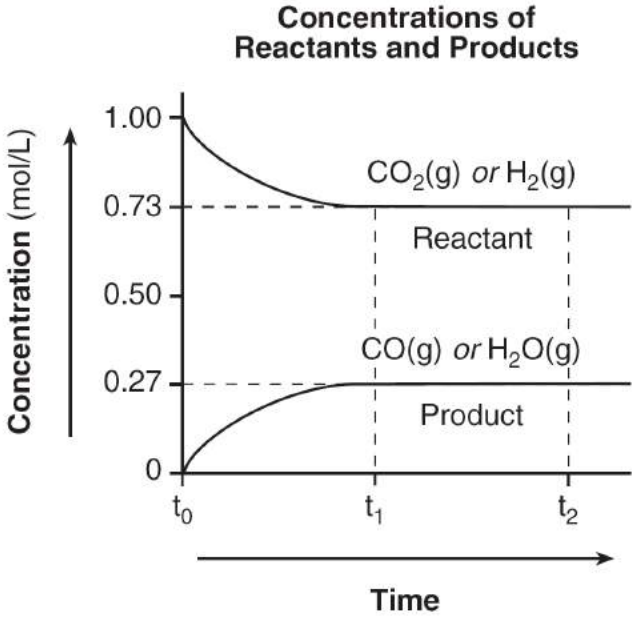
32. State the effect on the concentration of the ClO^- ion when there is a *decrease* in the concentration of the OH^- ion.

33. Explain why the container must be closed to maintain equilibrium.

34. Compare the rate of the forward reaction to the rate of the reverse reaction for this system.

Base your answers to questions **35** and **36** on the information below.

At 550°C, 1.00 mole of $CO_2(g)$ and 1.00 mole of $H_2(g)$ are placed in a 1.00-liter reaction vessel. The substances react to form CO(g) and $H_2O(g)$. Changes in the concentrations of the reactants and the concentrations of the products are shown in the graph below.

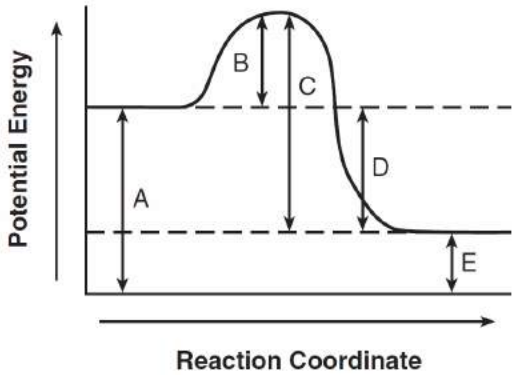


35. What can be concluded from the graph about the concentrations of the reactants and the concentrations of the products between time t_1 and time t_2?

36. Determine the change in the concentration of $CO_2(g)$ between time t_0 and time t_1.

Base your answers to questions **37** and **38** on the information below.

The chemical reaction between methane and oxygen is represented by the potential energy diagram and balanced equation below.



$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 890.4\ kJ$$

37. Explain, in terms of collision theory, why a lower concentration of oxygen gas *decreases* the rate of this reaction.

38. Which potential energy interval in the diagram represents the activation energy of the foward reaction?

Base your answers to questions **39** and **40** on the information below.

At room temperature, a reaction occurs when $KIO_3(aq)$ is mixed with $NaHSO_3(aq)$ that contains a small amount of starch. The colorless reaction mixture turns dark blue after a period of time that depends on the concentration of the reactants.

In a laboratory, 12 drops of a 0.02 M $NaHSO_3(aq)$ solution containing starch were placed in each of six test tubes. A different number of drops of 0.02 M $KIO_3(aq)$ and enough water to maintain a constant volume were added to each test tube and the time for the dark-blue color to appear was measured. The data were recorded in the table below.

Data Table

Test Tube	A	B	C	D	E	F
Number of Drops of 0.02 M KIO_3(aq)	2	4	6	8	10	12
Time for Dark-Blue Color to Appear (s)	210.	88	49	39	33	27

39. Identify *one* factor, other than the concentration of the reactants, that would affect the rate of this reaction.

40. State how increasing the number of drops of 0.02 M $KIO_3(aq)$ used in the reaction affects the rate of reaction.

Name:

Regents Chemistry

Practice Packet

Chapter 13: Acids & Bases

Chapter 13: Acids, Bases and Salts

Monoprotic acid - acids that contain only 1 hydrogen ion (H^+) HNO_3

Diprotic acid - acids that contain 2 hydrogen ion (H^+) H_2SO_4

Conjugate acid – the particle formed when a base gains a hydrogen ion

Conjugate base – the particle that remains when an acid has donated a hydrogen ion

Hydronium ion ($H_3O^{+)}$ – a water molecule that gains a hydrogen ion and becomes positivity charged

Arrhenius Acid – hydrogen-containing compounds that ionize or yield hydrogen ions (H^+)

Arrhenius Base – compounds that ionize hydroxide ions (OH^-)

Neutral Solution – an aqueous solution in which $[H^+] = [OH^-]$

Acidic Solution – a solution in which $[H^+]$ is greater than $[OH^-]$

Basic Solution - a solution in which $[H^+]$ is less than $[OH^-]$

pH – negative logarithm of the hydrogen-ion concentration $pH = -\log [H^+]$

Strong Acid - completely ionization (separate into ions) greater $[H^+]$

Strong Base - ionize slightly

Neutralization Reaction – a reaction in which an acid reacts with a base to produce salt and water (neutral)

Equivalence Point – when the number of moles of hydrogen ions equals the number of moles of hydroxide ions

Titration – process of adding known concentration to determine the concentration of another solution

End point – the point in which the indicator changes color in a titration

Characteristics of Acids & Bases

Chemistry 200
Video Lesson 13.1

Objective:

How can we recognize characteristics of acids and bases?

How can we determine the difference between Arrhenius and Bronsted/Lowry acids and bases?

Acids & Bases

Acids & Bases can be recognized by their properties.

Acids

- Dilute solutions have a sour taste
  ex: lemons --> citric acid
  vinegar --> acetic acid
  carbonated drinks --> carbonic acid
  tomato --> ascorbic acid

Bases

- Have a bitter taste & slippery, soapy feel.
  ex: soap, milk of magnesia

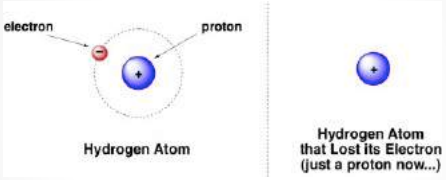
Acids

Two theories used to explain the behavior of acids:

Arrhenius Acid (Svante Arrhenius)

- a substance that releases hydrogen ions ***[H⁺]*** in an aqueous solution ex: HCl, H_2SO_4
- not all substances that contain hydrogen are acids. ***CH_4*** (methane) is ***not*** an acid because the hydrogen atoms do not ionize.

- A **Hydrogen Ion** is produced when a hydrogen atom loses an electron to become a positive ion



- A **Hydronium Ion $[H_3O]^+$** is formed when a hydrogen ion reacts w/ H_2O

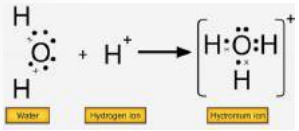


Table E

ex: $HCl_{(g)} \xrightarrow{H_2O} H_3O^+_{(aq)} + Cl^-_{(aq)}$

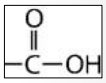
Arrhenius acids can be divided into 2 main categories

Inorganic Acid

- an acid that starts w/ hydrogen & does not have carbon
  ex: HCl, H_2SO_4

Organic Acid

- an acid that has carbon & ends w/ COOH
  (Carboxyl Group)



ex: CH_3COOH (ethanoic acid)
HCOOH (methanoic acid)

An Arrhenius acid can sometimes yield more than one hydrogen ion in aqueous solution

Monoprotic Acid

- acids that produce a single hydrogen ion (ionizes in 1 step)

ex: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Diprotic Acid

- acids that produce 2 hydrogen ions (ionizes in 2 steps)

ex: $H_2SO_4 \rightarrow H^+ + HSO_4^-$ (step 1)
$HSO_4^- \rightarrow H^+ + SO_4^{-2}$ (step 2)

Triprotic acids: 3 hydrogen ions ex: H_3PO_4

The **Brønsted-Lowry theory** focuses solely on the hydrogen ion as a proton & can be used to further explain acid behavior:

Brønsted-Lowry Acid

- any species (any particle: atom, molecule or ion) that can donate a proton ***[H⁺]*** to another species ***(a proton donor)***
- all Arrhenius acids are also Brønsted-Lowry acids

  ex: $HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

- according to Brønsted-Lowry, HCl is an acid because it ***donates*** a proton (H^+)

Bases

Theories used to explain acid behavior can also be used to explain bases as well

Arrhenius Base

- a substance that releases hydroxide ions ***[OH⁻]*** in an aqueous solution

ex: **NaOH** $\xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$

CH_3OH
C_2H_5OH

*An **Alcohol** should not be confused w/ an Arrhenius Base as it does not form a hydroxide ion [OH^-] in an (aq) solution, therefore it is **NOT** a base!!!*

Brønsted-Lowry Base

- any species that can accept a proton ***[H⁺]*** from another species --> ***a proton acceptor***
- according to Arrhenius, only metallic hydroxides are bases
- Brønsted-Lowry deUinitions explain why a substance that has no [OH^-], like NH_3, behaves like a base in H_2O

  ex: **NH_3** + $H_2O \rightarrow NH_4^+ + OH^-$

Amphiprotic (Amphoteric)

- a substance that can act as an acid or a base

ex: H_2O

$HCl + \boldsymbol{H_2O} \rightarrow H_3O^+ + Cl^-$

Base

$NH_3 + \boldsymbol{H_2O} \rightarrow NH_4^+ + OH^-$

Acid

Electrolytes

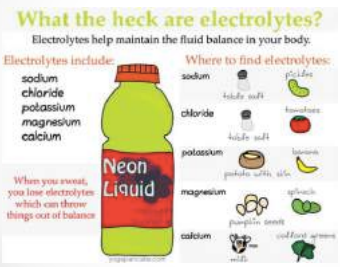
Chemistry 200
Video Lesson 13.2

Objective:

How do we determine which acids and bases are strong or weak electrolytes?

Electrolytes

- are substances that have mobile ions when put into solution (aq). This allows it to conduct electricity.



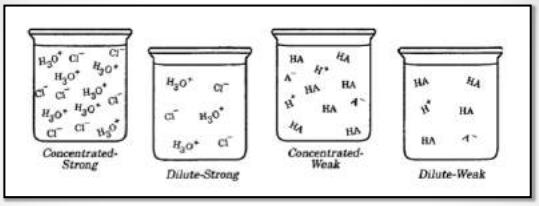
Ex: Acids
Bases
Salts

A ***Strong Acid*** is a strong electrolyte because it completely ionizes in water to produce a large number of H^+ ions in solution.

A ***Weak Acid*** is a weak electrolyte because it partially ionizes in water to produce a small number of H^+ ions in solution.

Table K
Common Acids

Formula	Name
HCl(aq)	hydrochloric acid
HNO_2(aq)	nitrous acid
HNO_3(aq)	nitric acid
H_2SO_3(aq)	sulfurous acid
H_2SO_4(aq)	sulfuric acid
H_3PO_4(aq)	phosphoric acid
H_2CO_3(aq) or CO_2(aq)	carbonic acid
CH_3COOH(aq) or $HC_2H_3O_2$(aq)	ethanoic acid (acetic acid)



A ***Strong Base*** is a strong electrolyte because it completely ionizes in water to produce a large number of OH^- ions in solution

A ***Weak Base*** is a weak electrolyte because it partially ionizes in water to produce a small number of OH^- ions in solution.

Table I. Common Bases

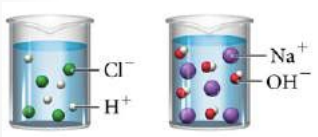
Formula	Name
NaOH(aq)	sodium hydroxide
KOH(aq)	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

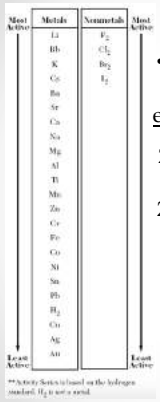
Acid/Base Reactions

- Acid react with a base to produce a salt and water

ex:

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow \boxed{NaCl_{(aq)}} + HOH_{(l)}$





Acid/Metal Reactions

- Acids react with a metal to produce a salt and hydrogen gas ($\mathbf{H_{2(g)}}$)

ex:

$2HNO_{3(aq)} + Ca_{(s)} \rightarrow Ca(NO_3)_{2(aq)} + H_{2(g)}$

$2HCl_{(aq)} + Cu_{(s)} \rightarrow CuCl_{2(aq)} + H_{2(g)}$

NO REACTION!!

Conjugate Acid-Base pairs

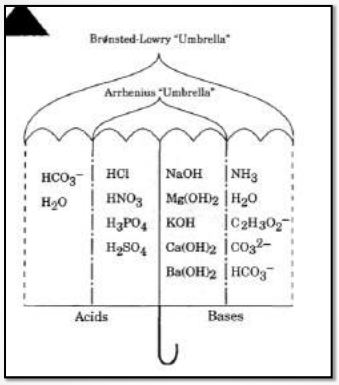
Video Lesson 13.3

Objectives

- Compare and contrast acids and bases as defined by the theories of Arrhenius and Bronsted-Lowry (alternate acid/base theory).

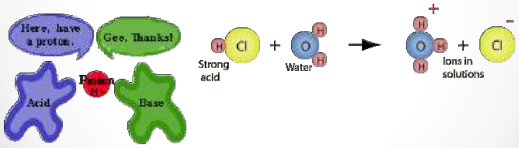
- QUESTION: Why is ammonia (NH_3) a base?

Acid Base Theories



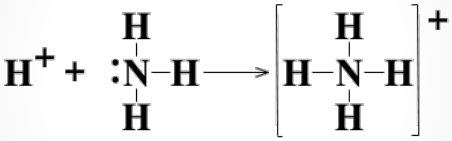
Alternate Acid-Base Theory
AKA Bronsted-Lowry Theory

- Explains the behavior of WEAK acids and BASES (Na_2CO_3 & NH_3)
- Bronsted-Lowry Acids are **PROTON DONORS**
  - The prefer to give away a proton (H^+)



Alternate Acid Base Theory

- Bronsted-Lowry Bases are PROTON ACCEPTORS

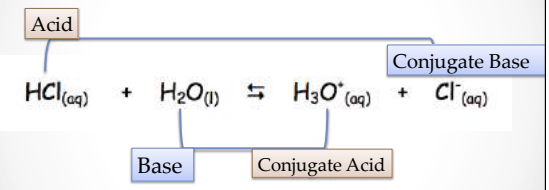


Conjugate Acid/Base Pairs

- An acid is a H^+ donor while a base is a H^+ acceptor.
- The substance produced when an acid has donated its proton is called the conjugate base.
- The substance produced when the base accepts a H^+ is called the conjugate acid

Practice

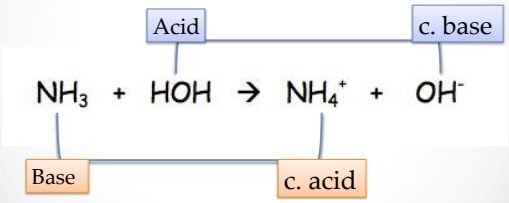
Label the acids and bases according to Bronsted-Lowry:



Conjugate Acid-Base Pairs

Practice

Label conjugate acid base pairs in the following equation



pH & Acid-Base Indicators

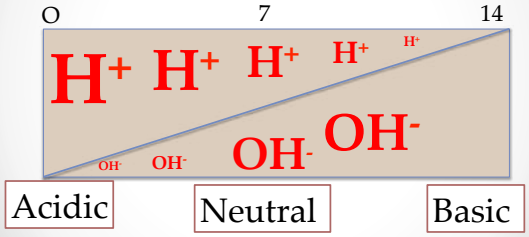
Video Lesson 13.4

Objectives

- Describe how [H^+] and [OH^-] are related in an aqueous solution.
- Classify a solution as neutral, acidic or basic.
- Describe the purpose of an acid-base indicator.

The pH Scale

- pH = direct measurement of H^+ ion concentration in a solution
- "power of hydrogen"



Logarithmic

- The pH scale is logarithmic
  - Based on exponents of the number 10
- The pH of a solution is the negative log of the [H^+] or [OH^-] ions

$pH = -\log[H^+]$

pH of .01 M HBr = $-\log(.01) = 2.00$

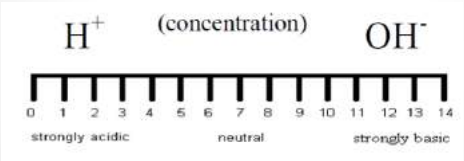
$pOH = -\log[OH^-]$

pH of .01 M LiOH = $14 - 2 = 12.00$

$pH = 14 - pOH$

A Tenfold Change

- Each change in a single pH unit signifies a tenfold change in [H^+] concentration



Ex: going from a ph of 4 to a pH of 5

Becoming _____ basic; [OH^-] _____ by _____

Becoming _____ acidic; [H^+] _____ by _____

Ex: going from a pH of 13 to a pH of 10

Becoming _____ acidic; [H^+] _____ by _____

Becoming _____ basic; [OH^-] _____ by _____

Acid-Base Indicators

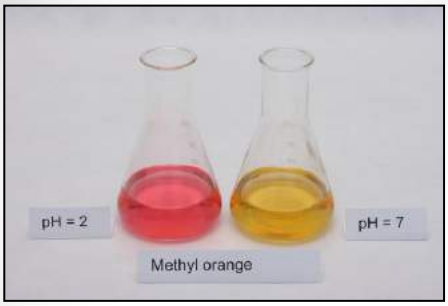
- Indicator
  - A substance that changes color as a result of a pH change
- Table M
- Example: Phenolphthalein → colorless up until a pH of 8, light pink from 8 to 9 and pink from pH of 9 up

pH = 5 pH = 11

Table M
Common Acid–Base Indicators

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.1–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8–9	colorless to pink
litmus	4.5–8.3	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

Source: *The Merck Index*, 14th ed., 2006, Merck Publishing Group



Indicator	pH of sample	Color indicator will turn
methyl orange	6.0	
bromothymol blue	2.0	
phenolphthalein	10	
Litmus	6.8	
bromocresol green	13	
thymol blue	1.2	

Acid-Base Titration

Video Lesson 13.5

Objectives

- Define products of an acid-base reaction.
- Explain how acid-base titration is used to calculate the concentration of an acid or base.
- Explain the concept of equivalence in neutralization reactions.

Neutralization Reactions

- Neutral
  - Neither acidic nor basic
  - Equal concentrations of H^+ **and** OH^-
  - Occurs when Arrhenius acid and an Arrhenius base react to form water and salt

Neutralization Reactions

Acid + Base → Salt + Water

Double Replacement!

Water

+1 -1 +1 -1 +1 -1 +1 -1

$HCl + KOH \rightarrow KCl + HOH$

Acid

Base

Salt
Or
Ionic Substance

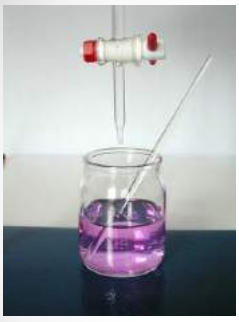
Magnesium hydroxide

- Stomach acid
  - pH between 2-3
- Heart burn
  - Happens when pH level drops in the stomach
- Milk of Magnesia brings the pH up
- $Mg(OH)_2$

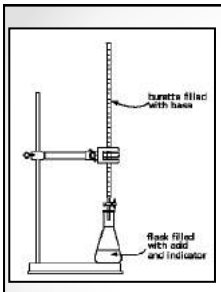
1. $HBr + KOH \rightarrow$
2. $NaOH + HC_2H_3O_2 \rightarrow$
3. $KOH + H_3PO_4 \rightarrow$
4. $2HNO_3 + Ca(OH)_2 \rightarrow$

Acid-Base Titration

- Titrations are used to calculate the concentration (Molarity) of an unknown solution.
- Lab process in which an unknown solution is systematically reacted with a solution of known concentration by adding measured volumes of an acid or a base to the unknown until neutralization occurs.



- In all neutralization reactions there must be a 1:1 ratio of moles of H^+ ions and moles of OH^- ions.
- Equivalence Point
  - $[H^+] = [OH^-]$



Titration formula (Table T): $M_A V_A = M_B V_B$

Where: M_A = molarity of acid (H^+)
V_A = volume of acid
M_B = molarity of base (OH^-)
V_B = volume of base

- Use this formula when you are dealing with a titration or neutralization word problem
- Make sure that all units are in agreement when plugging into formula (so they cancel out and you get the right answer!)

Example #1

- What is the concentration of a solution of HI if 0.3 L is neutralized by 0.6 L of 0.2 M solution of KOH?

Titration	$M_AV_A = M_BV_B$	M_A = molarity of H^+	M_B = molarity of OH^-
		V_A = volume of acid	V_B = volume of base

M_A = ?
V_A = 0.3 L
M_B = 0.2 M
V_B = 0.6 L

$M_A V_A = M_B V_B$
$M_A (0.3) = (0.2)(0.6)$
$M_A = \frac{(0.2)(0.6)}{(0.3)}$
M_A = 0.4M or moles/L

Example #2

- You have 50 mL of 1.0 M $H_2SO_{4\,(aq)}$. What volume of 0.5 M NaOH would be required to neutralize the acid? (Diprotic Acids yield 2 H^+ ions in solution!)

M_A = 2(1.0 M)
V_A = 50 mL
M_B = 0.5 M
V_B = ?

$2(M_A V_A) = M_B V_B$
$(2)(1.0)(50) = (0.5)V_B$
$V_B = \frac{(2)(1.0)(50)}{(0.5)}$
V_B = 200 mL

Sketch Notes

Video 13.1: Characteristics of Acids and Bases

Acid/Base/Salt Characteristics:

On the line to the left, write A if the statement is a property of an acid, write B if the property is that of a base, and write X, if it is a property of both acidic and basic solutions.

_____ 1) Often feels smooth and slippery

_____ 2) Has a sour taste

_____ 3) Stings in open wounds

_____ 4) Typically reacts vigorously with metals

_____ 5) Has a bitter taste

_____ 6) Turns litmus paper from blue to red

_____ 7) Is an electrolyte

_____ 8) Often looks like pure water

_____ 9) Turns litmus paper from red to blue

_____ 10) Typically does not react with metals

Complete the chart.

Substance	Acid, Base or Salt?	How do you know?
1. $NaOH$		
2. HCl		
3. $NaCl$		
4. HF		
5. K_2SO_4		
6. $Fe(OH)_3$		
7. H_3PO_4		

Answer the following questions

1. _____ Which species can conduct an electric current?
   1. $H_2O_{(s)}$
   2. $CH_3OH_{(aq)}$
   3. $NaOH_{(s)}$
   4. $HCl_{(aq)}$

2. _____ According to Arrhenius theory, which species does an acid produce in aqueous solution?
   1. hydroxide ions
   2. sodium ions
   3. hydrogen ions
   4. chloride ions

3. _____ Which substance is an Arrhenius acid?
   1. $Mg(OH)_{2(aq)}$
   2. $LiF_{(aq)}$
   3. $CH_3CHO_{(aq)}$
   4. $HBr_{(aq)}$

4. _____ According to the Arrhenius theory, when a base is dissolved in water, it produces a solution containing only one kind of negative ion. What is the name of this negative ion?
   1. hydrogen sulfate ion
   2. hydrogen carbonate ion
   3. hydride ion
   4. hydroxide ion

5. In terms of H^+ ions, explain the difference between monoprotic, diprotic and triprotic acids.

Arrhenius Acids and Bases

1. Which compounds are classified as Arrhenius acids?
   A) HCl and $NaOH$
   B) HNO_3 and $NaCl$
   C) NH_3 and H_2CO_3
   D) HBr and H_2SO_4

2. What can be explained by the Arrhenius theory?
   A) the behavior of many acids and bases
   B) the effect of stress on a phase equilibrium
   C) the operation of an electrochemical cell
   D) the spontaneous decay of some nuclei

3. Potassium hydroxide is classified as an Arrhenius base because KOH contains
   A) OH^- ions
   B) O^{2-} ions
   C) K^+ ions
   D) H^+ ions

4. When one compound dissolves in water, the only positive ion produced in the solution is H_3O^+ (aq). This compound is classified as
   A) a salt
   B) a hydrocarbon
   C) an Arrhenius acid
   D) an Arrhenius base

5. Given the equation:

   $HCl(g) + H_2O(\ell) \rightarrow X(aq) + Cl^-(aq)$
   Which ion is represented by *X*?
   A) hydroxide
   B) hydronium
   C) hypochlorite
   D) perchlorate

6. The only positive ion found in $H_2SO_4(aq)$ is the
   A) ammonium ion
   B) hydronium ion
   C) hydroxide ion
   D) sulfate ion

7. Which substance is an Arrhenius acid?
   A) $Ba(OH)_2$
   B) CH_3COOCH_3
   C) H_3PO_4
   D) $NaCl$

8. Which two formulas represent Arrhenius acids?
   A) CH_3COOH and CH_3CH_2OH
   B) $HC_2H_3O_2$ and H_3PO_4
   C) $KHCO_3$ and $KHSO_4$
   D) $NaSCN$ and Na_2S_2O

9. Which chemical equation represents the reaction of an Arrhenius acid and an Arrhenius base?
   A) $HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(\ell)$
   B) $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(\ell)$
   C) $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
   D) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaCl(aq)$

10. According to the Arrhenius theory, when a base dissolves in water it produces
    A) CO_3^{2-} as the only negative ion in solution
    B) OH^- as the only negative ion in solution
    C) NH_4^+ as the only positive ion in solution
    D) H^+ as the only positive ion in solution

11. Which substance yields hydroxide ion as the only negative ion in aqueous solution?
    A) $Mg(OH)_2$
    B) $C_2H_4(OH)_2$
    C) $MgCl_2$
    D) CH_3Cl

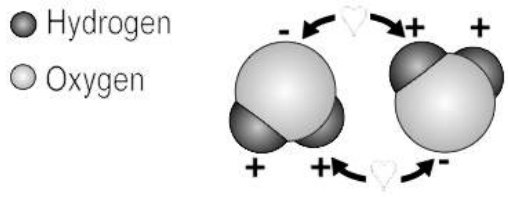
12. According to the Arrhenius theory, which list of compounds includes only bases?
    A) KOH, $Ca(OH)_2$, and CH_3OH
    B) KOH, NaOH, and LiOH
    C) LiOH, $Ca(OH)_2$, and $C_2H_4(OH)_2$
    D) NaOH, $Ca(OH)_2$, and CH_3COOH

6. Identify each of the following acids as monoprotic, diprotic or triprotic and state how many H^+ ions are released in solution.
    1. H_2CO_3
    2. H_3PO_4
    3. HNO_3
    4. CH_3COOH
    5. H_2SO_4
    6. HCl
    7. C_2H_5COOH
    8. H_2S

Video 13.2: Electrolytes

Answer the questions below based on the information above and on your knowledge of chemistry.

1. Using the diagram below, draw a diagram of the molecule formed when one hydrogen (proton) is pulled off one water molecule and attached to another.



2. What are the formulas of the ions formed?

3. Write an equation showing the formation of the ions from two molecules of water.

4. The hydronium ion (H_3O^+) and the hydroxide ion (OH^-) are formed by the reaction between water molecules. What would form from the reaction between hydronium and hydroxide ions? Write an equation showing the reaction.

5. How does the reaction between water molecules compare to the reaction between hydronium and hydroxide ions?

6. Pure water is actually a mixture of water molecules, hydronium ions, and hydroxide ions:

    a. How does the concentration of hydroxide ions compare to the concentration of hydronium ions in pure water? Explain

    b. How does the concentration of ions compare to the concentration of molecules in pure water?

Electrolytes

1. Which two compounds are electrolytes?

   A) CH_3OH and C_5H_{12}
   B) KOH and C_5H_{12}
   C) KOH and CH_3COOH
   D) CH_3OH and CH_3COOH

2. Which compounds can be classified as electrolytes?

   A) saturated hydrocarbons B) alcohols
   C) alkynes D) organic acids

3. Which laboratory test result can be used to determine if KCl(s) is an electrolyte?

   A) electrical conductivity of KCl(aq)
   B) pH of KCl(s)
   C) electrical conductivity of KCl(s)
   D) pH of KCl(aq)

4. Which sample of HCI(aq) contains the greatest number of moles of solute particles?

   A) 1.0 L of 2.0 M HCI(aq)
   B) 2.0 L of 2.0 M HCI(aq)
   C) 3.0 L of 0.50 M HCI(aq)
   D) 4.0 L of 0.50 M HCI(aq)

5. As water is added to a 0.10 M NaCl aqueous solution, the conductivity of the resulting solution

   A) increases because the concentration of ions decreases
   B) decreases, but the concentration of ions remains the same
   C) increases, but the concentration of ions remains the same
   D) decreases because the concentration of ions decreases

6. Which compound dissolves in water to form an aqueous solution that can conduct an electric current?

   A) C_2H_5OH B) $CC1_4$
   C) CH_3COOH D) CH_4

7. A student tested a 0.1 M aqueous solution and made the following observations:
   - conducts electricity
   - turns blue litmus to red
   - reacts with Zn(s) to produce gas bubbles

   Which compound could be the solute in this solution?

   A) LiOH B) LiBr C) CH_3OH D) HBr

8. Which of the following aqueous solutions is the best conductor of electricity?

   A) 1.0 M NaOH B) 0.10 M NaOH
   C) 0.10 M CH_3OH D) 1.0 M CH_3OH

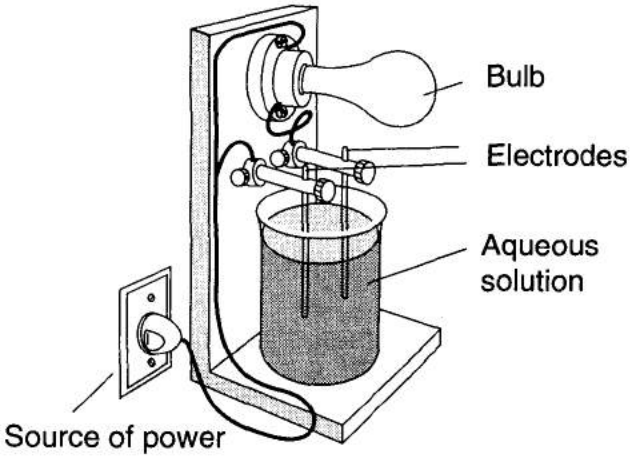
9. Based on Reference Table *F*, which of these salts is the best electrolyte?

   A) barium sulfate B) sodium nitrate
   C) silver chloride D) magnesium carbonate

10. An example of a nonelectrolyte is

    A) $K_2SO_4(aq)$ B) NaCl(aq)
    C) HCl(aq) D) $C_6H_{12}O_6(aq)$

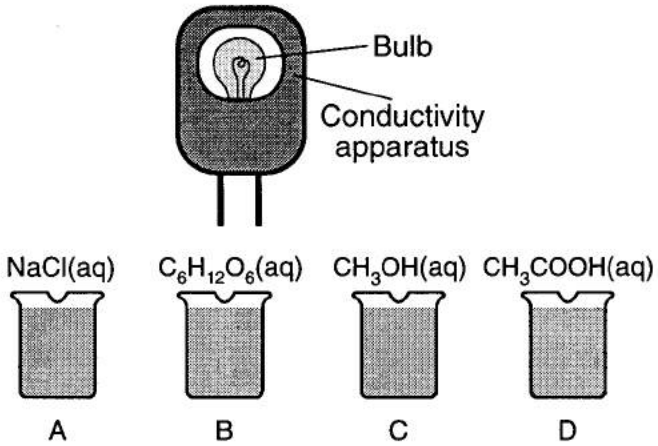
11. The diagram below shows an apparatus used to test the conductivity of various materials.



    Which aqueous solution will cause the bulb to light?

    A) $C_{12}H_{22}O_{11}(aq)$ B) LiOH(aq)
    C) $C_6HI_2O_6(aq)$ D) $CH_3OH(aq)$

12. Beakers *A, B, C,* and *D* shown below each contain a different solution.



    The bulb will glow when the conductivity apparatus is placed into which beakers?

    A) *A* and *D* B) *C* and *D*
    C) *B* and *C* D) *A* and *B*

Complete the following reactions, balance and indicate phase of the products

1. ___ $HNO_{3(aq)}$ + ___ $Mg_{(s)}$ →
2. ___ $H_2SO_{4(aq)}$ + ___ $Li_{(s)}$ →
3. ___ $HBr_{(aq)}$ + ___ $Zn_{(s)}$ →
4. Aluminum metal reacts w/ phosphoric acid.

5. Calcium metal reacts w/ hydrochloric acid

Using Table J answer the following questions.

1. Why does gold occur uncombined where as zinc does not?
   _____
2. Why was silver used to make coins in the past?
   _____
3. Why is copper used to make electrical wires and cables?
   _____
4. Why do we know little about the lifestyles of the people of the Iron Age?
   _____

Video 13.3: Conjugate Acid-Base Pairs

Operational Definition: Acids and bases are chemical species that exhibit distinctive sets of observable properties. Acids taste sour, bases are bitter.

Conceptual Definitions: Acids and bases are defined conceptually to help account for what is happening on the microscopic level.

Arrhenius Concept: An acid is a substance that when dissolved in water forms hydrogen ions (H^+). A base is a substance that when dissolved in water produces hydroxide ions (OH^-). This concept is limited because other chemicals have operational acid or base properties but do not form H^+ or OH^-. An example would be ammonia (NH_3) has basic properties but doesn't contain hydroxide ions.

Alternate Acid Base Concept (Bronsted-Lowry): An acid is a proton (H^+) donor and a base is a proton (H^+) acceptor.

$$HF + H_2O \Leftrightarrow H_3O^+ + F^-$$
Acid Base C. Acid C. Base

Fill in the following table. Identify the acid, base, conjugate acid and conjugate base in each of the equations.

	Equation	Acid	Base	Conjugate Acid	Conjugate Base
1.	$HCl + NH_3 \rightarrow NH_4^+ Cl^-$				
2.	$PO_4^{3-} + HNO_3 \rightarrow NO_3^- + HPO_4^{2-}$				
3.	$HCO_3^- + OH^- \rightarrow H_2O + CO_3^{2-}$				
4.	$NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$				
5.	$HPO_4^{2-} + H_2O \rightarrow OH^- H_2PO_4^-$				

6. Write the equation that shows ammonia, NH_3 reacting with hydrobromic acid, HBr. Label the acid, the base, the conjugate acid and the conjugate base.

7. Write the equation that shows the reaction of hydrogen sulfide, HS^- with hydroxide ion, OH^-. Label the acid, the base, the conjugate acid and the conjugate base.

Video 13.4: pH & Acid-Base Indicators

Indicators: Given the pH of the following substances, use table M of your reference to determine what color the indicator will turn when placed in each substance.

Solution	pH Range	Methyl Orange	Bromothymol blue	Phenolphthalein	Litmus	Bromocresol green	Thymol blue	Acid/Base
Vinegar	1.3							
Soap	8.4							
Cola	3.2							
Ammonia	12							
Coffee	5.2							

Alternate Acid-Base Theory

1. According to one acid-base theory, NH_3 acts as a base when an NH_3 molecule

   1) accepts an H^+ ion
   2) donates an H^+ ion
   3) accepts an OH^- ion
   4) donates an OH^- ion

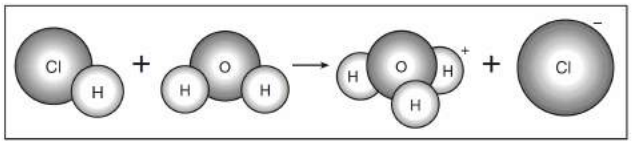
2. According to one acid-base theory, a base is an

   1) H^+ acceptor 2) H^+ donor
   3) Na^+ acceptor 4) Na^+ donor

3. A substance that dissolves in water and produces hydronium ions as the only positive ions in the solution is classified as

   1) an alcohol 2) an acid
   3) a base 4) a salt

4. Given the diagram representing a reaction:



   According to one acid-base theory, the water acts as

   1) a base because it accepts an H^+
   2) a base because it donates an H^+
   3) an acid because it accepts an H^+
   4) an acid because it donates an H^+

5. Given the balanced equation representing a reaction:

$$HSO_4^{\ -}(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + SO_4^{\ 2-}(aq)$$

   According to one acid-base theory, the $H_2O(\ell)$ molecules act as

   1) a base because they accept H^+ ions 2) a base because they donate H^+ ions
   3) an acid because they accept H^+ ions 4) an acid because they donate H^+ ions

6. One acid-base theory defines a base as an

   1) H^+ donor 2) H^+ acceptor
   3) H donor 4) H acceptor

7. One alternate acid-base theory states that an acid is an

   1) H^+ donor 2) H^+ acceptor
   3) OH^- donor 4) OH^- acceptor

8. Given the balanced equation representing a reaction:

$NH_3(g) + H_2O(\ell) \rightarrow NH_4^+(aq) + OH^-(aq)$

According to one acid-base theory, the $NH_3(g)$ molecules act as

1) an acid because they accept H^+ ions
2) an acid because they donate H^+ ions
3) a base because they accept H^+ ions
4) a base because they donate H^+ ions

9. Given the equation representing a reaction at equilibrium:

$NH_3(g) + H_2O(\ell) \leftrightarrow NH_4^+(aq) + OH^-(aq)$

The H^+ acceptor for the forward reaction is

1) $H_2O(\ell)$
2) $NH_3(g)$
3) $NH_4^+(aq)$
4) $OH^-(aq)$

10. Given the reaction:

$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

The water acts as the

1) base
2) acid
3) proton acceptor
4) electron donor

pH & Acid-Base Indicators

1. When the pH of an aqueous solution is changed from 1 to 2, the concentration of hydronium ions in the solution is

   1) decreased by a factor of 2
   2) decreased by a factor of 10
   3) increased by a factor of 2
   4) increased by a factor of 10

2. When the pH of a solution is changed from 4 to 3, the hydronium ion concentration of the solution

   1) decreases by a factor of 10
   2) increases by a factor of 10
   3) decreases by a factor of 100
   4) increases by a factor of 100

3. When the hydronium ion concentration of a solution is increased by a factor of 10, the pH value of the solution

   1) decreases 1 pH unit
   2) decreases 10 pH units
   3) increases 1 pH unit
   4) increases 10 pH units

4. When the hydronium ion concentration of a solution is increased by a factor of 10, the pH value of the solution

   1) decreases 1 pH unit
   2) decreases 10 pH units
   3) increases 1 pH unit
   4) increases 10 pH units

5. When the pH value of a solution is changed from 2 to 1, the concentration of hydronium ions

   1) decreases by a factor of 2
   2) increases by a factor of 2
   3) decreases by a factor of 10
   4) increases by a factor of 10

6. Which change in pH represents a hundredfold increase in the concentration of hydronium ions in a solution?

   1) pH1 to pH 2 2) pH 1 to pH 3
   3) pH 2 to pH 1 4) pH 3 to pH 1

7. What is the pH of a solution that has a hydronium ion concentration 100 times greater than a solution with a pH of 4?

   1) 5 2) 2 3) 3 4) 6

8. As the pH of a solution is changed from 3 to 6, the concentration of hydronium ions

   1) increases by a factor of 3
   2) increases by a factor of 1000
   3) decreases by a factor of 3
   4) decreases by a factor of 1000

9. Which pH indicates a basic solution?

   1) 1 2) 5 3) 7 4) 12

10. Which of these pH numbers indicates the highest level of acidity?

    1) 5 2) 8 3) 10 4) 12

11. Phenolphthalein is pink in an aqueous solution having a pH of

    1) 5 2) 2 3) 7 4) 12

12. What is the color of the indicator thymol blue in a solution that has a pH of 11?

    1) red 2) blue
    3) pink 4) yellow

13. Three samples of the same solution are tested, each with a different indicator. All three indicators, bromthymol blue, bromcresol green and thymol blue, appear blue if the pH of the solution is

    1) 4.7 2) 6.0 3) 7.8 4) 9.9

14. Based on the results of testing colorless solutions with indicators, which solution is most acidic?

    1) a solution in which bromthymol blue is blue
    2) a solution in which bromcresol green is blue
    3) a solution in which phenolphthalein is pink
    4) a solution in which methyl orange is red

15. Which indicator would best distinguish between a solution with a pH of 3.5 and a solution with a pH of 5.5

    1) bromthymol blue 2) bromcresol green
    3) litmus 4) thymol blue

Modeling Neutralization Reactions

Engage:

- Predict what will happen to the pH when equal volumes of acid and base are combined?
- Test pH of solutions then mix together. Test pH of resulting solution after mixing. Make observations
- Test conductivity before and after

Activity:

1. Write the balanced neutralization reaction for the reaction of HCl and NaOH similar to the one you saw in the "Upset Tummy" phenomena.
2. Using the key, draw particle models to represent the reaction before and after mixing. Draw five dissociated acid particles in beaker 1 and five dissociated base particles in beaker 2.
3. Match each hydrogen ion in beaker one with one hydroxide ion in beaker two. Then draw the number of water molecules that form in beaker three.

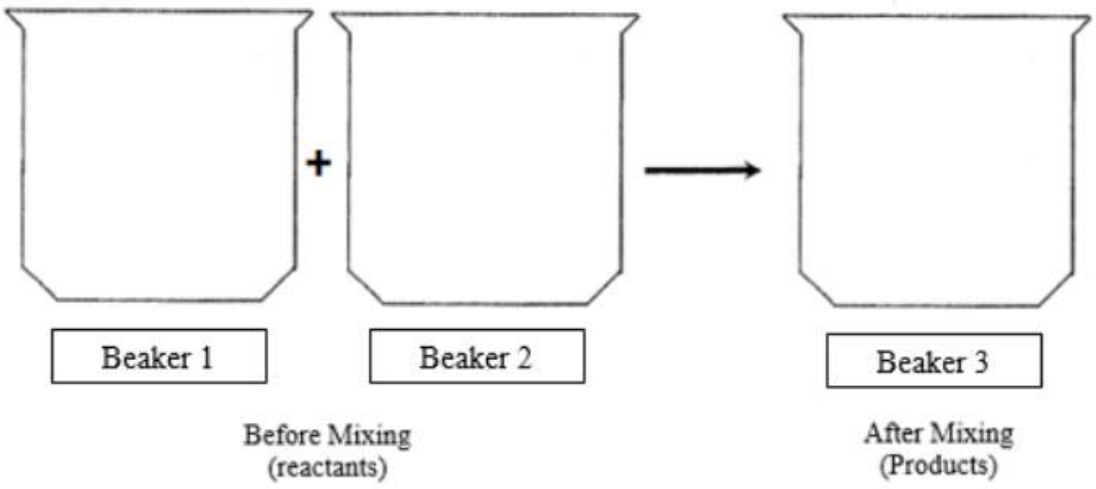
Key:

H^+ =

Cl^{-1} =

Na^+ =

OH^{-1} =



4. What was the pH of the resulting solution?

5. In a second neutralization reaction **H_2SO_4** and **NaOH** are mixed. Using the key below, draw three dissociated acid particles in beaker one and three dissociated base particles in beaker two.

Modeling Neutralization Reactions

Engage:

- Predict what will happen to the pH when equal volumes of acid and base are combined?
- Test pH of solutions then mix together. Test pH of resulting solution after mixing. Make observations
- Test conductivity before and after

Activity:

1. Write the balanced neutralization reaction for the reaction of HCl and NaOH similar to the one you saw in the "Upset Tummy" phenomena.
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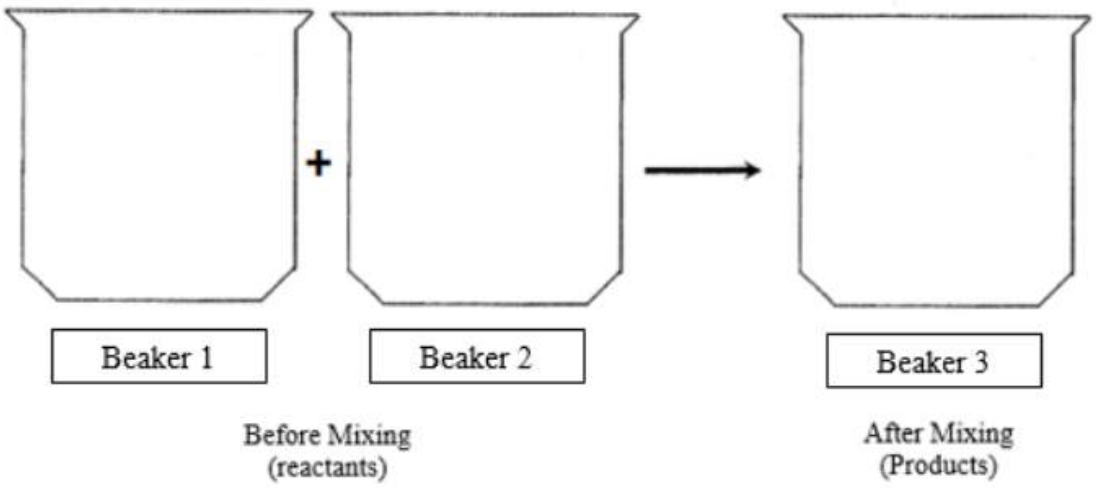

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5. In a second neutralization reaction **H_2SO_4** and **NaOH** are mixed. Using the key below, draw three dissociated acid particles in beaker one and three dissociated base particles in beaker two.

6. Match each hydrogen ion in beaker one with one hydroxide ion in beaker two. Then draw the number of water molecules that form in beaker three. Draw the remaining ions in breaker three.

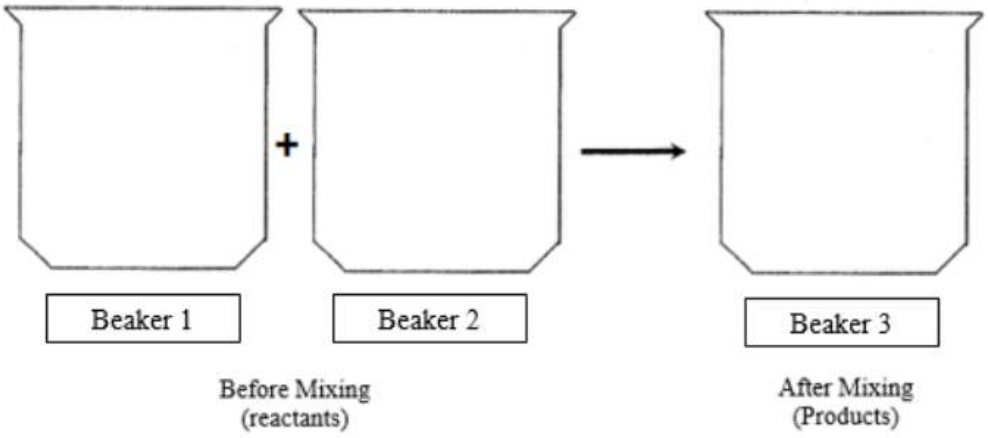
Key:

H^+ =

SO_4^{-2} =

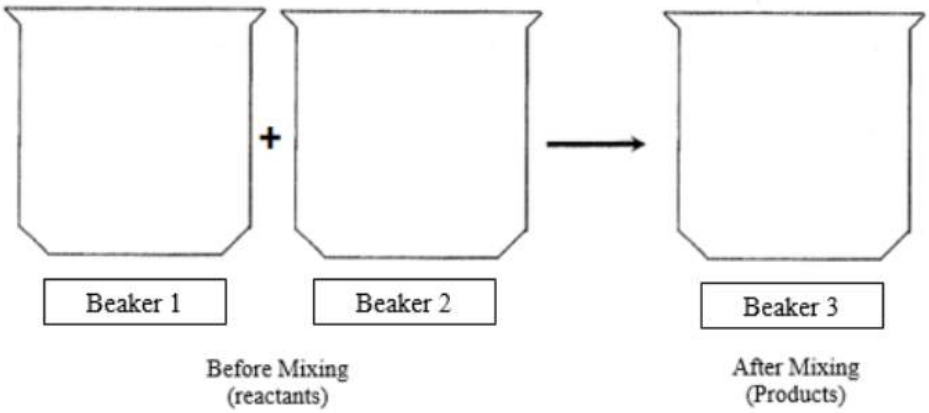
Na^+ =

OH^{-1} =



7. What additional ion(s) would be needed to neutralize the solution? How many more would you need?

8.Draw a model below to justify that neutralization has occurred.



8. Write the **balanced** equation for the complete neutralization of H_2SO_4 and NaOH .

6. Match each hydrogen ion in beaker one with one hydroxide ion in beaker two. Then draw the number of water molecules that form in beaker three. Draw the remaining ions in breaker three.

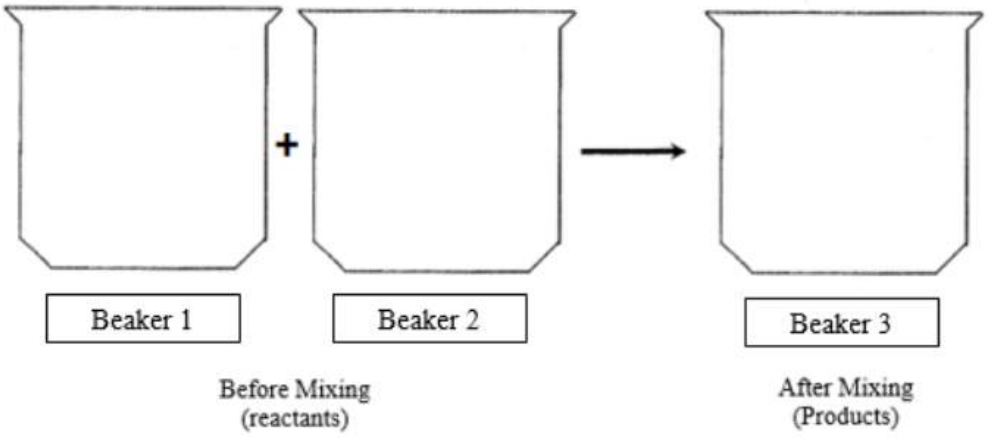
Key:

H^+ =

SO_4^{-2} =

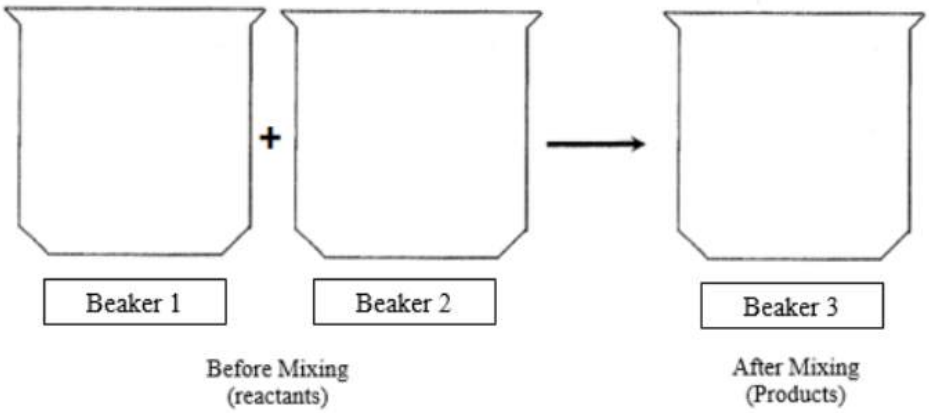
Na^+ =

OH^{-1} =



7. What additional ion(s) would be needed to neutralize the solution? How many more would you need?

8.Draw a model below to justify that neutralization has occurred.



8. Write the **balanced** equation for the complete neutralization of H_2SO_4 and NaOH .

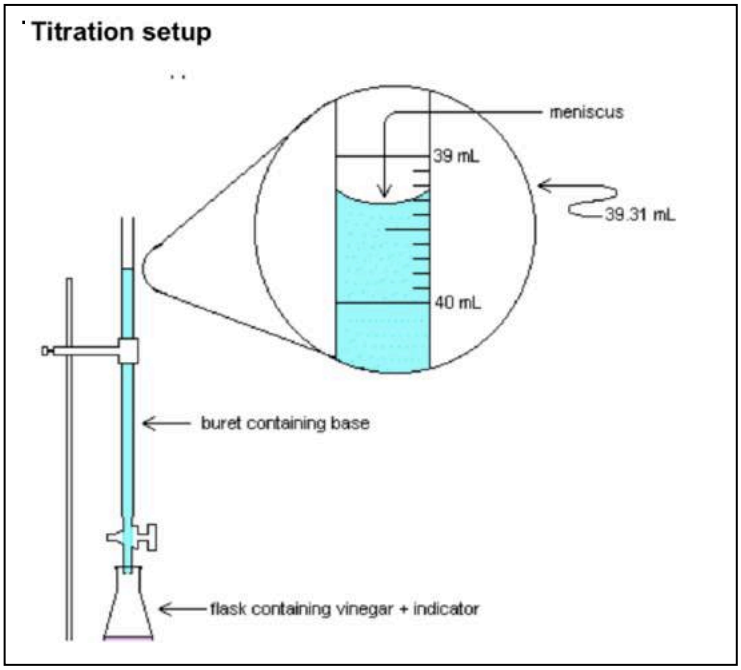
Titration and Neutralization Calculations

Do Now: Using your knowledge of chemistry, fill in each blank to complete the statements.

1. The chemicals HCl, HBr, HNO_3, H_2SO_4 are all categorized as acids because when dissolved in water they produce a _____ ion.
2. The chemicals NaOH, LiOH, and $Ca(OH)_2$ are all categorized as bases because when dissolved in water they produce a _____ ion.
3. When acids react with bases the reaction is known as a _____ reaction because the properties of the original acid and base are lost to form a _____ and water.
4. Calculate the Molarity of an acid created by dissolving 73.0g of HCl in 200.0mL if solution.

5. Calculate the moles of NaOH dissolved in a 2.0M basic solution with a total volume of 500.0mL.

READ ME! Titrations are procedures used to determine the concentration or Molarity (M) of an unknown acid or a base. The acid and base solutions are combined together in a ratio of 1 mole of Acid for every 1 mole of Base, thus neutralizing the reaction. Observe the titration set up below. A student placed 20.00mL of an unknown acid and two drops of phenolphthalein indicator in the flask. Then the student added the known base dropwise until the solution was neutralized (indicated by the color change of the phenolphthalein). Answer questions 1 through 4.



Read the volume of the base that was released from the burette

If the base concentration was known to be 2.00M, calculate the number of moles of basic ion are dissolved in the solution.

If the solution in the flask is now neutral, how many moles of acid ion are in the flask?

Calculate the molarity of the unknown

STOP

Using the equation on Reference Table T, you can solve for either the molarity/concentration (M) or a volume added (V) using the titration formula:

$$M_A V_A = M_B$$

M_A= molarity of H^+ V_A = volume of acid

M_B = molarity of OH^- V_B = volume of base

5. What unit of measurement is obtained when the Molarity of a solution is multiplied by the volume (in liters) of the solution?

6. Explain why the Titration formula is mathematically accurate for finding the concentration of solution when the solution is neutralized.

Notice that the Molarity formula requires the volume to be calculated in Liters. However, the titration formula has no such requirement, as long as the volume unit of measurement is the same for both V_A and V_B. Since the titration formula has two volumes, the volume unit of measurement will cancel out. Therefore, converting the units will also cancel out. Perform the calculations in question 7 to prove this rule.

7. A 100.00mL solution of 2.00M HCl is neutralized by 50.00mL of NaOH. Calculate the Molarity of the unknown base.
   a. Use the titration formula and calculate the Molarity with the volume values in mL.

   b. Use the titration formula and calculate the Molarity with the volume values in mL.

   c. Compare your answers to questions 7a and 7b.

8. A 25.0-milliliter sample of HNO_3 (aq) is neutralized by 32.1 milliliters of 0.150 M KOH (aq). Calculate the concentration of the acid.

9. The titration formula can also be used to determine the volume needed to neutralize two known solutions. Calculate the volume of 0.200 M NaOH needed to neutralize 100. mL of 0.100 M HCl.

For each question, the two pH values are being compared. How many times stronger or weaker is the pH of the solution?

1. pH 5 → pH 3 _____
2. pH 8 → pH 4 _____
3. pH 10 → pH 7 _____
4. pH 14 → pH 7 _____
5. pH 3 → pH 6 _____

Video 13.5: Neutralization & Titrations

Complete and balance each of the acid base neutralization reactions below.

1. _____ H_2SO_4 + _____ $Mg(OH)_2$ →
2. _____ HNO_3 + _____ $Al(OH)_3$ →
3. _____ H_3PO_4 + _____ $Ca(OH)_2$ →
4. _____ HI + _____ KOH →
5. _____ HBr + _____ $Ba(OH)_2$ →
6. _____ HCl + _____ KOH →
7. _____ H_3PO_4 + _____ LiOH →
8. _____ HF + _____ $Ca(OH)_2$ →
9. In a titration of $HClO_4$ with NaOH, 100.0 mL of the base was required to neutralize 20.0 mL of 5.0 M $HClO_4$. What is the molarity of the NaOH? (Be sure to write the neutralization reaction.)

10. In a titration of HNO_3 with NaOH, 60.0 mL of 0.020 M NaOH was needed to neutralize 15.0 mL of HNO_3. What is the molarity of the acid? (Write the neutralization reaction.)

11. In a titration, 20.0 milliliters of 0.15 M HCl(aq) is exactly neutralized by 18.0 milliliters of KOH(aq).
    (a) Complete the equation below for the neutralization reaction by writing the formula of *each* product.
        KOH(aq) + HCl(aq) → _____ + _____

    (b) Compare the number of moles of H^+(aq) ions to the number of moles of OH^- (aq) ions in the titration mixture when the HCl(aq) is exactly neutralized by the KOH(aq).

    (c) Determine the concentration of the KOH(aq).

    (d) What is the new pH of the solution?

12. In a laboratory activity, 0.500 mole of NaOH(s) is completely dissolved in distilled water to form 400. milliliters of NaOH(aq). This solution is then used to titrate a solution of HNO_3(aq).
    (a) Identify the negative ion produced when the NaOH(s) is dissolved in distilled water.

    (b) Calculate the molarity of the NaOH(aq). Your response must include *both* a correct numerical setup and the calculated result.

    (c) If 26.4 milliliters of the NaOH solution is needed to exactly neutralize 44.0 milliliters of the HNO_3 solution, what is the molarity of the HNO_3 solution?

    (d) Complete the equation below representing this titration reaction by writing the formulas of the products.
        NaOH(aq) + HNO_3(aq) → _____ + _____

13. If 10.0 mL of 0.300 M KOH are required to neutralize 30.0 mL of stomach acid (HCl), what is the molarity of the stomach acid? (Write the neutralization reaction.)

14. If it takes 50 mL of 0.5 M KOH solution to completely neutralize 125 mL of sulfuric acid solution (H_2SO_4), what is the concentration of the H_2SO_4 solution?

Titration Practice:
A titration was set up and used to determine the unknown molar concentration of a solution of NaOH. A 1.2 M HCl solution was used as the titration standard. The following data were collected.

	Trial 1	Trial 2	Trial 3	Trial 4
Volume of 1.2 M HCl	10.0 mL	10.0 mL	10.0 mL	10.0 mL
Intial buret reading of NaOH	0.0 mL	12.2 mL	23.2 mL	35.2 mL
Final buret reading of NaOH	12.2 mL	23.2 mL	35.2 mL	47.7 mL
Volume of NaOH used (mL)				
Molarity of NaOH (M)				

1. Calculate the volume of NaOH used to neutralize the acid for each trial. Record in the data table above. Show one sample calculation below.

2. Using the $M_AV_A = M_BV_B$ formula, calculate the molarity of the base for each trial. Record in the data table above. Show one sample calculation below.

3. Calculate the average molarity of the NaOH using your results from question 2. Your answer must include the correct number of significant figures and the correct units of measure.

Name: **Acid/Base/Salt Review**

1. Which laboratory test result can be used to determine if KCl(s) is an electrolyte?

   A) pH of KCl(aq)
   B) pH of KCl(s)
   C) electrical conductivity of KCl(aq)
   D) electrical conductivity of KCl(s)

2. Which substance is an electrolyte?

   A) CCl_4 B) C_2H_6
   C) HCl D) H_2O

3. Which two compounds are electrolytes?

   A) $C_6H_{12}O_6$ and CH_3CH_2OH
   B) $C_6H_{12}O_6$ and HCl
   C) NaOH and HCl
   D) NaOH and CH_3CH_2OH

4. A substance is classified as an electrolyte because

   A) it has a high melting point
   B) it contains covalent bonds
   C) its aqueous solution conducts an electric current
   D) its aqueous solution has a pH value of 7

5. Which of the following aqueous solutions is the best conductor of electricity?

   A) 0.10 M CH_3OH B) 1.0 M CH_3OH
   C) 0.10 M NaOH D) 1.0 M NaOH

6. A student was given four unknown solutions. Each solution was checked for conductivity and tested with phenolphthalein. The results are shown in the data table below.

Solution	Conductivity	Color with Phenolphthalein
A	Good	Colorless
B	Poor	Colorless
C	Good	Pink
D	Poor	Pink

   Based on the data table, which unknown solution could be 0.1 M NaOH?

   A) *A* B) *B* C) *C* D) *D*

7. Which compound is an Arrhenius acid?

   A) CaO B) HCI C) K_2O D) NH_3

8. When one compound dissolves in water, the only positive ion produced in the solution is H_3O^+ (aq). This compound is classified as

   A) a salt B) a hydrocarbon
   C) an Arrhenius acid D) an Arrhenius base

9. Which substance is always a product when an Arrhenius acid in an aqueous solution reacts with an Arrhenius base in an aqueous solution?

   A) HBr B) H_2O
   C) KBr D) KOH

10. Given the equation:

    $$HCl(g) + H_2O(\ell) \rightarrow X(aq) + Cl^-(aq)$$

    Which ion is represented by *X*?

    A) hydroxide B) hydronium
    C) hypochlorite D) perchlorate

11. An aqueous solution of lithium hydroxide contains hydroxide ions as the only negative ion in the solution. Lithium hydroxide is classified as an

    A) aldehyde B) alcohol
    C) Arrhenius acid D) Arrhenius base

12. According to the Arrhenius theory, an acid is a substance that

    A) changes litmus from red to blue
    B) changes phenolphthalein from colorless to pink
    C) produces hydronium ions as the only positive ions in an aqueous solution
    D) produces hydroxide ions as the only negative ions in an aqueous solution

13. Which compound releases hydroxide ions in an aqueous solution?

    A) CH_3COOH B) CH_3OH
    C) HCl D) KOH

14. Which formula represents a hydronium ion?

    A) H_3O^+ B) NH_4+
    C) OH^- D) HCO_3^-

15. A solution with a pH of 2.0 has a hydronium ion concentration ten times greater than a solution with a pH of

    A) 1.0 B) 0.20 C) 3.0 D) 20

16. Which change in pH represents a hundredfold increase in the concentration of hydronium ions in a solution?

A) pH1 to pH 2 B) pH 1 to pH 3
C) pH 2 to pH 1 D) pH 3 to pH 1

17. What is the pH of a solution that has a hydronium ion concentration 100 times greater than a solution with a pH of 4?

A) 5 B) 2 C) 3 D) 6

18. Which pH indicates a basic solution?

A) 1 B) 5 C) 7 D) 12

19. Which of these pH numbers indicates the highest level of acidity?

A) 5 B) 8 C) 10 D) 12

20. Based on the results of testing colorless solutions with indicators, which solution is most acidic?

A) a solution in which bromthymol blue is blue
B) a solution in which bromcresol green is blue
C) a solution in which phenolphthalein is pink
D) a solution in which methyl orange is red

21. Which indicator would best distinguish between a solution with a pH of 3.5 and a solution with a pH of 5.5

A) bromthyrnol blue B) bromcresol green
C) litmus D) thymol blue

22. Which indicator, when added to a solution, changes color from yellow to blue as the pH of the solution is changed from 5.5 to 8.0?

A) bromcresol green B) bromthymol blue
C) litmus D) methyl orange

23. The table below shows the color of the indicators methyl orange and litmus in two samples of the same solution.

Results of Acid-Base Indicator Tests

Indicator	Color Result from the Indicator Test
methyl orange	yellow
litmus	red

Which pH value is consistent with the indicator results?

A) 1 B) 5 C) 3 D) 10

24. Which indicator is yellow in a solution with a pH of 9.8?

A) methyl orange B) bromthymol blue
C) bromcresol green D) thymol blue

25. One acid-base theory defines a base as an

A) H^+ donor B) H^+ acceptor
C) H donor D) H acceptor

26. One alternate acid-base theory states that an acid is an

A) H^+ donor B) H^+ acceptor
C) OH^- donor D) OH^- acceptor

27. Which statement describes an alternate theory of acids and bases?

A) Acids and bases are both H^+ acceptors.
B) Acids and bases are both H^+ donors.
C) Acids are H^+ acceptors, and bases are H^+ donors.
D) Acids are H^+ donors, and bases are H^+ acceptors.

28. Given the reaction at equilibrium:

$$HSO_4^- + NH_3 \leftrightarrow SO_4^{2-} + NH_4^+$$

What are the two species that are acids?

A) NH_3 and NH_4^+
B) NH_3 and SO_4^{2-}
C) HSO_4^- and SO_4^{2-}
D) HSO_4^- and NH_4^+

29. According to one acid-base theory, water acts as an acid when an H_2O molecule

A) accepts an H^+ B) donates an H^+
C) accepts an H^- D) donates an H^-

30. Which compound is produced when HCl(aq) is neutralized by $Ca(OH)_2(aq)$?

A) $CaCl_2$ B) CaH_2
C) HClO D) $HClO_2$

31. Which word equation represents a neutralization reaction?

A) base + acid $\rightarrow$ salt + water
B) base + salt $\rightarrow$ water + acid
C) salt + acid $\rightarrow$ base + water
D) salt + water $\rightarrow$ acid + base

32. Which equation represents a neutralization reaction?

A) $4Fe(s) + 3O_2(g) \rightarrow Fe_2O_3(s)$
B) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$
C) $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(\ell)$
D) $AgNO_3(aq) + KCl(aq) \rightarrow KNO_3(aq) + AgCl(s)$

33. During which process can 10.0 milliliters of a 0.05 M HCl(aq) solution be used to determine the unknown concentration of a given volume of NaOH(aq) solution?

A) evaporation B) distillation
C) filtration D) titration

34. A student completes a titration by adding 12.0 milliliters of NaOH(aq) of unknown concentration to 16.0 milliliters of 0.15 M HCl(aq). What is the molar concentration of the NaOH(aq)?

A) 0.11 M B) 0.20 M
C) 1.1 M D) 5.0 M

35. A 25.0-milliliter sample of $HNO_3(aq)$ is neutralized by 32.1 milliliters of 0.150 M KOH(aq). What is the molarity of the $HNO_3(aq)$?

A) 0.117 M B) 0.150 M
C) 0.193 M D) 0.300 M

36. Which volume of 0.10 M NaOH(aq) exactly neutralizes 15.0 milliliters of 0.20 M $HNO_3(aq)$?

A) 1.5 mL B) 7.5 mL
C) 3.0 mL D) 30. mL

37. What volume of 0.120 M $HNO_3(aq)$ is needed to completely neutralize 150.0 milliliters of 0.100 M NaOH(aq)?

A) 62.5 mL B) 125 mL
C) 180. mL D) 360. mL

Base your answers to questions **38** and **39** on the information below.

In one trial of an investigation, 50.0 milliliters of HCI(aq) of an unknown concentration is titrated with 0.10 M NaOH(aq). During the titration, the total volume of NaOH(aq) added and the corresponding pH value of the reaction mixture are measured and recorded in the table below.

Titration Data

Total Volume of NaOH(aq) Added (mL)	**pH Value of Reaction Mixture**
10.0	1.6
20.0	2.2
24.0	2.9
24.9	3.9
25.1	10.1
26.0	11.1
30.0	11.8

38. Write a balanced equation that represents this neutralization reaction.

39. In another trial, 40.0 milliliters of HCI(aq) is completely neutralized by 20.0 milliliters of this 0.10 M NaOH(aq). Calculate the molarity of the titrated acid in this trial. Your response must include *both* a numerical setup and the calculated result.

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40. Base your answer to the following question on the information below.

In liquid water, an equilibrium exists between $H_2O(\ell)$ molecules, $H^+(aq)$ ions, and $OH^-(aq)$ ions. A person experiencing acid indigestion after drinking tomato juice can ingest milk of magnesia to reduce the acidity of the stomach contents. Tomato juice has a pH value of 4. Milk of magnesia, a mixture of magnesium hydroxide and water, has a pH value of 10.

Compare the hydrogen ion concentration in tomato juice to the hydrogen ion concentration in milk of magnesia.

Base your answers to questions **41** through **43** on the information below.

A student used blue litmus paper and phenolphthalein paper as indicators to test the pH of distilled water and five aqueous household solutions. Then the student used a pH meter to measure the pH of the distilled water and each solution. The results of the student's work are recorded in the table below.

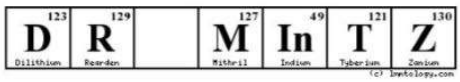
Testing Results

Liquid Tested	Color of Blue Litmus Paper	Color of Phenolphthalein Paper	Measured pH Value Using a pH Meter
2% milk	blue	colorless	6.4
distilled water	blue	colorless	7.0
household ammonia	blue	pink	11.5
lemon juice	red	colorless	2.3
tomato juice	red	colorless	4.3
vinegar	red	colorless	3.3

41. Explain, using the reference table, in terms of the pH range for color change why litmus is *not* appropriate to differentiate the acidity levels of tomato juice and vinegar.

42. Identify the liquid tested that has the *lowest* hydronium ion concentration.

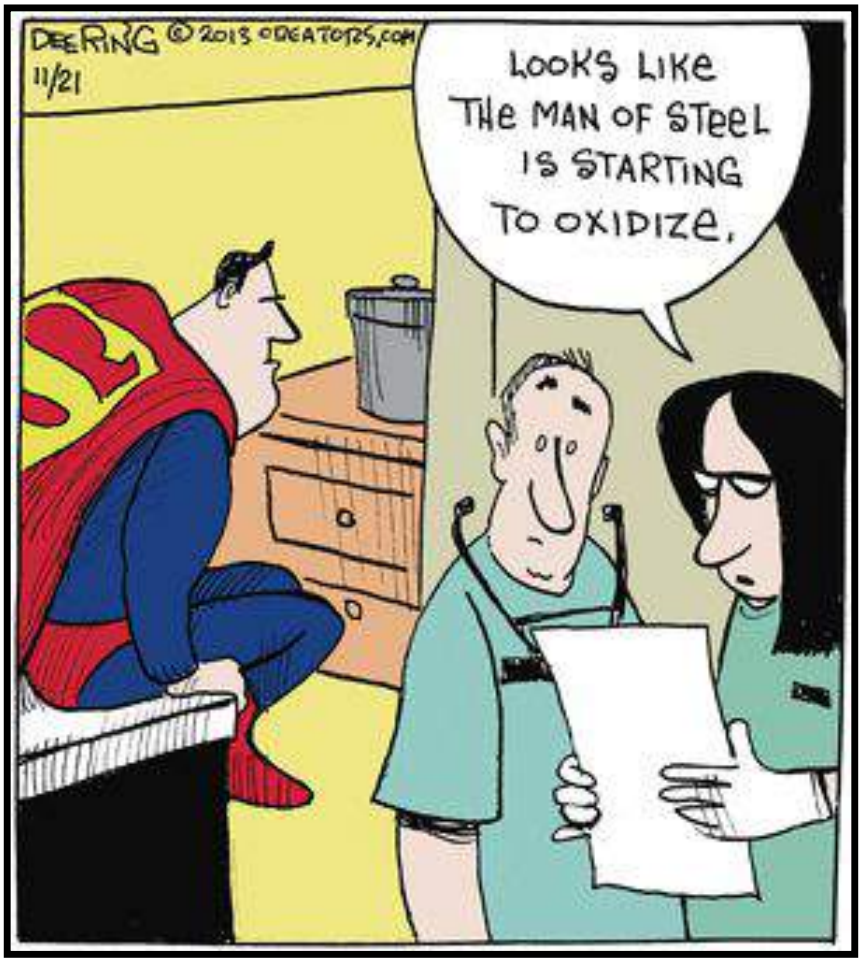
43. Based on the measured pH values, identify the liquid tested that is 10 times more acidic than vinegar.



Regents Chemistry: Mrs. Mintz

Practice Packet

Chapter 14: Oxidation Reduction & Electrochemistry



Chapter 14: Electrochemistry

Anode – the electrode at which oxidation occurs (negative)

Cathode – The electrode at which reaction occurs (positive)

Electrode – a conductor in a circuit that carries electrons to or from a substance other than a metal

Electrochemical Cell – any device that converts chemical energy into electrical energy or electrical energy into chemical energy

Electrolytic Cell – a cell that uses electricity from an outside source to force a nonsponstanous redox reaction to occur; Examples: recharging batteries, electroplating

Electroplating – an electrolytic process that involved oxidizing a source metal into a solution with the use of an external power source and then reducing the metal ion onto a metallic object that is to be plated ("silver plated")

Half Reaction – a reaction that describes the change in oxidation number and subsequent gain or loss of electrons that occurs during oxidation or reduction

Oxidation – a process that involves complete or partial loss of electrons or gain of oxygen; results in an increase in the oxidation number

Oxidizing Agent – the species that is reduced and therefore removes the electrons from the species that was oxidized

Oxidation Number – the charge of an ion

Redox reaction – another name for an oxidation-reduction reaction; a reaction that involves the transfer of electrons between reactants

Reduction – a process that involves a complete or partial gain of electrons or the loss of oxygen; it results in a decrease in the oxidation number

Reducing Agent – the species that is oxidized and therefore gives electrons to the species that was reduced.

Salt Bridge – a tube containing a strong electrolyte used to separate the half-cells in a voltaic cell; it allows the passage of ions form one half-cell to the other but prevents the solutions from mixing

Species – the symbol and charge of an element or ion in a redox reaction

Voltaic Cell – an electrochemical cell that produces electric current as a result of a spontaneous redox reaction, used to make batteries. Consists of two half-cells connected by a salt bridge and two electrodes that connect to a load (device) that uses the electricity produced by the cell.

Assigning Oxidation Numbers

Chemistry 200
Video Lesson 14.1

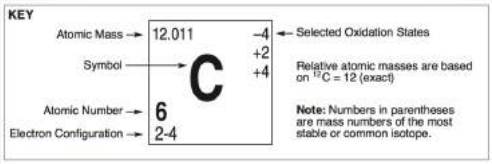
Objective:

How do we assign atoms the correct oxidation number?

Oxidation & Reduction (Redox)

Oxidation number

- a number assigned to keep track of electron(s) gained or lost in a redox rxns.



Rules for Assigning Oxidation Numbers (states)

1. Free elements are assigned an oxidation state of zero

   The hydrogen in H_2, the sodium in Na, and the sulfur in S_8 all have oxidation numbers of **zero.**

2. The Group 1 metals (Alkali Metals) in compounds always have an oxidation state of +1
3. The Group 2 metals (Alkali Earth Metals) in compounds always have an oxidation state of +2
4. The oxidation state for any simple one-atom ion(monoatomic) is equal to its charge.

In $MgCl_2$, the Mg^{2+} *ion* has an oxidation number of +2. Each of the Cl^- *ions* has an oxidation number of *-1*.

In $FeCl_2$, the Fe^{2+} *ion* has an oxidation number of +2, while in $FeCl_3$ the Fe^{3+} *ion* has an oxidation number of +3.

5. Fluorine in compounds is always assigned an oxidation state of -1.
6. Hydrogen in compounds is always assigned an oxidation number of +1. (HCl). If combined w/ a metal, Hydrogen has an oxidation # of -1. (LiH)

7. Oxygen in compounds is assigned an oxidation state of -2.

When combined w/ Fluorine, Oxygen has an oxidation state of +2.

8. The sum of the oxidation states of all the atoms in a species must be equal to the net charge on the species.

If the species is neutral, the sum of the oxidation states is zero.

If the species has a charge, the sum is equal to that charge.

Determine the oxidation number of sulfur in H_2SO_4

$\frac{(+2)|+6|(-8)}{+1|\boxed{+6}|-2} = 0$

H_2SO_4

Determine the oxidation number of nitrogen in NO_3^{-1}

$\frac{+5|(-6)}{\boxed{+5}|-2} = -1$

NO_3^{-1}

Redox Reactions

Chemistry 200
Video Lesson 14.2

Objective:

How do we identify substances being oxidized or reduced and the oxidizing and reducing agents.

How do we use this information to create half reactions ?

Redox

Redox reaction

- a chemical rxn where electron(s) are transferred from one atom to another

Oxidation

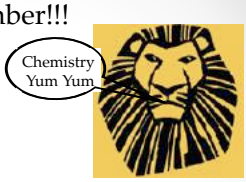
- loss of electron(s) by an atom or ion & an increase in oxidation # ***($Ca \rightarrow Ca^{+2} + 2e^-$)***

Reduction

- gain of electron(s) by an atom or ion & a decrease in oxidation # ***($5e^- + Mn^{+7} \rightarrow Mn^{+2}$)***

*****Oxidation & Reduction occur simultaneously, they cannot occur separately*****

Easy way to remember!!!



LEO says GER

LEO = Loss of Electrons is Oxidation

GER = Gain of Electrons is Reduction

Oxidizing Agent

- the substance that causes the ***oxidation*** of other substances. They accept electron(s) easily & therefore are reduced

ex$_1$: $5e^- + Mn^{+7} \rightarrow Mn^{+2}$

(gain of electrons, oxidation # reduced)

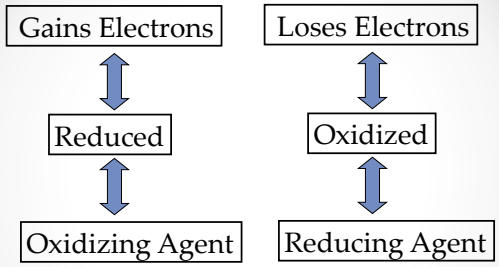
Reducing Agent

- the substance that allows another substance to be ***reduced.*** They lose electron(s) & therefore are oxidized

ex: $Ca \rightarrow Ca^{+2} + 2e^-$

(loss of electrons, oxidation # increases)

The species oxidized, the species reduced, the oxidizing agent or reducing agent must always be a reactant!!



Writing Half Reactions

- They show either the oxidation or reduction portion of a redox rxn, including electrons gained or lost

Half rxn for oxidation --> electrons are products
Half rxn for reduction --> electrons are reactants

THERE'S AN EASIER WAY TO REMEMBER!!!

***** Electrons go on the side w/ the more positive oxid. # *****

Creating Half Reactions from a Redox Reaction

1. Assign oxidation numbers to determine the substance oxidized & the substance reduced.
2. Write each half reaction, balance by atom, then by electrons. The ONLY exception is a half rxn the contains a diatomic element.
3. Identify the species oxidized, the species reduced the oxidizing agent & the reducing agent.

+1 +5 -6
+1 +5 -2 +1 -1 0

$KClO_3 \rightarrow KCl + O_2$

Oxid: $2O^{-2} \rightarrow O_2 + 4e^-$ *(reducing agent)*

Red: $6e^- + Cl^{+5} \rightarrow Cl^{-1}$ *(oxidizing agent)*

Electrochemical Cells

Video Lesson 14.4

Objectives

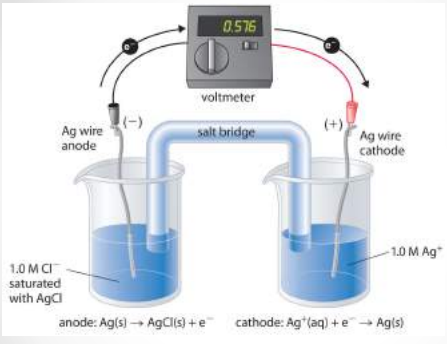
- Name the type of reactions involved in electrochemical processes.
- Describe how a voltaic cell produces electrical energy.

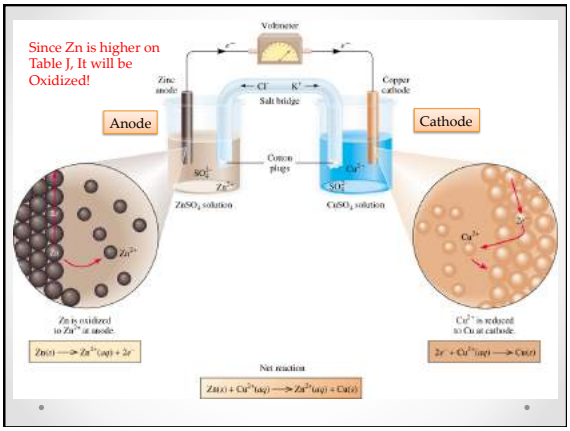
Electrochemical Process

- An electrochemical process is any conversion between chemical and and electrical energy (charged particles).
- All electrochemical processes involve redox reactions.
- The two half reactions must be physically separated to be used as a source of electrical energy.

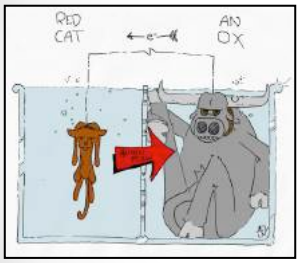
Voltaic Cells

- To obtain a useful current, we separate the oxidizing and reducing agents so that electron transfer occurs thru an external wire.
- What is needed – Spontaneous Reaction
  - Salt Bridge
    - connects the 2 containers & provides a passage for ions from one solution to another
  - Electrodes
    - one of 2 surfaces that conduct electricity, the site of oxidation & reduction rxns





Flow of Electrons



Reduction occurs at the cathode

Oxidation occurs at the anode

From Anode to Cathode

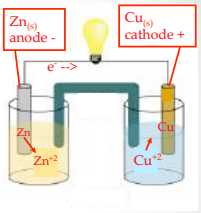
Voltaic Cell

Overall Redox Rxn

$Zn_{(s)} + Cu^{+2}_{(aq)} \rightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$

Oxid: $Zn_{(s)} \rightarrow Zn^{+2}_{(aq)} + 2e^-$

Red: $2e^- + Cu^{+2}_{(aq)} \rightarrow Cu_{(s)}$



Species oxidized Zn

Reducing agent Zn

Species reduced Cu^{+2}

Oxidizing agent Cu^{+2}

Anode Zn

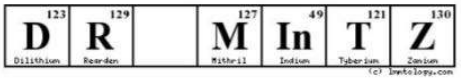
Cathode Cu

Sketch Notes

Name: _____ Date: _____ Per: _____

Assigning Oxidation Numbers

	Formula	Element and Oxidation Number					
1.	NO_2	**N**		**O**			
2.	KCl	**K**		**Cl**			
3.	MnO_2	**Mn**		**O**			
4.	H_2SO_4	**H**		**S**		**O**	
5.	K_3PO_4	**K**		**P**		**O**	
6.	HNO_3	**H**		**N**		**O**	
7.	Fe_2O_3	**Fe**		**O**			
8.	$CaCl_2$	**Ca**		**Cl**			
9.	$Na_2S_2O_3$	**Na**		**S**		**O**	
10.	CH_4	**C**		**H**			



Video 14.1: Oxidation Numbers

Determine the oxidation number of **EACH** element in each of the following substances:

1. $KMnO_4$
2. S_8
3. $Na_2Cr_2O_7$
4. $CuCl_2$
5. H_2SO_4
6. PO_4^{-3}
7. zinc oxide
8. water
9. calcium hydride
10. carbon dioxide

Video 14.2: Redox Reactions

Writing Half Reactions

Step 1: Balance the half reaction by atom using coefficients
Step 2: Balance the half reaction by charge using electrons (e^-)

Electrons always go on the side with the more positive oxidation #

Half rxns are written on a per atom basis. The only exception is if the half rxn contains a diatomic element

$Mg \rightarrow Mg^{+2}$

$O_2 \rightarrow O^{-2}$

$Cl^{+5} \rightarrow Cl^{-1}$

$Mn^{+2} \rightarrow Mn^{+5}$

$N^{-3} \rightarrow N_2$

$S^{+6} \rightarrow S^{-2}$

$3Cu^{+1} \rightarrow 3Cu$

$6H^{+1} \rightarrow 3H_2$

Name:

14.1 Oxidation Numbers

____ 1. What is the oxidation state of nitrogen in the compound NH_4Br?

1) –1 2) +2 3) –3 4) +4

____ 2. What is the oxidation number of sulfur in $Na_2S_2O_3$?

1) –1 2) +2 3) +6 4) +4

____ 3. Given the balanced equation representing a reaction:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

The oxidation state of chlorine in this reaction changes from

1) –1 to +1
2) –1 to +5
3) + I to –1
4) +5 to –1

____ 4. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-} ?

1) +6 2) +2 3) +3 4) +8

____ 5. What is the oxidation number assigned to manganese in $KMnO_4$?

1) +7 2) +2 3) +3 4) +4

____ 6. What is the oxidation state of nitrogen in $NaNO_2$?

1) +1 2) +2 3) +3 4) +4

____ 7. What is the oxidation number of chromium in $K_2Cr_2O_7$?

1) +12 2) +2 3) +3 4) +6

____ 8. Given the reaction that occurs in an electrochemical cell:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

During this reaction, the oxidation number of Zn changes from

1) 0 to +2
2) 0 to –2
3) +2 to 0
4) –2 to 0

____ 9. In which substance does hydrogen have an oxidation number of zero?

1) LiH
2) H_2O
3) H_2S
4) H_2

____ 10. In which compound does carbon have an oxidation state of –4?

1) CO
2) CO_2
3) CCl_4
4) CH_4

Creating Half Reactions from Redox Reactions

Step 1: Assign oxidation numbers to determine the substances that have a change in oxidation number
Step 2: Identify the two half reactions & balance each half reaction
Step 3: Label each half reaction & identify the species oxidized, the species reduced, the oxidizing agent & the reducing agent

Note: The reaction provided to you may be balanced or unbalanced. Remember that half reactions are made on a per atom basis. The ONLY exception is a half reaction that contains a diatomic element

$$CuSO_4 + Mn \rightarrow Mn(SO_4)_2 + Cu$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$AgF + Fe \rightarrow FeF_3 + Ag$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$Na + Cl_2 \rightarrow NaCl$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$H_2 + O_2 \rightarrow H_2O$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$S + O_2 \rightarrow SO_3$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$Fe + HCl \rightarrow H_2 + FeCl_3$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$KClO_3 \rightarrow KCl + O_2$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

Name:

14.2 Half Reactions

1. An oxidation-reduction reaction involves the

1) sharing of electrons
2) sharing of protons
3) transfer of electrons
4) transfer of protons

2. During which process does an atom gain one or more electrons?

1) transmutation
2) reduction
3) oxidation
4) neutralization

3. Which half-reaction correctly represents reduction?

1) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
2) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
3) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
4) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

4. In a redox reaction, the total number of electrons lost is

1) less than the total number of electrons gained
2) greater than the total number of electrons gained
3) equal to the total number of electrons gained
4) equal to the total number of protons gained

5. Which ion is most easily reduced?

1) Zn^{2+} 2) Mg^{2+} 3) Co^{2+} 4) Ca^{2+}

6. Given the balanced ionic equation representing a reaction:

$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$

Which half-reaction represents the reduction that occurs?

1) $Al \rightarrow Al^{3+} + 3e^-$
2) $Al^{3+} + 3e^- \rightarrow Al$
3) $Cu \rightarrow Cu^{2+} + 2e^-$
4) $Cu^{2+} + 2e^- \rightarrow Cu$

7. Which equation represents an oxidation-reduction reaction?

1) $H^+ + OH^- \rightarrow H_2O$
2) $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$
3) $Zn + Sn^{4+} \rightarrow Zn^{2+} + Sn^{2+}$
4) $3AgNO_3 + Li_3PO_4 \rightarrow Ag_3PO_4 + 3LiNO_3$

8. Which half-reaction shows conservation of charge?

1) $Cu + e^- \rightarrow Cu^+$
2) $Cu^{2+} + 2e^- \rightarrow Cu$
3) $Cu^+ \rightarrow Cu + e^-$
4) $Cu^{2+} \rightarrow Cu + 2e^-$

9. Which balanced equation represents an oxidation-reduction reaction?

1) $Ba(NO_3)_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaNO_3$
2) $H_3PO_4 + 3KOH \rightarrow K_3PO_4 + 3H_2O$
3) $Fe(s) + S(s) \rightarrow FeS(s)$
4) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

10. Which expression correctly represents a balanced reduction half-reaction?

1) $Na^+ + e^- \rightarrow Na$
2) $Na \rightarrow Na^+ + e^-$
3) $Cl_2 + 2e^- \rightarrow Cl^-$
4) $2\,Cl^- \rightarrow Cl_2 + 2e^-$

Lesson 14.3: Activity Series and Spontaneity

Table J: The table is arranged with the most reactive metals at the top of the table and the least reactive metals at the bottom of the table. What this means is that a metal listed on the table will react with the compound of a metal that is below it. For example, Zn is above Cu on the table. This means that Zn will replace Cu in a compound. Zn will be oxidized by a compound containing Cu. The Cu^{+} ion will be reduced by elemental Zn^{0}. Metals higher on table J are more likely to be oxidized.

The reaction: $Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$ will spontaneously occur.

For each example below, if a reaction will occur spontaneously based on the element positions in the Activity Series, complete the equation and balance it. If there is no reaction, write no reaction.

1. $Ag_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow$
2. $Pb_{(s)} + AuCl_{(aq)} \rightarrow$
3. $Au_{(s)} + LiCl_{(aq)} \rightarrow$
4. $Mg_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow$

Determine if a spontaneous reaction will occur. If one will occur, write the balanced oxidation and reduction half reactions.

1. $Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$
   - Is the reaction spontaneous? _____
   - If yes, write the oxidation ½ reaction:
   - If yes, write the reduction ½ reaction:

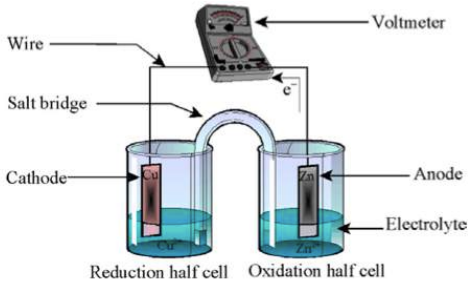
2. $3\ Zn + 2\ Al(NO_3)_3 \rightarrow 2\ Al + 3\ Zn(NO_3)_2$
   - Is the reaction spontaneous? _____
   - If yes, write the oxidation ½ reaction:
   - If yes, write the reduction ½ reaction:

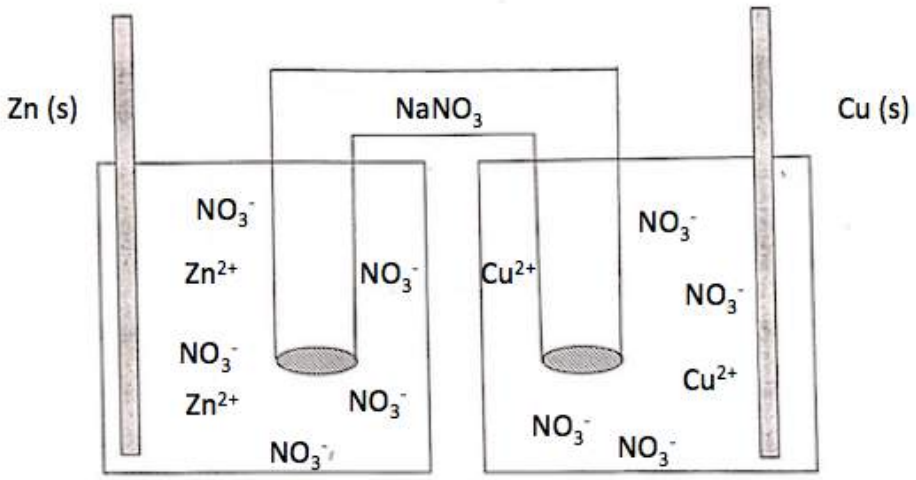
Table J
Activity Series**

Most	Metals	Nonmetals	Most
	Li	F_2	
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	$^{**}H_2$		
	Cu		
	Ag		
	Au		
Least			Least

Video 14.4: Electrochemical Cells

A voltaic cell is an electrochemical cell. A voltaic cell produces electric current as a result of a spontaneous redox reaction, used to make batteries. They consist of two half-cells connected by a salt bridge and two electrodes that connect to a load that uses the electricity produced by the cell.) An example an electrochemical cell is a battery.





Overall Redox Reaction:

Oxidation ½ Reaction:

Species Oxidized: _____

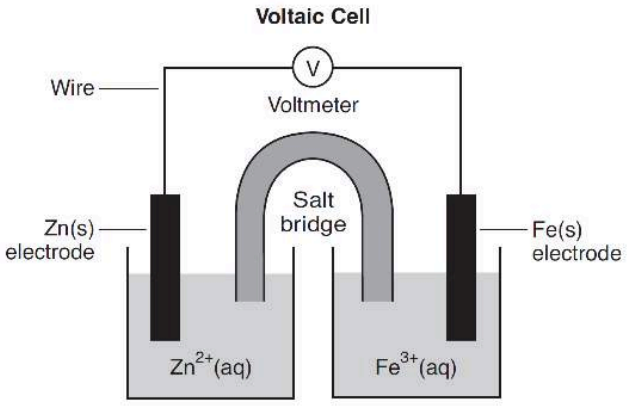
Species Reduced: _____

Reduction ½ Reaction:

Anode: _____

Cathode: _____

An operating voltaic cell has zinc and iron electrodes. The cell and the unbalanced ionic equation representing the reaction that occurs in the cell are shown below.



$$Zn(s) + Fe^{3-}(aq) \rightarrow Zn^{2-}(aq) + Fe(s)$$

1. Write the balanced oxidation half-reaction for this cell.

2. Write the balanced reduction half reaction for this cell.

3. If 6 moles of Zn react, how many moles of electrons will be transferred?

4. Explain, in terms of Zn atoms and Zn ions, why the mass of the Zn electrode decreases as the cell operates.

5. Identify the subatomic particles that flow through the wire as the cell operates.

LESSON 14.4: Electrochemical Cells

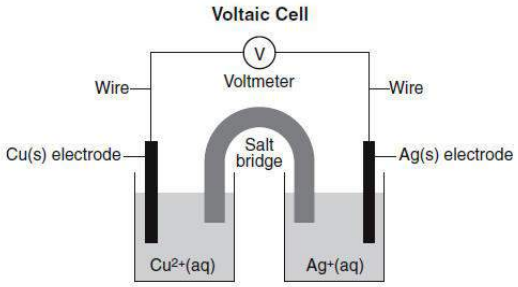
Objective:

- ***Determine the flow of electrons in a battery (voltaic cell)***
- ***Identify the anode and cathode in a voltaic cell***

Voltaic Cells (batteries)

<u>Directions</u>: In each of the following, determine which element oxidized easier on table J (higher up on table J). Then label the ***anode, cathode, direction of e- flow,*** *(remember electrons flow from high to low),* ***which electrode increases and decreases in mass*** *and then* ***write the half reactions in the spaces provided.***

1.

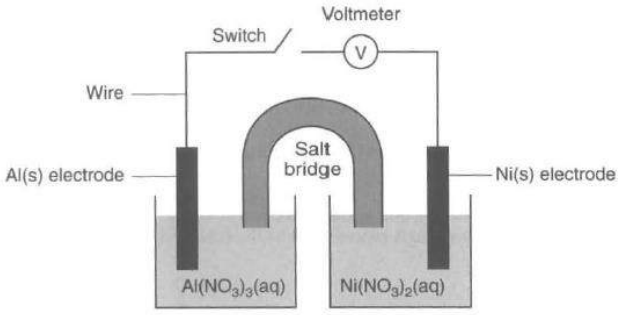


$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

ox:

red:

2.

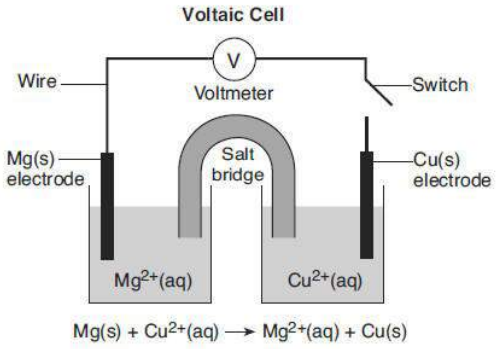


$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$

ox:

red:

3.

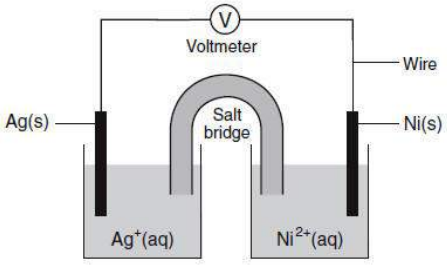


$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$

ox:

red:

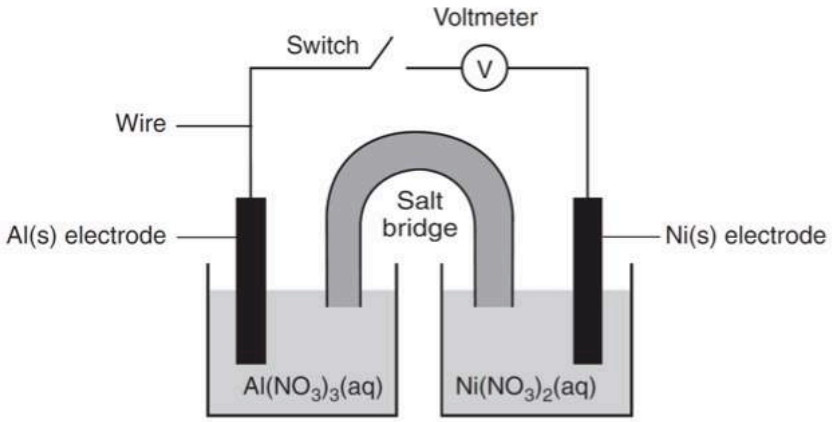
4.



ox:

red:

A student constructs an electrochemical cell during a laboratory investigation. When the switch is closed, electrons flow through the external circuit. The diagram below represents this cell and the reaction that occurs.



$$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$$

1. State, in terms of energy, why this cell is a voltaic cell.

   _____

   _____

2. Determine the number of moles of $Al_{I(s)}$ needed to completely react with 9.0 moles of $Ni^{2+}{}_{(aq)}$ ions.

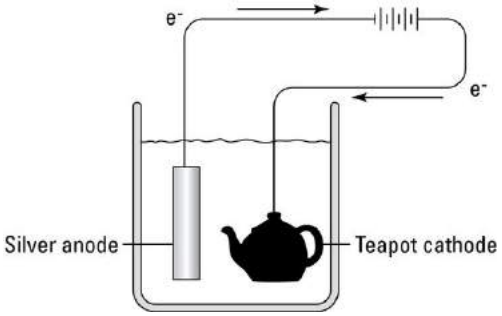
3. Write the balanced half-reaction equation for the oxidation that occurs when the switch is closed.

4. State the direction of electron flow through the wire and when the switch is closed.

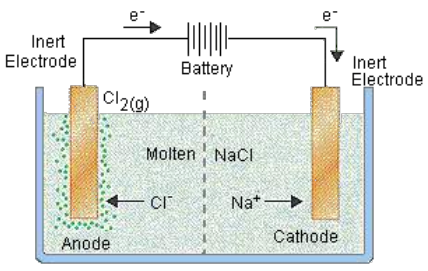
   _____

Lesson 14.5: Electrolytic Cells.

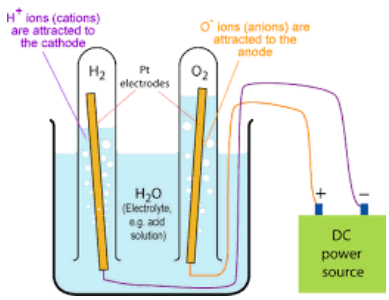
Electrolytic cells use electrical energy to force a nonspontaneous chemical reaction to occur. Reduction occurs at the cathode. The ***cathode*** is the electrode where electrons are sent and is the ***negative electrode***. Oxidation occurs at the anode. The ***anode*** is the electrode where electrons are drawn away from and is the ***positive electrode.***



Electroplating

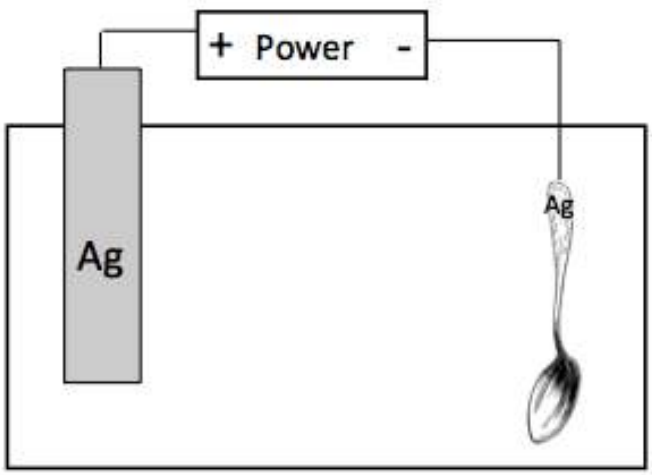


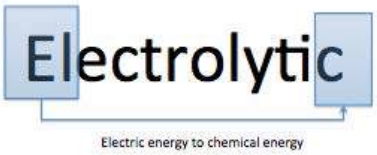
Molten Salt



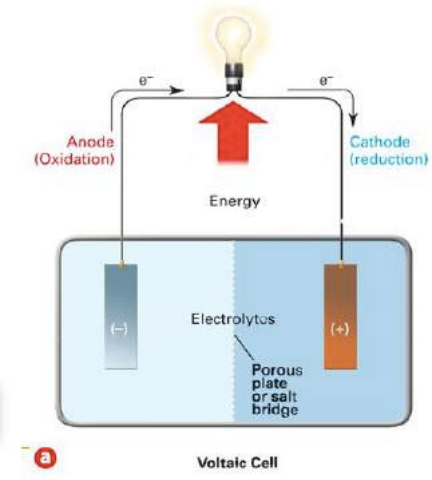
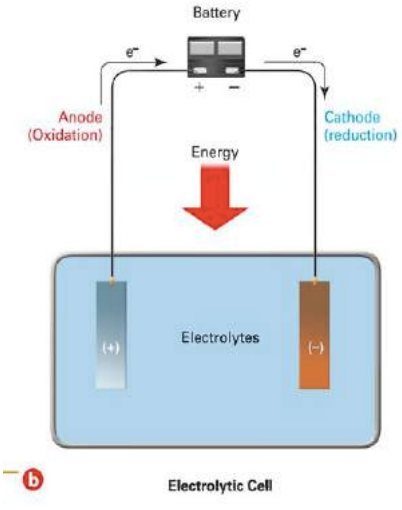
Electrolysis of Water

Electroplating:



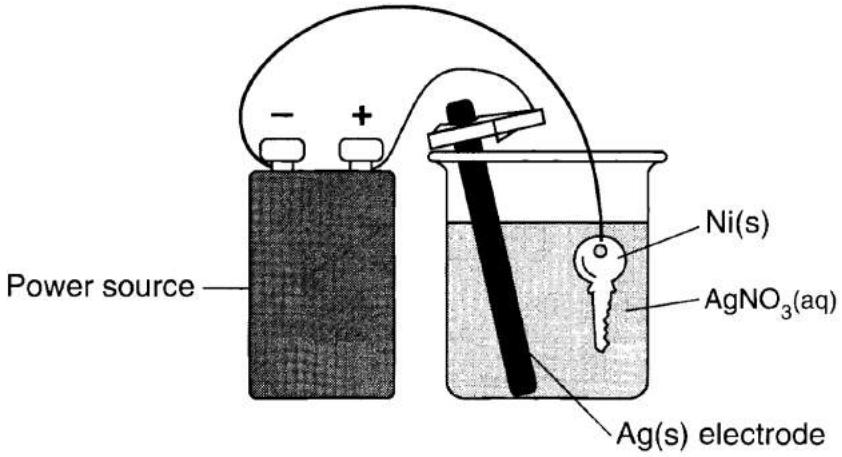


Complete the chart

Electrochemical Cells	
Voltaic Cell	***Electrolytic Cell***
Example:	Example: Electroplating
Redox Reactions	
	2 electrodes
Anode = oxidation Negative (-)	Anode =
Cathode =	Cathode = Reduction Negative (-)
Electrons flow from anode to cathode	
	Use electrical energy for force a nonspontaneous redox reaction to occur (chemical)
	

Base your answers to questions **1** through **3** on the information below.

The diagram below represents an operating electrolytic cell used to plate silver on a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.



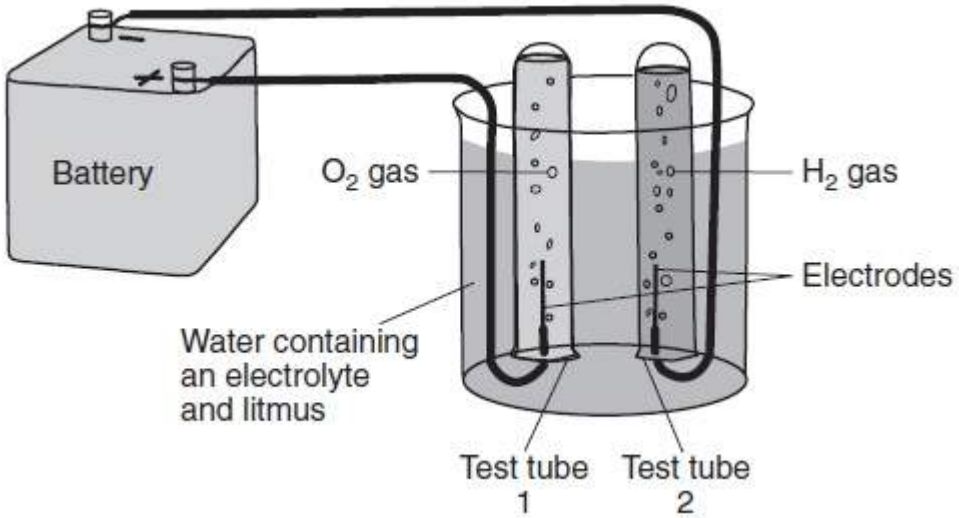
1. Explain, in terms of Ag atoms and $Ag^+(aq)$ ions, why the mass of the silver electrode *decreases* as the cell operates

2. State the purpose of the power source in the cell.

3. Identify the cathode in the cell.

Base your answers to questions **4** and **5** on the information below.

The diagram below shows a system in which water is being decomposed into oxygen gas and hydrogen gas. Litmus is used as an indicator in the water. The litmus turns red in test tube 1 and blue in test tube 2.



The oxidation and reduction occurring in the test tubes are represented by the balanced equations below.

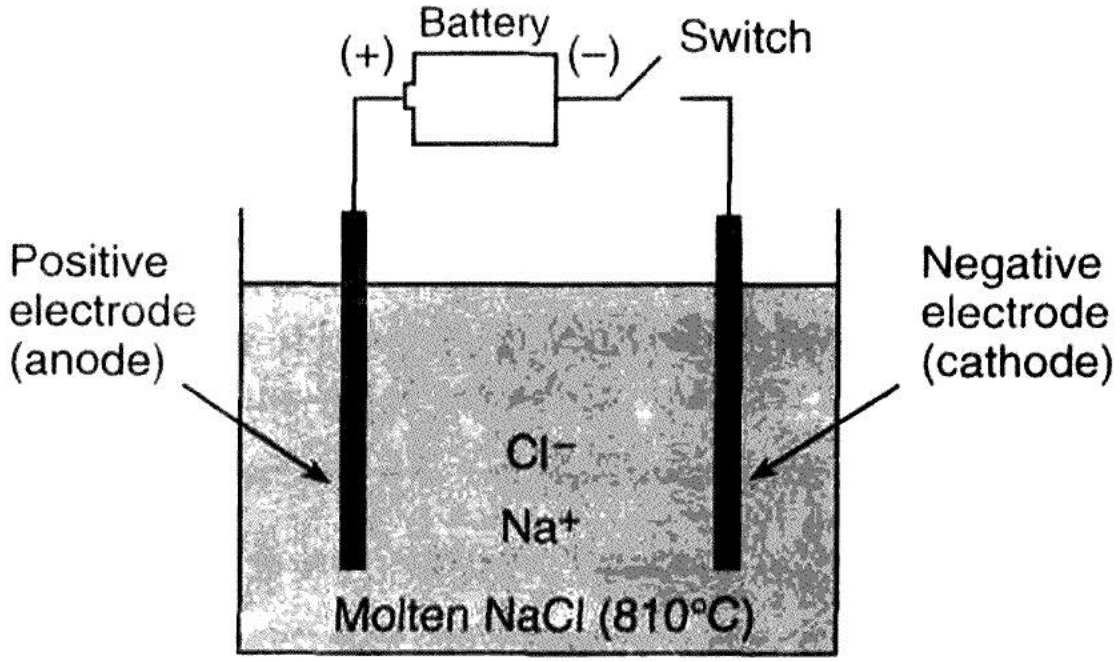
Test tube 1: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Test tube 2: $4H_2O(/) + 4e^- \rightarrow 2H_2(g) + 4OH^-(aq)$

4. Determine the change in oxidation number of oxygen during the reaction in test tube 1.

5. Identify the information in the diagram that indicates this system is an electrolytic cell.

Base your answers to questions **6** through **8** on the diagram and balanced equation below, which represent the electrolysis of molten NaCl.



$$2NaCl \longrightarrow Cl_2 + 2Na$$

6. What is the purpose of the battery in this electrolytic cell?

7. When the switch is closed, which electrode will attract the sodium ions?

8. Write the balanced half-reaction for the reduction that occurs in this electrolytic cell.

Redox Review

1. What is the oxidation state of nitrogen in the compound NH_4Br?

   A) –1 B) +2 C) –3 D) +4

2. Given the balanced equation representing a reaction:

   $$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

   The oxidation state of chlorine in this reaction changes from

   A) –1 to +1 B) –1 to +5
   C) + l to –1 D) +5 to –1

3. During which process does an atom gain one or more electrons?

   A) transmutation B) reduction
   C) oxidation D) neutralization

4. Which half-reaction correctly represents reduction?

   A) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
   B) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
   C) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
   D) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

5. In a redox reaction, the total number of electrons lost is

   A) less than the total number of electrons gained
   B) greater than the total number of electrons gained
   C) equal to the total number of electrons gained
   D) equal to the total number of protons gained

6. Half-reactions can be written to represent all

   A) double-replacement reactions
   B) neutralization reactions
   C) fission and fusion reactions
   D) oxidation and reduction reactions

7. Which half-reaction equation represents the reduction of an iron(II) ion?

   A) $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   B) $Fe^{2+} + 2e^- \rightarrow Fe$
   C) $Fe^{3+} + e^- \rightarrow Fe^{2+}$
   D) $Fe \rightarrow Fe^{2+} + 2e^-$

8. Given the balanced equation representing a reaction:

   $$Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$$

   During this reaction, the oxidation number of Fe changes from

   A) +2 to 0 as electrons are transferred
   B) +2 to 0 as protons are transferred
   C) +3 to 0 as electrons are transferred
   D) +3 to 0 as protons are transferred

9. Given the balanced equation representing a redox reaction:

   $$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$$

   Which statement is true about this reaction?

   A) Each Al loses $2e^-$ and each Cu^{2+} gains $3e^-$.
   B) Each Al loses $3e^-$ and each Cu^{2+} gains $2e^-$.
   C) Each Al^{3+} gains $2e^-$ and each Cu loses $3e^-$.
   D) Each Al^{3+} gains $3e^-$ and each Cu loses $2e^-$.

10. Which half-reaction correctly represents reduction?

    A) $Ag \rightarrow Ag^+ + e^-$
    B) $F_2 \rightarrow 2\ F^- + 2e^-$
    C) $Au^{3+} + 3e^- \rightarrow Au$
    D) $Fe^{2+} + e^- \rightarrow Fe^{3+}$

11. Which balanced equation represents a redox reaction?

    A) $AgNO_3(aq) + NaCI(aq) \rightarrow AgCI(s) + NaNO_3(aq)$
    B) $H_2CO_3(aq) \rightarrow H_2O(\ell) + CO_2(g)$
    C) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(\ell)$
    D) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

12. Which equation represents an oxidation-reduction reaction?

    A) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
    B) $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$
    C) $MgCrO_4 + BaCl_2$ ® $MgC1_2 + BaCrO_4$
    D) $Zn(NO_3)_2 + Na_2CO_3 \rightarrow 2NaNO_3 + ZnCO_3$

13. Which metal is more active than H_2?

A) Ag B) Au C) Cu D) Pb

14. Which reaction occurs spontaneously?

A) $Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaCl(aq)$
B) $Cl_2(g) + 2NaF(aq) \rightarrow F_2(g) + 2NaCl(aq)$
C) $I_2(s) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaI(aq)$
D) $I_2(s) + 2NaF(aq) \rightarrow F_2(g) + 2NaI(aq)$

15. Given the unbalanced ionic equation:

$$3Mg + \_\_Fe^{3+} \rightarrow 3Mg^{2+} + \_\_Fe$$

When this equation is balanced, both Fe^{3+} and Fe have a coefficient of

A) 1, because a total of 6 electrons is transferred
B) 2, because a total of 6 electrons is transferred
C) 1, because a total of 3 electrons is transferred
D) 2, because a total of 3 electrons is transferred

16. Which half-reaction shows conservation of charge?

A) $Cu + e^- \rightarrow Cu^+$
B) $Cu^{2+} + 2e^- \rightarrow Cu$
C) $Cu^+ \rightarrow Cu + e^-$
D) $Cu^{2+} \rightarrow Cu + 2e^-$

17. Given the balanced equation representing a reaction:

$$2Fe + 3Cu^{2+} \rightarrow 2Fe^{3+} + 3Cu$$

When the iron atoms lose six moles of electrons, how many moles of electrons are gained by the copper ions?

A) 12 moles B) 2 moles
C) 3 moles D) 6 moles

18. A voltaic cell spontaneously converts chemical energy to

A) electrical energy
B) geothermal energy
C) mechanical energy
D) nuclear energy

19. In a voltaic cell, chemical energy is converted to

A) electrical energy, spontaneously
B) electrical energy, non-spontaneously
C) nuclear energy, spontaneously
D) nuclear energy, non-spontaneously

20. Which half-reaction can occur at the anode in a voltaic cell?

A) $Ni^{2+} + 2e^- \rightarrow Ni$
B) $Sn + 2e^- \rightarrow Sn^{2+}$
C) $Zn \rightarrow Zn^{2+} + 2e^-$
D) $Fe^{3+} \rightarrow Fe^{2+} + e^-$

21. Given the balanced ionic equation representing the reaction in an operating voltaic cell:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

The flow of electrons through the external circuit in this cell is from the

A) Cu anode to the Zn cathode
B) Cu cathode to the Zn anode
C) Zn anode to the Cu cathode
D) Zn cathode to the Cu anode

22. Reduction occurs at the cathode in

A) electrolytic cells, only
B) voltaic cells, only
C) both electrolytic cells and voltaic cells
D) neither electrolytic cells nor voltaic cells

23. Given the balanced equation representing the reaction occurring in a voltaic cell:

$$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$$

In the completed external circuit, the electrons flow from

A) Pb(s) to Zn(s)
B) $Pb^{2+}(aq)$ to $Zn^{2+}(aq)$
C) Zn(s) to Pb(s)
D) $Zn^{2+}(aq)$ to $Pb^{2+}(aq)$

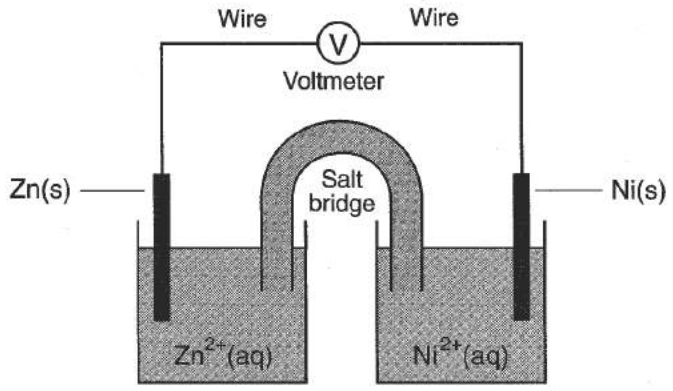
24. Which statement is true about oxidation and reduction in an electrochemical cell?

A) Both occur at the anode.
B) Both occur at the cathode.
C) Oxidation occurs at the anode and reduction occurs at the cathode.
D) Oxidation occurs at the cathode and reduction occurs at the anode.

25. Which statement describes one characteristic of an operating electrolytic cell?

A) It produces electrical energy.
B) It requires an external energy source.
C) It uses radioactive nuclides.
D) It undergoes a spontaneous redox reaction.

26. The diagram below represents an operating electrochemical cell and the balanced ionic equation for the reaction occurring in the cell.



$Zn(s) + Ni^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Ni(s)$

Which statement identifies the part of the cell that conducts electrons and describes the direction of electron flow as the cell operates?

A) Electrons flow through the salt bridge from the Ni(s) to the Zn(s).
B) Electrons flow through the salt bridge from the Zn(s) to the Ni(s).
C) Electrons flow through the wire from the Ni(s) to the Zn(s).
D) Electrons flow through the wire from the Zn(s) to the Ni(s).

27. Which statement describes electrolysis?

A) Chemical energy is used to produce an electrical change.
B) Chemical energy is used to produce a thermal change.
C) Electrical energy is used to produce a chemical change.
D) Thermal energy is used to produce a chemical change.

28. Given the balanced equation representing a reaction occurring in an electrolytic cell:

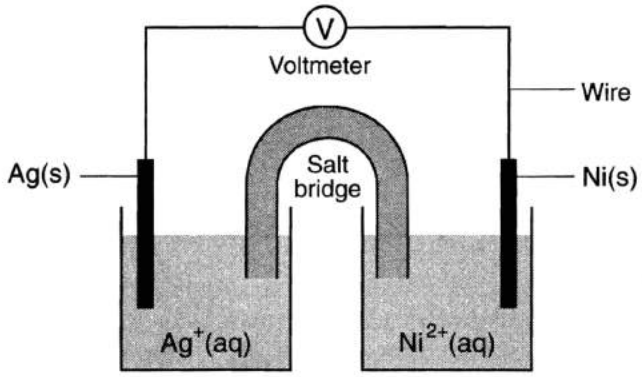
$$2NaCl(\ell) \rightarrow 2Na(\ell) + Cl_2(g)$$

Where is $Na(\ell)$ produced in the cell?

A) at the anode, where oxidation occurs
B) at the anode, where reduction occurs
C) at the cathode, where oxidation occurs
D) at the cathode, where reduction occurs

Base your answers to questions **29** through **31** on the information below.

The diagram below represents an operating voltaic cell at 298 K and 1.0 atmosphere in a laboratory investigation. The reaction occurring in the cell is represented by the balanced ionic equation below.



$$2Ag^+(aq) + Ni(s) \longrightarrow 2Ag(s) + Ni^{2+}(aq)$$

29. Determine the total number of moles of $Ni^{2+}(aq)$ ions produced when 4.0 moles of $Ag^+(aq)$ ions completely react in this cell

30. Identify the anode in this cell.

31. Write a balanced half-reaction equation for the reduction that occurs in this cell.

32. Base your answer to the following question on the information below.

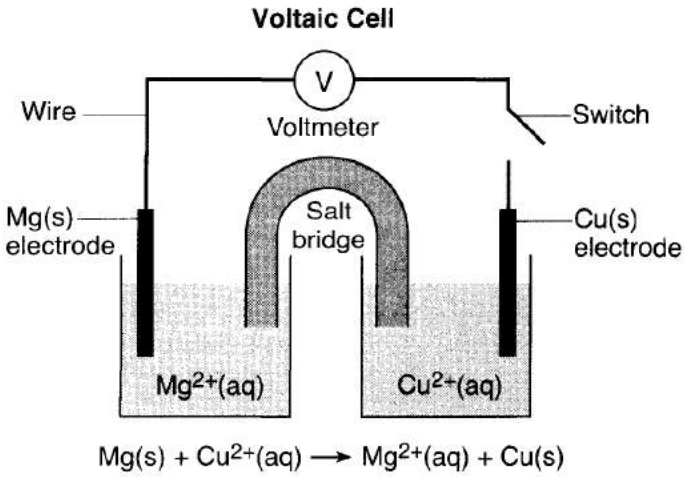
In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.

$$Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

Write a balanced half-reaction equation for the oxidation that occurs.

Base your answers to questions **33** through **35** on the information below.

A voltaic cell with magnesium and copper electrodes is shown in the diagram below. The copper electrode has a mass of 15.0 grams.



When the switch is closed, the reaction in the cell begins. The balanced ionic equation for the reaction in the cell is shown below the cell diagram. After several hours, the copper electrode is removed, rinsed with water, and dried. At this time, the mass of the copper electrode is greater than 15.0 grams.

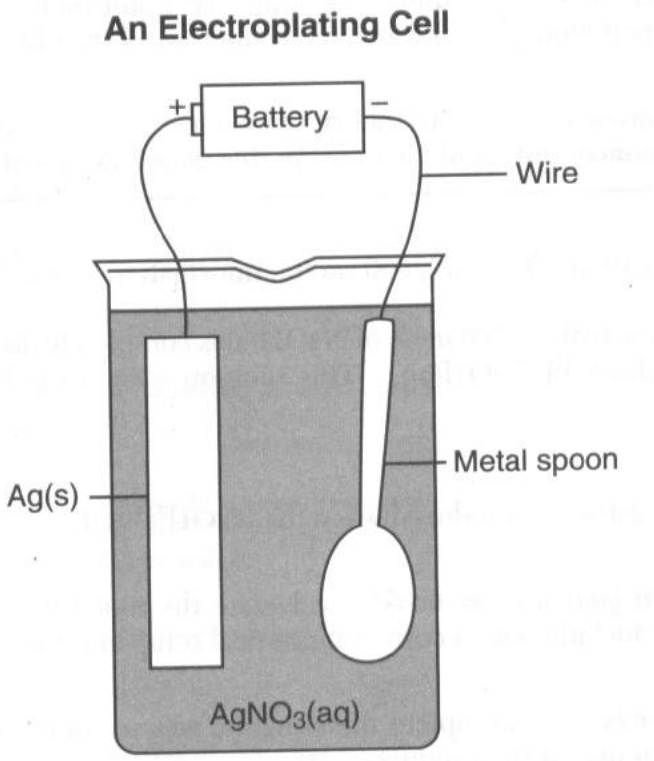
33. State the purpose of the salt bridge in this cell.

34. State the directions of electron flow through the wire between the electrodes when the switch is closed.

35. Explain, in terms of copper ions and copper atoms, why the mass of the copper electrode increases as the cell operates. Your response must include information about *both* copper ions and copper atoms.

36. Base your answer to the following question on the information below.

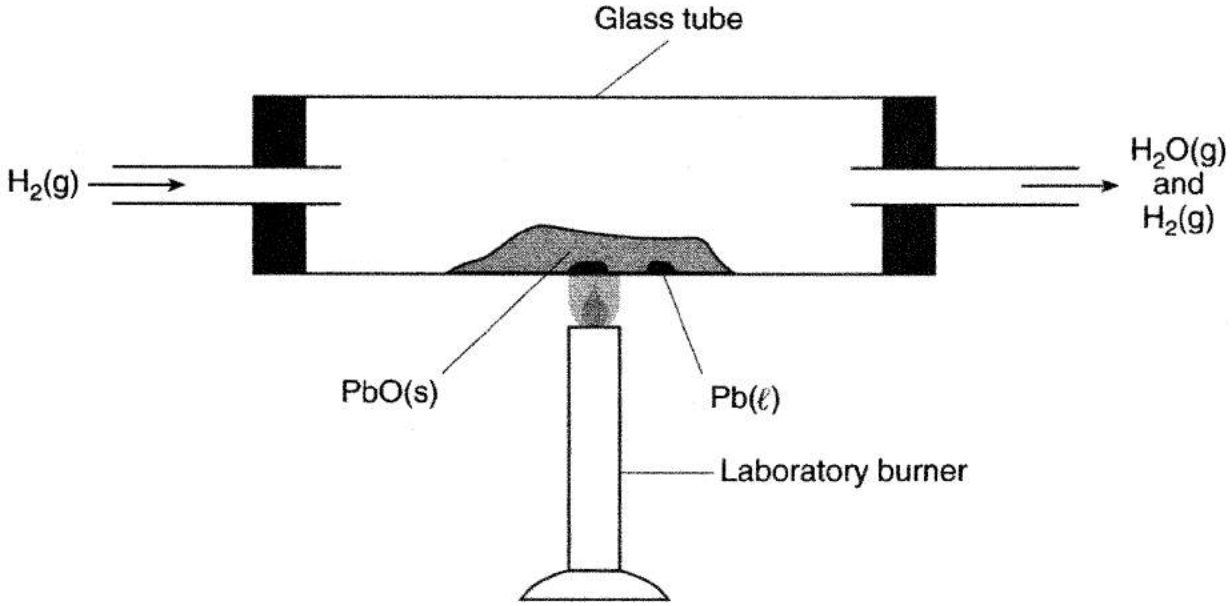
Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The diagram below shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in $AgNO_3(aq)$.



Explain the purpose of the battery in this cell.

Base your answers to questions **37** through **40** on the information below and on your knowledge of chemistry.

In a laboratory apparatus, a sample of lead(II) oxide reacts with hydrogen gas at high temperature. The products of t his reaction are liquid lead and water vapor. As the reaction proceeds, water vapor and excess hydrogen gas leave the glass tube. The diagram and balanced equation below represent this reaction.



$$PbO(s) + H_2(g) + \text{heat} \longrightarrow Pb(\ell) + H_2O(g)$$

37. State *one* change in reaction conditions, other than adding a catalyst, that would cause the rate of this reaction to increase.

38. Explain why the reaction that occurs in this glass tube can *not* reach equilibrium.

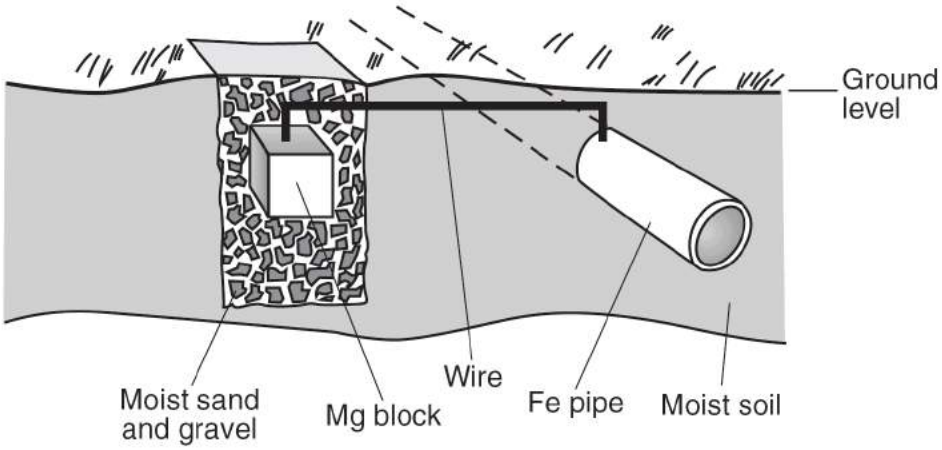
39. Write a balanced half-reaction equation for the reduction of the Pb^{2+} ions in this reaction.

40. Determine the change in oxidation number for the hydrogen that reacts.

Base your answers to questions **41** and **42** on the information below.

Underground iron pipes in contact with moist soil are likely to corrode. This corrosion can be prevented by applying the principles of electrochemistry. Connecting an iron pipe to a magnesium block with a wire creates an electrochemical cell. The magnesium block acts as the anode and the iron pipe acts as the cathode. A diagram of this system is shown below.

Cross-Sectional View of Underground Pipe Protection System

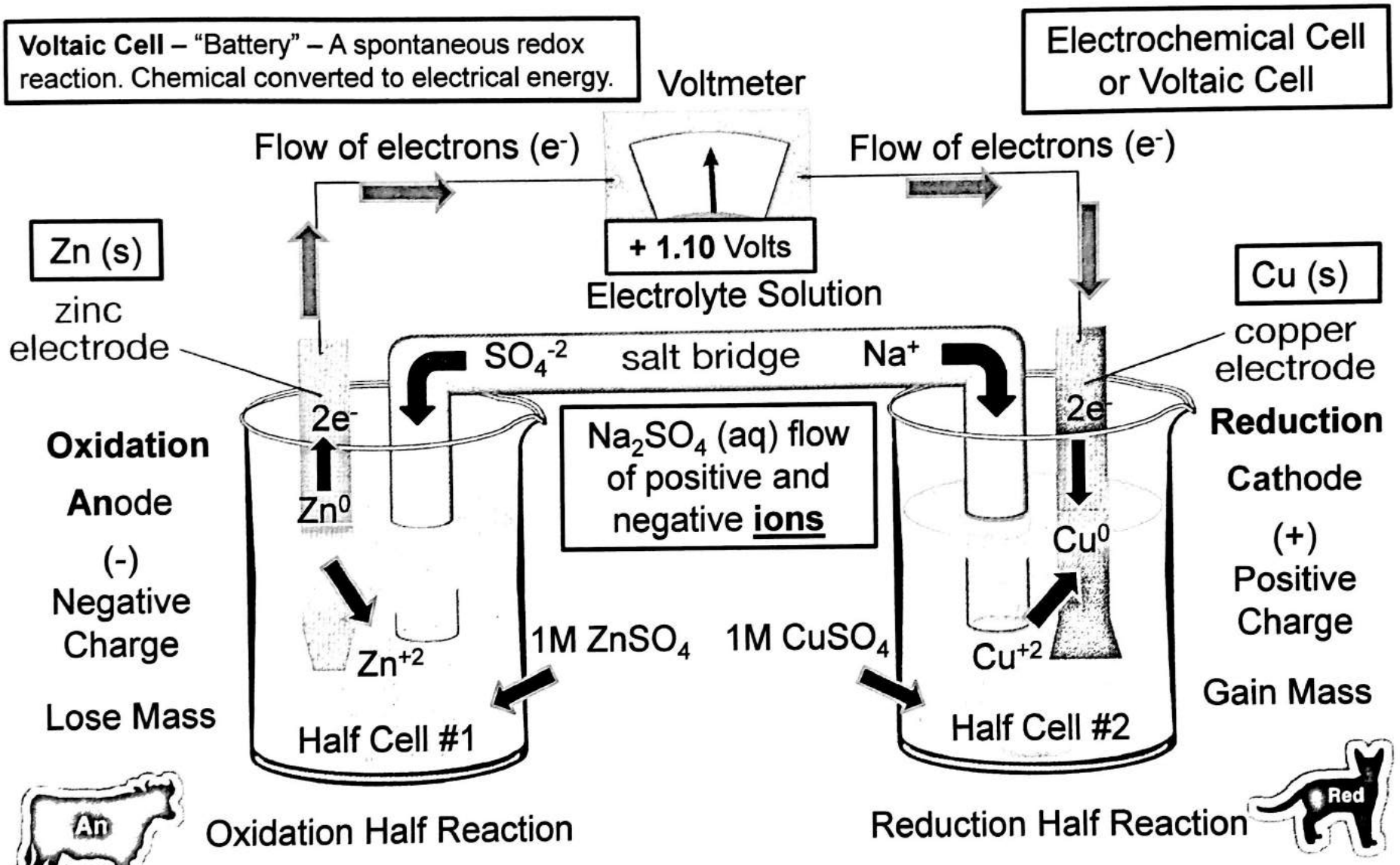


41. Explain, in terms of reactivity, why magnesium is preferred over zinc to protect underground iron pipes. Your response must include *both* magnesium and zinc.

42. State the direction of the flow of electrons between the electrodes in this cell.

Voltaic Cell – "Battery" – A spontaneous redox reaction. Chemical converted to electrical energy.

Electrochemical Cell or Voltaic Cell



Oxidation Half Reaction

Zn^0 (s) → Zn^{+2} (aq) + $2e^-$

E^0 = + 0.76

Reduction Half Reaction

Cu^{+2} (aq) + $2e^-$ → Cu^0 (s)

E^0 = + 0.34

Net ionic reaction: Zn^0 (s) + Cu^{+2} (aq) → Zn^{+2} (aq) + Cu^0 (s)