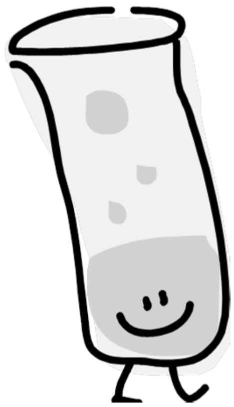


Subject Test Chemistry

Strategy Notes

APPLYME

Test Basics

- Scoring: 200 – 800
- Timing: 60 minutes
- Questions: 85 multiple choice (3 types)
- 42 seconds per question!
- All questions are worth the same amount
  - Right: +1 point
  - Blank: 0 points
  - Wrong: -1/4 point
- NO calculator allowed

Question Types

- Part A: Classification questions (20-24)
  - 5 answer choices followed by 3 to 5 questions
  - Answer choices may be used once, more than once, or not at all

Directions: Each set of lettered choices below refers to the numbered questions or statements immediately following it. Select the one lettered choice that best answers each question or best fits each statement and then fill in the corresponding oval on the answer sheet. A choice may be used once, more than once, or not at all in each set.

Questions 1–3

(A) Zinc
(B) Iron
(C) Helium
(D) Copper
(E) Fluorine

1. A highly electronegative element
2. Forms colored solutions when dissolved in water
3. Normally exists as a diatomic molecule but can react to form a 2- ion

Question Types

- Part B: Relationship-Analysis questions (16-17)
  - Determine whether Statement I and Statement II are true or false
  - Decide whether the 1nd statement is the reason for the 1st statement being true
  - Numbered starting with 101

Question Types

<u>Directions:</u> Each question below consists of two statements, statement I in the left-hand column and statement II in the right-hand column. For each question, determine whether statement I is true or false <u>and</u> whether statement II is true or false and fill in the corresponding T or F ovals on your answer sheet. <u>Fill in oval CE only if statement II is a correct explanation of statement I.</u>

	Statement I		Statement II
101.	A 1.0 M solution of HCl has a low pH.	BECAUSE	HCl contains chlorine.
102.	An atom of chlorine is smaller than an atom of sulfur.	BECAUSE	Chlorine has a greater effective nuclear charge than sulfur.

Question Types

- Part C: Five Choice Completion questions

<u>Directions:</u> Each of the questions or incomplete statements below is followed by five suggested answers or completions. For each question, select the one choice that is the best answer to the question and then fill in the corresponding oval on the answer sheet.

24. Which of the following molecules does not match its geometric shape?

(A)	BF3	trigonal planar
(B)	CHCl3	tetrahedral
(C)	H2O	V shape (bent)
(D)	CO2	linear
(E)	PCl3	trigonal planar

General Strategies

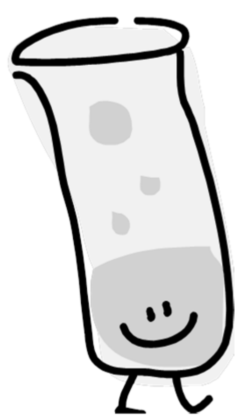
- Pacing: Focus your time on the questions with familiar concepts
- Double pass: Not sure, skip it and come back later
- Use POE: Eliminate 2 or more, then GUESS
- If you're doing lots of math, you're doing something wrong

Topics

- Chemistry Basics
- The Structure of Matter
- The States of Matter
- Reaction Types
- Stoichiometry
- Equilibrium and Reaction Rates
- Thermodynamics
- Descriptive Chemistry
- Laboratory

Barron’s Topics

1. Introduction to Chemistry
2. Atomic Structure and the Periodic Table of the Elements
3. Bonding
4. Chemical Formulas
5. Gases and the Gas Laws
6. Stoichiometry (Chemical Calculations) and the Mole Concept
7. Liquids, Solids, and Phase Changes
8. Chemical Reactions and Thermochemistry
9. Rates of Chemical Reactions
10. Chemical Equilibrium
11. Acids, Bases, and Salts
12. Oxidation-Reduction
13. Some Representative Groups and Families
14. Carbon and Organic Chemistry
15. The Laboratory



Subject Test Chemistry

Chemistry Basics

Strategy Notes

APPLYME

Naming Simple Compounds

- Know your polyatomic ions.
- If the compound **starts with H**, it is an acid. Use the **naming acids** rules.
- If the compound **starts with C** and contains quite a few H's and perhaps some O's, it is organic. Use the **naming organic compounds** rules.
- If the compound **starts with a metal**, it is most likely ionic. Use the **naming binary ionic compounds** rules.
- If the compound **starts with a nonmetal other than H or C**, use the **naming binary molecular compounds** rules.

Naming Polyatomic Ions

Name of polyatomic ion	Formula and charge
Ammonium ion	
	$C_2H_3O_2^-$
Cyanide ion	
	OH^-
Nitrate ion	NO_3^-
	ClO_3^-
Sulfate ion	
Carbonate ion	
	PO_4^{3-}

Naming Acids

Acid formula	Acid name
HCl	
	Hypochlorous acid
	Chlorous acid
	Chloric acid
	Hyperchloric acid (or perchloric acid)
HNO_3	
	Hydrobromic acid
H_3PO_4	
H_3PO_3	
	Hydrocyanic acid
$HC_2H_3O_2$	
	Carbonic acid
	Hydroiodic acid
HF	

Naming Organic Compounds

No. of carbon atoms = *n*	Prefix or stem	*-ane* C_nH_{2n+2}	*-ene* C_nH_{2n}	*-yne* C_nH_{2n-2}	*-anol* $C_nH_{2n+1} + OH$
1	*meth-*		Must have 2 carbons		CH_3OH
2	*eth-*				
3	*prop-*		C_3H_6		
4	*but-*				
5	*pent-*	C_5H_{12}			
6	*hex-*				
7	*hept-*				$C_7H_{15}OH$
8	*oct-*			C_8H_{14}	
9	*non-*				
10	*dec-*				

Naming Binary Ionic Compounds

- The positive ion name is given *first* (remember, if it's a transition metal, the Roman numeral indicating its charge is part of its name), followed by the name of the negative ion.
- *No* prefixes are used.

Naming Binary Molecular Compounds

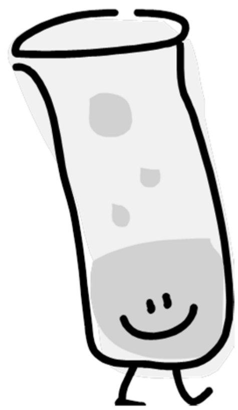
Subscript	Prefix
1	*mono-* (usually used only on the second element, such as carbon monoxide or nitrogen monoxide)
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*
6	*hexa-*
7	*hepta-*
8	*octa-*
9	*nona-*
10	*deca-*

Writing Chemical Formulas

Ag	Si^{2+}	Cu^{+}	Ba	NH_4	P^{5+}	Mn^{7+}
N						
O						
Br						
S						
SO_4						
ClO_2						
PO_3						

Writing Chemical Names

Ag	Si^{2+}	Cu^{+}	Ba	NH_4	P^{5+}	Mn^{7+}
N						
O						
Br						
S						
SO_4						
ClO_2						
PO_3						



Subject Test Chemistry

Structure of Matter

Strategy Notes

APPLYME

Atomic Structure

Example:

The atomic number of a certain element is 11, and its atomic mass number is 23. How many protons and neutrons does this atom have, and what is its chemical symbol?

Nuclear Reactions

- **Radioactivity** is the spontaneous disintegration of an unstable atomic nucleus and the subsequent emission of radiation.
- Types of radioactive decay
  - Alpha decay: results in a decrease in the atomic number by 2 and a decrease in the atomic mass by 4
  - Beta decay: causes an increase in the atomic number by 1 but no change in mass number
  - Gamma decay: consists of the emission of pure electromagnetic energy; no particles are emitted during this process; do not affect charge or mass
  - Positron emission: converts a proton into a neutron; the positron is emitted and the neutron remains behind in the nucleus, decreasing the atomic number by 1

Nuclear Reactions

- Alpha decay

$$^{234}_{92}\text{U} \rightarrow {}^{4}_{2}\text{He} + {}^{230}_{90}\text{Th}$$

- Beta decay

$$^{233}_{92}\text{U} \rightarrow {}^{0}_{-1}\beta + {}^{233}_{93}\text{Np}$$

- Gamma decay: not included in reactions
- Positron emission

$$p \rightarrow {}^{1}_{-0}n + {}^{0}_{1}e$$

Nuclear Reactions

Example:

Complete the balanced equation by determining the missing term.

$$^{80}_{33}\text{As} \rightarrow {}^{0}_{-1}\beta + ?$$

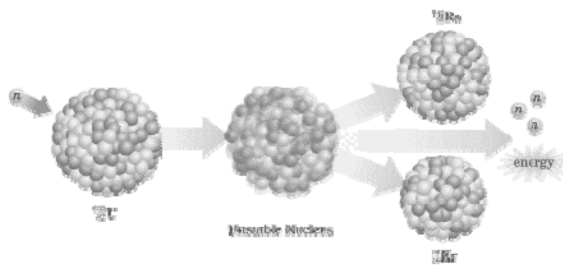
Nuclear Reactions

Example:

Write the equation for the alpha decay of radium-221.Write the equation for the beta decay of sulfur-35.

Nuclear Reactions

- Types of nuclear reactions
  - Fusion reactions: two light nuclei are combined to form a heavier, more stable nucleus.
  - Fission reactions: a heavy nucleus is split into two nuclei with smaller mass numbers.



Nuclear Reactions

Example:

Is the following process an example of fission or fusion?

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}n \rightarrow {}^{146}_{57}\mathrm{La} + {}^{87}_{35}\mathrm{Ba} + 3\,{}^{1}_{0}n$$

Half-Lives

Example:

A radioactive substance has a half-life of 20 minutes. If we begin with a 500 g sample, how much of the original sample remains after two hours?

Quantum Mechanical Model of Atom

- Max Planck determined that energy is released and absorbed by atoms in certain fixed amounts known as **quanta.**
- Albert Einstein determined that radiant energy is also quantized—he called the discrete energy packets **photons**. Einstein's theory was that electromagnetic radiation (light, for example) has characteristics of both a wave and a stream of particles.

Quantum Mechanical Model of Atom

- Niels Bohr proposed that the neutrons and protons are contained in a small, dense nucleus, which the electrons orbit in defined spherical orbits.
- He referred to these orbits as "shells" or "energy levels" and designated each by an integer.
- Bohr theorized that energy in the form of photons must be absorbed in order for an electron to move from a lower energy level to a higher one, and is emitted when an electron travels from a higher energy level to a lower one.
- In the Bohr model, the lowest energy state available for an electron is the **ground state**, and all higher-energy states are **excited states**.

Quantum Mechanical Model of Atom

- Werner Heisenberg put forth his **uncertainty principle**, which states that, at any one time, it is impossible to calculate both the momentum and the location of an electron in an atom; it is only possible to calculate the **probability** of finding an electron within a given space.
- According to the **Pauli exclusion principle**, *no two electrons in an atom can have the same set of four quantum numbers*. This means no atomic orbital can contain more than *two* electrons, and if the orbital does contain two electrons, they must be of opposite spin.

Quantum Mechanical Model of Atom

Principal quantum number (n)	Has positive values of 1, 2, 3, etc. As n increases, the orbital becomes larger—this means that the electron has a higher energy level and is less tightly bound to the nucleus.
Second quantum number or azimuthal quantum number (l)	Has values from 0 to n – 1. This defines the shape of the orbital, and the value of l is designated by the letters *s*, *p*, *d*, and *f*, which correspond to values for l of 0, 1, 2, and 3. In other words, if the value of l is 0, it is expressed as *s*; if l = 1 = *p*, l = 2 = *d*, and l = 3 = *f*.
Magnetic quantum number (m_l)	Determines the orientation of the orbital in space relative to the other orbitals in the atom. This quantum number has values from -l through 0 to +l.
Spin quantum number (m_s)	Specifies the value for the spin and is either +1/2 or -1/2. No more than two electrons can occupy any one orbital. In order for two electrons to occupy the same orbital, they must have opposite spins.

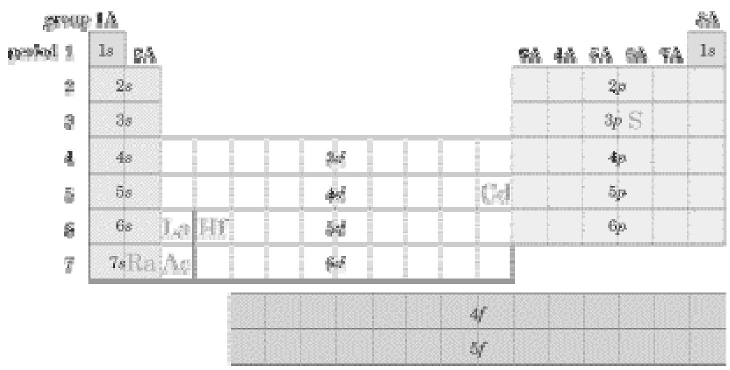
Electron Configurations

- **Hund's rule** states that the most stable arrangement of electrons is that which allows the maximum number of unpaired electrons.

Energy level principal quantum number, *n*	Number of sublevels	Names of sublevels
1	1	*s*
2	2	*s*, *p*
3	3	*s*, *p*, *d*
4	4	*s*, *p*, *d*, *f*

Electron Configurations

Sublevel	*s*	*p*	*d*	*f*
Number of orbitals	1	3	5	7
Maximum number of electrons	2	6	10	14
Quantum number, *l*	0	1	2	3



Electron Configurations

Example:

Using the periodic table, determine the electron configuration for sulfur.

Orbital Notation

Example:

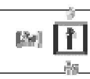
Determine the orbital notation of sulfur.

Orbital Notation

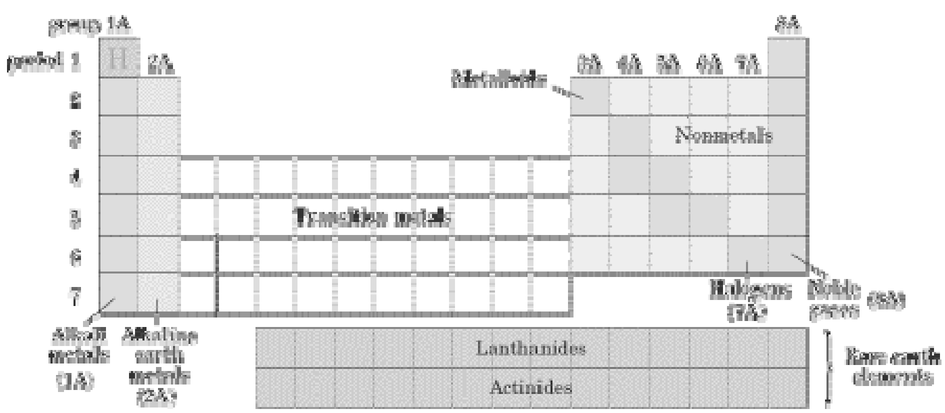
Example:

Which element has this set of quantum numbers: n = 5, l = 1, m_l = -1, and m_s = -1/2?

Electron Configurations

Element	Valence electron configuration	Valence orbital notation	Set of quantum numbers
			
	[Ar] $3d^6$		
			
			5, 1, 0, +1/2
	$4p^5$		
			6, 0, 0, -1/2

Periodic Table and Periodic Properties



Metals

- Malleable, ductile, and have luster
- Oxidize (rust and tarnish) readily and form *positive* ions (cations)
- Excellent conductors of both heat and electricity
- Types of metals
  - Transition metals: known for their ability to refract light as a result of their unpaired electrons; ionic solutions of these metals are usually colored, so these metals are often used in pigments.
  - Rare earth metals (actinides and lanthanides): fill the *f* orbitals; rarely found in nature; uranium is the last naturally occurring element; the rest are man-made

Non-Metals

- Do not conduct electricity well because they do not have free electrons
- All the elemental gases are included in the nonmetals
- Notice that hydrogen is placed with the metals because it has only one valence electron, but it is a nonmetal.

Specific Families

- **Alkali metals (1A)**—The most reactive metal family, these must be stored under oil because they react violently with water! They dissolve and create an alkaline, or basic, solution, hence their name.
- **Alkaline earth metals (2A)**—These also are reactive metals, but they don't explode in water; pastes of these are used in batteries.
- **Halogens (7A)**—Known as the "salt formers," they are used in modern lighting and always exist as diatomic molecules in their elemental form.
- **Noble gases (8A)**—Known for their extremely slow reactivity, these were once thought to never react; neon, one of the noble gases, is used to make bright signs.

Atomic Radius

- Atomic radii decrease moving across a period from left to right

Example:

Which ion is larger, F^- or O^{2-}?

Ionization Energy

- The energy required to remove an electron from the atom in the gas phase
- Ionization energy increases as we move across a period
- Ionization energy decreases as you move down a group or family

Example:

Which of the following elements has the highest ionization energy: K, Ca, Ga, As, or Se?

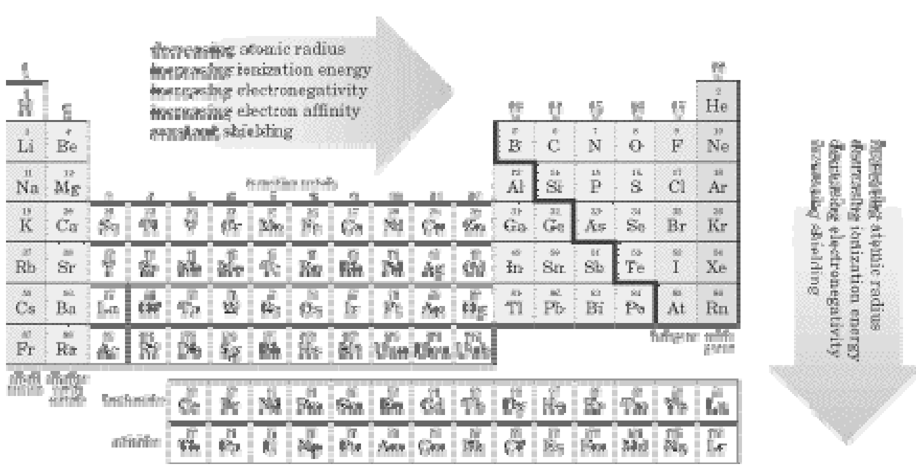
Electron Affinity

- The amount of energy released when an electron is added to the atom in its gaseous state—when an electron is added to an atom, the atom forms a negative ion
- Electron affinity becomes more negative as we move across a period

Electronegativity

- A measure of the attraction an atom has for electrons when it is involved in a chemical bond
- Elements that have high ionization energy and high electron affinity will also have high electronegativity since their nuclei strongly attract electrons
- Electronegativity increases from left to right as we move across a period and decreases as we move down any group or family

Periodic Table Summary



Chemical Bonds

- Strong attractive forces that hold the atoms in most molecules together
- Determine physical properties of molecules
  - Melting point
  - Hardness
  - Electrical and thermal conductivity
  - Solubility
- Bonds formed by interacting valence (outer) electrons
- Bonding occurs according to the Octet Rule

Octet Rule

- All noble gases have 8 valence electrons (s^2p^6) and are chemically stable
- An atom tends to bond with other atoms until it has 8 electrons in its outermost shell, thereby forming a stable electron configuration similar to that of the noble gas elements

Exceptions to Octet Rule

Element	Exception
Hydrogen	Can only have 2 valence electrons
Lithium	Bonds to attain 2 valence electrons
Beryllium	Bonds to attain 4 valence electrons
Boron	Bonds to attain 6 valence electrons
Phosphorus	More than 8 with *d* orbitals
Sulfur	More than 8 with *d* orbitals

- Atoms with fewer than eight electrons (e.g., H)
- Elements in periods 4 and higher; can be surrounded by more than four valence pairs in certain compounds

Types of Bonds

- Ionic bonds: usually form between metals and nonmetals
- Covalent bonds: usually between 2 nonmetals
- Metallic bonds: exist only in metals, such as aluminum, gold, copper, and iron

Ionic Bonds

- Result of an electrostatic attraction between ions that have opposite charges (cations (+) and anions (-))
- Elements that participate in ionic bonds are often from opposite ends of the periodic table and have an electronegativity difference greater than 1.67 (Group I and II → Group VII)
- An electron is actually *transferred* from the less electronegative atom to the more electronegative element
- Example: NaCl (table salt)

Ionic Compounds

- Strong ionic bonds
- Exist in solid state under standard conditions (hard and brittle)
- High melting and boiling points
- Conduct electricity in liquid and aqueous states but not as a solid
- Form crystals

Ionic Compounds

- Sodium (Na): Silvery gray metal composed of neutral atoms which react vigorously with water or air
- Chlorine (Cl): Neutral greenish-yellow, poisonous, diatomic gas
- Sodium chloride (NaCl): Common white crystalline table salt
  - Melting point: 801°C
- Magnesium (Mg): Solid, slightly shiny metal
- Oxygen (O): Neutral, clear, diatomic gas
- Magnesium oxide (MgO): Fine white powder
  - Melting point: 2800°C

Ionic Compounds

Example:

Write the correct name for the following ionic compounds: MgO, FeS, LiH, H_2S.

Covalent Bonds

- Form when electrons are *shared* between atoms rather than transferred from one atom to another (to achieve a noble gas electron configuration)
- Sharing rarely occurs equally because of course no two atoms have the same electronegativity value
- Nonpolar: electronegativity difference between the two atoms involved falls between 0 and 0.4
- Polar: electronegativity difference falls between 0.4 and 1.67
- Element with the higher electronegativity attracts the electron pair more strongly

Types of Covalent Bonds

- Single bond: one pair of shared electrons
  - One sigma (σ) bond: electron density concentrated along the line that represents the bond joining the two atoms
- Double bond: two pairs of shared electrons
  - One σ and a pi (π) bond
  - Pi bond: electron density concentrated above and below the line that represents the bond joining the two atoms
- Triple bond: three pairs of shared electrons
  - One σ and two π bonds

Types of Covalent Bonds

- Characterized by
  - Bond length
    - Average distance between the 2 nuclei of the atoms involved in the bond
    - As the number of shared electrons increases, the bond length decreases
    - Triple < double < single
  - Bond energy
    - Energy required to separate two bonded atoms
    - As the number of shared electrons increases, the bond energy (strength) decreases
    - Triple > double > single

Covalent Compounds

- Held together by weak Van der Waal's forces
- Either gases, volatile liquids, or soft solids
- Low melting and boiling points
- Do not dissolve in water, only in organic solvents
- Bad conductors of electricity
- Low density (less dense than water)

Covalent Compounds

- Hydrogen (H): diatomic clear gas at room temperature
- Oxygen (O): diatomic clear gas at room temperature
- Water (H_2O): odorless colorless liquid at room temperature
- Nitrogen (N): a colorless, odorless and tasteless gas; makes up around 78% of the air you breathe
- Ammonia (NH_3): colorless gas with a characteristic pungent smell; easily liquefied due to the hydrogen bonding between molecules

Covalent Compounds

Example:

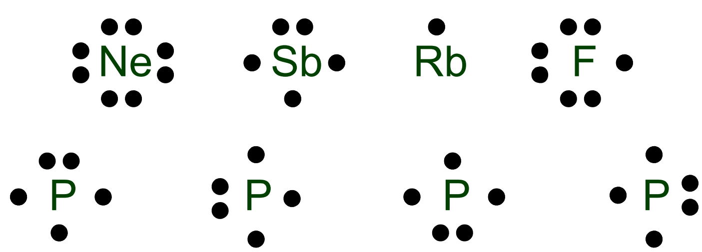
Write the chemical formula for the following covalent compounds: hexaboron silicide, chlorine dioxide, hydrogen iodide, phosphorus triiodide, dinitrogen trioxide.

Metallic Bonds

- Exist only in metals, such as aluminum, gold, copper, and iron
- Each atom is bonded to several other metal atoms, and their electrons are free to move throughout the metal structure
- Responsible for the unique properties of metals, such as their high conductivity

Lewis Dot Structures

- Used to show bonding of valence (outer) electrons
- Number of valence electrons is group on periodic table



Lewis Dot Structures

- Only C, N, O, P, and S can form double bonds

Lewis Dot Structures

Example:

Draw the Lewis structures for the following molecules: HF, N_2, NH_3, CH_4, CF_4, and NO^+.

Lewis Dot Structures

		valence electrons	single bonds	remaining electrons	final Lewis structure	number of electrons
a.	HF	1 + 7 = 8	H—F	6	H—F̤̈:	H: 2 F: 8
b.	N_2	5 + 5 = 10	N—N	8	:N≡N:	N: 8 N: 8
c.	NH_3	5 + 3(1) = 8	H—N—H | H	2	H—N̈—H | H	H: 2 N: 8
d.	CH_4	4 + 4(1) = 8	H | H—C—H | H	0	H | H—C—H | H	H: 2 C: 8
e.	CF_4	4 + 4(7) = 32	F | F—C—F | F	24	:F̤̈: | :F̤̈—C—F̤̈: | :F̤̈:	F: 8 C: 8
f.	NO^+	5 + 6 − 1 = 10	N—O	8	$[:N \equiv O:]^+$	N: 8 O: 8

Lewis Dot Structures

Example:

Which one of the following molecules contains a triple bond: PF_3, NF_3, C_2H_2, H_2CO, or HOF?

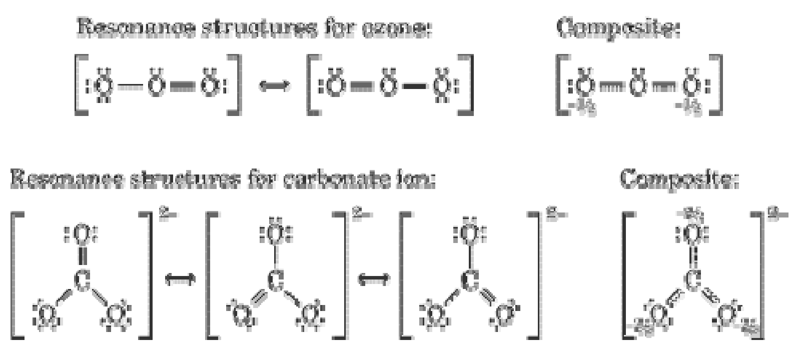
Lewis Dot Structures

Example:

How many sigma (σ) bonds and how many pi (π) bonds does the molecule ethene, C_2H_4, contain?

Resonance Structures

- Two or more Lewis structures that describe molecule
- Composite represents true structure for molecule



Molecular Gemeotry

- A molecule will assume the shape that most minimizes electron pair repulsions
- Two types of electron sets:
  - Electrons can exist in bonding pairs, which are involved in creating a single or multiple covalent bond,
  - Electrons can exist in nonbonding pairs, which are pairs of electrons that are not involved in a bond, but are localized to a single atom

Molecular Geometry

- Valence shell electron-pair repulsion (VSEPR) theory uses Lewis structures to predict the molecular geometry of bonded molecules
- <u>VSEPR Theory</u>: The 3D arrangement of atoms surrounding a central atom is determined by the repulsions between the bonding and nonbonding electron pairs in the valence shell of the central atom
- Electrons want to be as far apart as possible

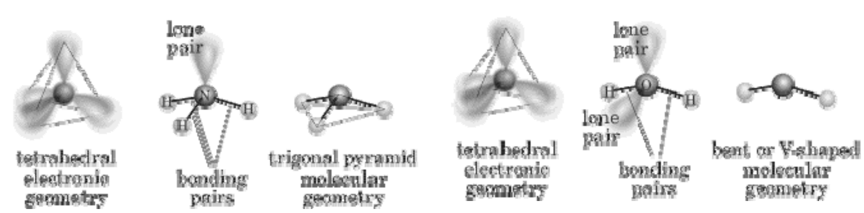
Molecular Geometry

- Use the structural pair geometry to determine the molecular geometry by following these steps:
  – Draw the Lewis dot structure for the molecule and count the total number of single bonds, multiple bonds, and unpaired electrons.
  – Determine the structural pair geometry for the molecule by arranging the electron pairs so that the repulsions are minimized (based on the table).
  – Use the table to determine the molecular geometry.

Molecular Geometry

Lone pairs have more repulsive force than do shared electron pairs, and thus they force the shared pairs to squeeze more closely together

- Ammonia (NH_3), which has three sigma bonds and a lone pair, is trigonal pyramidal
- Water (H_2O) has two lone pairs and its molecular geometry is "bent"



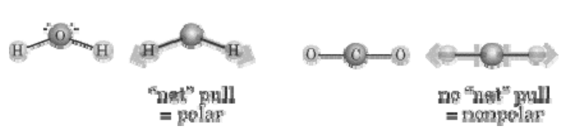
Molecular Geometry

Example:

Draw the dot formula for SeF_4 and determine the shape.

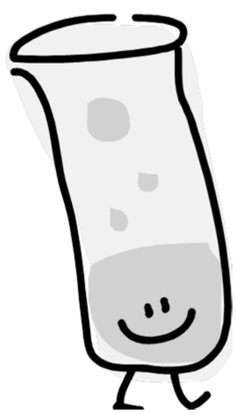
Molecular Polarity

- Polarity: an uneven distribution of electron pairs between the two bonded atoms
- One of the atoms is slightly more negative than the other
- Dipoles: polar molecules that have a slightly positive charge on one end and a slightly negative charge on the other
  - Water: two lone electron pairs on the oxygen atom establish a negative pole on this bent molecule, while the bound hydrogen atoms constitute a positive pole
- Molecules can also contain polar bonds and not be polar
  - Carbon dioxide: both of the C—O bonds in carbon dioxide are polar, but they're oriented such that they cancel each other out, and the molecule itself is not polar



Valence Bond Theory

- Orbital hybridization results from a blending of atomic orbitals (*s* and *p*) to create energy orbitals that have energy in between the energy of lone orbitals
- Changes the shape (angles within) the molecule
- Refer to table and book



Subject Test Chemistry

States of Matter

Strategy Notes

APPLYME

Intra- and Intermolecular Forces

- Along with the bonds that exist within each individual molecule, whether a particular group of bonded molecules takes the form of a solid, liquid, or gas depends on the presence and type of bonds between molecules
- Two types of intermolecular forces:
  – Dipole-dipole forces (including hydrogen bonds)
  – London dispersion forces

Dipole-Dipole Forces

- Take place when two or more neutral, polar molecules are oriented such that their positive (+) and negative (-) ends are close to each other
- Fairly strong type of intermolecular force
- Molecules held together by dipole-dipole forces tend to be in the solid or liquid state
- For molecules that are about the same size and weight, the strength of the dipole-dipole forces increases as the degree of polarity increases
- The more polar a molecule is, the stronger the dipole-dipole forces it will form with itself and other molecules

Hydrogen Bonding

- Unique case of the dipole-dipole attraction
- Not true bonds: just strong attractive forces between the hydrogen on one molecule and a highly electronegative atom on a nearby molecule
- Most commonly form between hydrogen atoms and fluorine, oxygen, or nitrogen.

London Dispersion Forces

- Relatively weak forces of attraction that exist between nonpolar molecules and noble gas atoms, like argon (a noble gas) and octane (a hydrocarbon; C_8H_{18})
- Caused by instantaneous dipole formation: electron distribution in the individual molecules suddenly becomes asymmetrical, and the newly formed dipoles are now attracted to one another.

London Dispersion Forces

- Polarizability: ease with which the electron cloud of an atom can be distorted to become
- The greater the number of electrons an electron has, the farther they will be from the nucleus, and the greater the chance for them to shift positions within the molecule
- Larger nonpolar molecules tend to have stronger London dispersion forces
- For nonpolar molecules, the farther you go down the group, the stronger the London dispersion forces
  - Fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid

Solids

- Molecules are generally held together by ionic or strong covalent bonding
- Attractive forces between the atoms, ions, or molecules in solids are very strong
- Particles in a solid are held in fixed positions and have very little freedom of movement.
- Solids have definite shapes and definite volumes and are not compressible to any extent
- Two main categories of solids:
  - Crystalline solids: the atoms, ions, or molecules that make up the solid exist in a regular, well-defined arrangement
  - Amorphous solids: do not have much order in their structures

Crystalline Solids

- Unit cell: smallest repeating pattern of crystalline solids; all identical and repeating
- Four types:
  - Ionic solids: Made up of positive and negative ions and held together by electrostatic attractions; very high melting points and brittleness and are poor conductors in the solid state (e.g., table salt, NaCl)
  - Molecular solids: Made up of atoms or molecules held together by London dispersion forces, dipole-dipole forces, or hydrogen bonds; low melting points and flexibility and are poor conductors (e.g., sucrose)
  - Covalent-network (also called atomic) solids: Made up of atoms connected by covalent bonds; the intermolecular forces are covalent bonds as well; very hard with very high melting points and poor conductors (e.g., diamond and graphite)
  - Metallic solids: Made up of metal atoms that are held together by metallic bonds; high melting points, can range from soft and malleable to very hard, and are good conductors of electricity

Amorphous Solids

- Do not have much order in their structures.
- Though their molecules are close together and have little freedom to move, they are not arranged in a regular order
- Common examples are glass and plastics.

Liquids

- Made up of molecules that contain covalent bonds and have strong intermolecular attractive forces.
- Atoms and molecules have more freedom of movement than do those in solids.
- Liquids have no definite shape but do have a definite volume, and they are not easily compressible

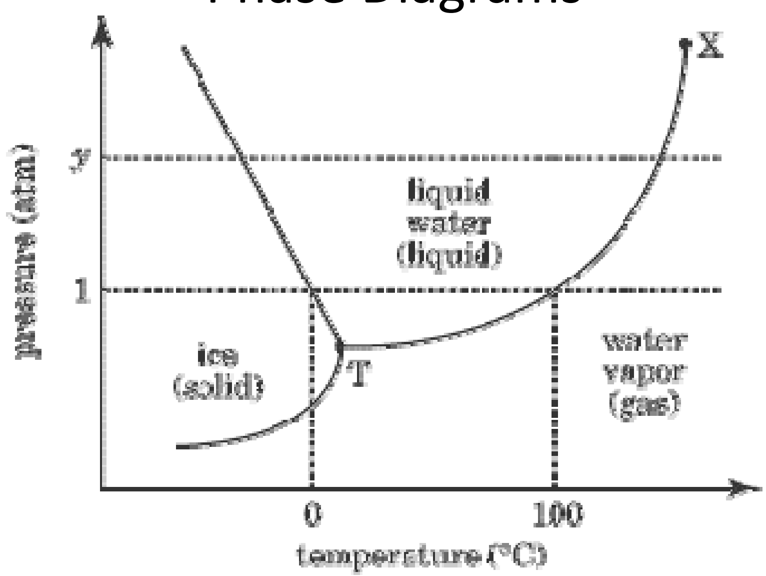
Gases

- Generally consist of atoms and molecules that are covalently bonded, and their intermolecular forces are very weak
- Molecules of a gas are highly separated, so gases are mostly empty space
- A gas has no definite shape—it will take the shape of the container that holds it, and gases are easily compressible.

Phase Changes

- Fusion: to turn from a solid into a liquid
- Vaporization: to turn from a gas to a liquid
- In order for a substance to move between the states of matter, energy must be gained or lost
- Heat of fusion (H_{fus}): the amount of energy that must be put into the substance for it to melt (kJ/mol or cal/g)
- Heat of vaporization (H_{vap}): the amount of energy needed to cause the transition from liquid to gas
- Phase diagrams show changes in states of matter

Phase Diagrams



Phase Diagrams

- Triple point: intersection of all three lines is known as the triple point
  - All three phases of matter are in equilibrium with each other
- Critical point: at this point and beyond, the substance is forever in the vapor phase
- Why does water boil at a lower temperature at higher altitudes?

Phase Diagrams

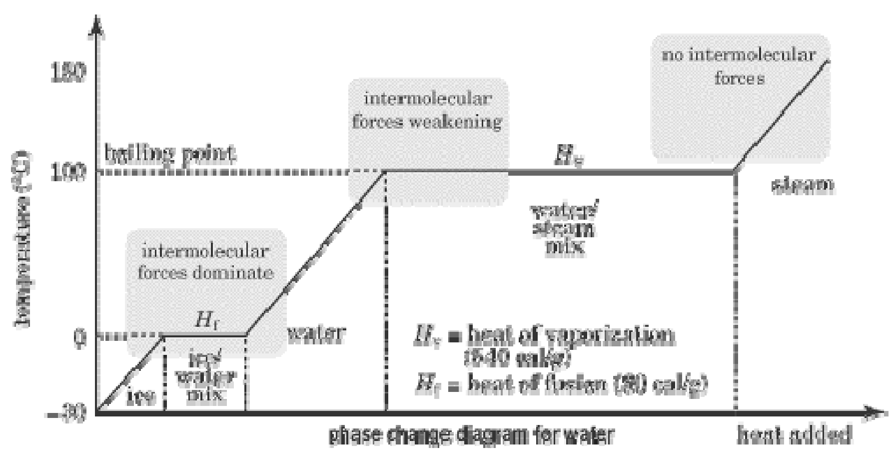
- If we put a liquid into a closed container, the evaporation of the liquid will cause an initial increase in the total pressure of the system, and then the pressure of the system will become a constant
- The value of this final pressure is unique to each liquid and is known as the liquid's vapor pressure
- Water and other liquids that have low vapor pressures are said to be nonvolatile
- Substances like rubbing alcohol and gasoline, which have relatively high vapor pressures, are said to be volatile

Phase Diagrams

Example:

What happens to water when the pressure remains constant at 1 atm but the temperature changes from -10ºC to 75ºC?

Heating Curves



Specific Heat

- The heat required to raise the temperature of 1 g of a substance by 1ºC
- Will be given on exam for a particular substance
- Related equation:

$$\textbf{\textit{energy (in calories)}} = mC_p\,DT$$

where m = the mass of the substance (in grams)

C_p = the specific heat of the substance (in cal/g ºC)

DT = the change in temperature of the substance (in either Kelvins or ºC, but make sure all your units are compatible!)

Specific Heat

Example:

If you had a 10.0 g piece of ice at -10ºC, under constant pressure of 1 atm, how much energy would be needed to melt this ice and raise the temperature to 25.0ºC? (The specific heat for ice is 0.485 cal/g ºC.)

The Gas Laws: Ideal Gases

- In an ideal state:
  - All gas particles are in constant, random motion
  - All collisions between gas particles are perfectly elastic (meaning that the kinetic energy of the system is conserved)
  - The volume of the gas molecules in a gas is negligible
  - Gases have no intermolecular attractive or repulsive forces
  - The average kinetic energy of the gas is directly proportional to its Kelvin temperature and is the same for all gases at a specified temperature

The Gas Laws: Ideal Gases

- Four measurable properties used for gases:

1. Quantity (amount) of the gas in moles (n)
2. Temperature, T, of gases must always be converted to the Kelvin temperature scale (the absolute temperature scale)
3. Volume, V, of a gas in liters (L)
4. Pressure, P, of a gas in atmospheres (atm)
   - Standard temperature and pressure (STP) = 273K (or 0ºC) and 1 atm

The Gas Laws: Ideal Gases

Example:

Which of the following statements is not true of ideal gases?

1. The volume occupied by gas particles is only significant at very low pressures.
2. Gas molecules occupy an insignificant volume compared to the volume of the container that holds them.
3. The particles of a gas move in random straight line paths until a collision occurs.
4. The collisions that occur between gas particles are considered elastic.
5. At a given temperature, all gas molecules within a sample possess the same average kinetic energy.

Measuring the Pressure of a Gas

- Memorize these for the exam so you can convert units where necessary:

760 mmHg

760 torr

1.00 atm

101,325 Pa

101.325 kPa

Boyle’s Law

- The volume of a confined gas at a fixed temperature is inversely proportional to the pressure exerted on the gas
- *PV* = a constant

$$\boxed{P_1V_1 = P_2V_2}$$

Boyle’s Law

Example:

Sulfur dioxide (SO_2) gas is a component of car exhaust and power plant discharge, and it plays a major role in the formation of acid rain. Consider a 3.0 L sample of gaseous SO_2 at a pressure of 1.0 atm. If the pressure is changed to 1.5 atm at a constant temperature, what will be the new volume of the gas?

Charles's Law

- If a given quantity of gas is held at a constant pressure, its volume is directly proportional to the absolute temperature
- As the temperature of the gas increases, the gas molecules will begin to move around more quickly and hit the walls of their container with more force—thus the volume will increase
- Temperature must be in Kelvins!

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

Charles's Law

Example:

A sample of gas at 15ºC and 1 atm has a volume of 2.50 L. What volume will this gas occupy at 30ºC and 1 atm?

Avogadro's Law

- The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles (n) of the gas
- Equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules

$$\boxed{V = constant \times n}$$

The Ideal Gas Law

$$\boxed{PV = nRT}$$

where P = pressure (atm), V = volume (L), n = number of moles (mol), R = 0.08206 L · atm/mol · K, and T = temperature (K)

The Ideal Gas Law

Example:

A 16.0 g sample of methane gas, CH_4, the gas used in chemistry lab, has a volume of 5.0 L at 27ºC. Calculate the pressure.

Density of Gases

- Since gases are mostly empty space, the densities of gases are reported in g/L

$$\boxed{Density = \frac{FW}{22.4}}$$

where the formula weight (*FW*) is in g/mol, and the standard molar volume is 22.4 L/mol

Density of Gases

Example:

What is the density of helium gas at STP?

Dalton's Law of Partial Pressures

- The pressure of a mixture of gases is the sum of the pressures that each of the individual gases would exert if it were alone

$$\boxed{P_T = P_1 + P_2 \ldots P_n}$$

Graham's Law of Diffusion and Effusion

- The rates of effusion of two gases are inversely proportional to the square roots of their molar masses at the same temperature and pressure
- Effusion: the passage of a gas through a tiny orifice into an evacuated chamber
- Rate of effusion measures the speed at which the gas travels through the tiny hole into a vacuum
  - (Diffusion: the spread of a gas throughout a space or throughout a second substance)

$$\boxed{\frac{Rate\ of\ effusion\ of\ gas\ 1}{Rate\ of\ effusion\ of\ gas\ 2} = \sqrt{\frac{FW_2}{FW_1}}}$$

Solutions

- Solution: a homogenous mixture of two or more substances that exist in a single phase
- Two main parts of solutions
  - Solute: component of a solution that is dissolved in the solvent; usually present in a smaller amount than the solvent
  - Solvent: component into which the solute is dissolved, and it is usually present in greater concentration
- Aqueous solution: solutions where water is the solvent

Solutions

- **Like dissolves like**
- Polar, ionic substances are soluble in polar solvents
- Nonpolar solutes are soluble in nonpolar solvents
- Examples
  - Alcohol and water, which are both polar, can form a solution
  - Iodine and carbon tetrachloride, which are both nonpolar, make a solution
  - Iodine will <u>not</u> readily dissolve in polar water

Solutions

- Particles are really small—anywhere from 0 to 100 nm
- Never settle on standing
- Cannot be separated by filtering
- Light will pass through a solution unchanged

Solutions

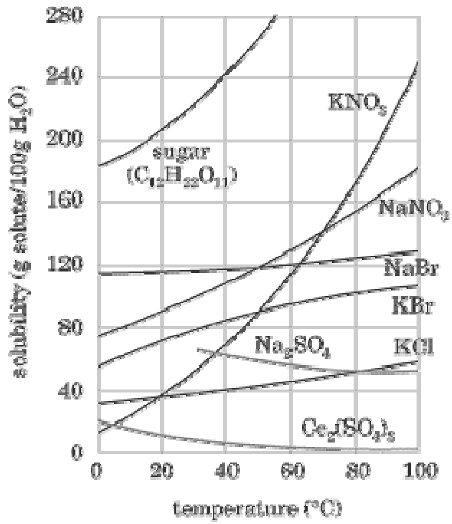
- Not considered solutions
  - Colloid: small particles will not settle and cannot be filtered, but do scatter light
    - Examples: gelatin, fog, smoke, and shaving cream
  - Suspension: larger particles that will settle on standing, can often be separated by a filter, and may scatter light
    - Examples: muddy water, paint, and some medicines, like Pepto-Bismol

The Solution Process

- In order for a solute to be dissolved in a solvent, the attractive forces between the solute and solvent particles must be great enough to overcome the attractive forces within the pure solvent and pure solute
- Solute and solvent molecules in a solution are expanded compared to their position within the pure substances.
- Expansion involves a change in the energy of the system
  - Endothermic: the separation of the solute particles from one another prior to dissolving
  - Exothermic: the solute and solvent combine with each other
- If the energy released is greater than the energy absorbed, the solution forms and is stable
- After dissolving, the solute is said to be fully <u>solvated</u> (usually by dipole-dipole or ion-dipole forces), and when the solvent is water, the solute is said to be <u>hydrated</u>

Solutions

- Solubility: the maximum amount of material that will dissolve in a given amount of solvent at a given temperature to produce a stable solution
- Most solids increase in solubility with an increase in temperature
- Gases decrease in solubility with an increase in temperature



Degrees of Saturation

- Three degrees of saturation
  - Unsaturated: the solvent is capable of dissolving more solute
  - Saturated: solvent has dissolved the maximum amount of solute that it can at the given temperature
    - Dynamic equilibrium: the processes of dissolving and precipitation are happening at the same rate.
  - Supersaturated: the solvent contains more solute than it can theoretically hold at a given temperature
    - Formed by heating a solution and dissolving more solute, then cooling the solution down slowly
    - Unstable and crystallize readily

Molarity

- Molarity (M, moles/L): is a measure of the number of moles of solute per liter of solution
- Most common concentration unit used in chemistry
- [NaCl] = 0.75 which means that 0.75 mole of NaCl is dissolved per 1.00 L of solution

$$M = \frac{moles\ of\ solute}{liters\ of\ solution}$$

Molarity

Example:

Calculate the molarity of a solution prepared by dissolving 20.0 g of solid NaOH in enough water to make 100 mL of solution.

Dilution

- Dilution: the process of taking a more concentrated solution and adding water to make it less concentrated
- The more concentrated solution before the dilution is performed is known as the *stock* solution

$$\boxed{M_1V_1 = M_2V_2}$$

Dilution

Example:

What volume of 6.0 M sulfuric acid (H_2SO_4) must be used to prepare 2.0 L of a 0.10*M* H_2SO_4 solution?

Mass Percent (Weight Percent)

- Mass percent of a solution is another way of expressing its concentration

- $\boxed{Mass\ percent = \frac{grams\ of\ solute}{g\ of\ solute + g\ of\ solvent} \times 100\%}$

Mass Percent (Weight Percent)

Example:

A solution is prepared by mixing 5.00 g ethanol (C_2H_5OH) with 100.0 g water. Calculate the mass percent of ethanol in this solution.

Molality

- Molality (m): a measure of the number of moles of solute per kilogram of solvent
- Whereas the *molarity* of a solution is dependent on the volume of the solution, the *molality* is dependent on the mass of the solvent in the solution

$$\boxed{m = \frac{moles\ of\ solute}{kilograms\ of\ solvent}}$$

Molality

Example:

A solution is prepared by mixing 80.0 g of sodium hydroxide (NaOH) with 500.0 g of water. Calculate the molality of this solution.

Electrolytes

- Electrolytes: solutions that conduct an electric current
- Three classes of electrolytes (solutions that conduct a current): acids, bases, and salts
- The greater the degree of dissociation of the solute, the greater the conductivity of the solution
- Strong electrolytes: solutes that dissociate completely in solution; strong acids, strong bases, and soluble salts
- Nonelectrolytes: predominantly covalently bonded, generally will not produce ions in solution, and therefore are considered nonconductors
- Weak electrolytes: solutes that dissociate only a little in solution; Weak acids, weak bases, and slightly soluble salts

Colligative Properties

- Colligative properties: depend on the number of solute particles present per solvent molecule
- The concentration of solute in a solution can affect various physical properties of the solvent including its freezing point, boiling point, and vapor pressure

Freezing Point Depression

- Freezing point: the temperature at which the vapor pressure of the solid and the liquid states of that substance are equal
- If the vapor pressure of the liquid is lowered, the freezing point decreases
- A solution's freezing point is depressed below that of a pure solvent → molecules cluster in order to freeze; if they act as a solvent, solute molecules get in the way and prevent them from clustering tightly together
- The more ions in solution, the greater the effect on the freezing point

Freezing Point Depression

$$\boxed{DT_f = K_f \times m_{solute} \times i}$$

DT_f = the change in freezing point

K_f = molal freezing point depression constant for the substance (for water = 1.86ºC/*m*)

m = molality of the solution

i = number of ions in solution (this is equal to 1 for covalent compounds and is equal to the number of ions in solution for ionic compounds)

Boiling Point Elevation

- Boiling point: the temperature at which the vapor pressure equals atmospheric pressure
- The boiling point is increased because vapor pressure is lowered by the addition of a nonvolatile solute → Since the solute particles get in the way of the solvent particles trying to escape the substance as they move around faster, it will take more energy for the vapor pressure to reach atmospheric pressure, and thus the boiling point increases.

Boiling Point Elevation

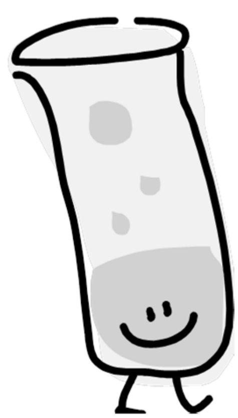
$$\boxed{DT_b = K_b \times m_{solute} \times i}$$

K_b = molal boiling point elevation constant (for water = 0.51˚C/*m*)

Boiling Point Elevation

Example:

Calculate the freezing point and boiling point of a solution of 100 g of ethylene glycol ($C_2H_6O_2$) in 900 g of water.



Subject Test Chemistry

Reaction Types

Strategy Notes

APPLYME

The Basics

Reactant → Product

Symbol	Meaning
→	"Yields" or "produces"
+	"Reacts with" or "and"
(g)	Gaseous state
(l)	Liquid state
(s)	Solid state
(aq)	Aqueous state (dissolved in water)
number subscript	Represents the number of atoms of the element it's to the right of
number coefficient	How many molecules or moles of the substance are reacting
$\xrightarrow{Pt}$	A substance named above the arrow represents a catalyst in the reaction
↑	A gas is produced
↓	A precipitate is formed
kJ or J	Energy term (kilojoules or joules)
⇔	Reversible equation; equilibrium
$\xrightarrow{\Delta}$	A delta above the reaction arrow indicates that heat is added to the reaction

Balancing Chemical Equations

Example:

Write the balanced equation for the reaction between chlorine and sodium bromide, which produces bromine and sodium chloride.

Balancing Chemical Equations

Example:

Write the balanced equation for the reaction between aluminum sulfate and calcium chloride, which produces aluminum chloride and calcium sulfate.

Types of Reactions

- Synthesis reaction: Two or more elements or compounds combine to form a single product (A + B $\rightarrow$ C)

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$
$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_3(aq)$$
$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

Types of Reactions

- Decomposition reaction: A single reactant, a compound, breaks into two or more parts (C → A + B)

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$
$$H_3CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$
$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$
$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Types of Reactions

- Single replacement or displacement reaction: a more active element replaces a less active element in a compound (A + BC → AC + B or A + BC → AB + C)

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$
$$Cl_2(g) + 2NaI(aq) \rightarrow 2NaCl(aq) + I_2(s)$$

Types of Reactions

- Double replacement or displacement reaction: two compounds react to form two new compounds (AB + CD → AC + BD)

$$Pb(NO_3)2(aq) + 2KI(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$$
$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

Types of Reactions

- Combustion reaction: A hydrocarbon is burned in the presence of oxygen gas to form carbon dioxide (in a complete combustion) or carbon monoxide (in an incomplete combustion, due to a limited amount of oxygen) ($C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(l)$)

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Types of Reactions

- **Hydrolysis reaction**: A reaction that involves water.

$$X^-(aq) + H_2O(l) \longleftrightarrow HX(aq) + OH^-(aq)$$

Net Ionic Equations

- Net ionic equations: equations that show only the soluble, strong electrolytes reacting (these are represented as ions) and omit the spectator ions, which go through the reaction unchanged
- Based on solubility rules
- If the ion is soluble, it won't form a precipitate, and this means it doesn't react and should be left out of the net ionic equation

Solubility Rules

1. Most alkali metal (Group 1A) compounds *and* compounds are soluble.
2. Cl^-, Br^-, I^- compounds are soluble, except when they contain Ag^+, Hg_2^{2+}, or Pb^{2+}.
3. F^- compounds are soluble, except when they contain group 2A metals.
4. NO_3^-, ClO_3^-, ClO_4^-, and CH_3COO^- compounds are soluble.
5. compounds are soluble, except when they include Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, Pb^{2+}, or Hg_2^{2+}.
6. CO_3^{2-}, PO_4^{3-}, $C_2O_4^{2-}$, CrO_4^{2-}, S^{2-}, OH^-, and O^{2-} compounds are insoluble.
7. Group 2A metal oxides are classified as strong bases even though they are not very soluble.

Other Rules

- If an insoluble precipitate or gas can be formed in a reaction, it probably will be.
- Oxides (except group 1A) are insoluble, and when reacted with water, they form either acids (nonmetal oxides) or bases (metal oxides).
- There are six strong acids that completely ionize: HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$. All other acids are weak and are written together, as molecules.
- The strong bases that ionize are oxides and hydroxides of group 1A and 2A metals. All other oxides and hydroxides are considered weak and written together, as molecules.

Net Ionic Equations

Example:

Write the net ionic equation for a mixture of solutions of silver nitrate and lithium bromide.

Net Ionic Equations

Example:

Hydrochloric acid and sodium hydroxide are mixed. Write the net ionic equation.

Net Ionic Equations

Example:

Chlorine gas is bubbled into a solution of potassium iodide; write the net ionic equation.

The Chemistry of Acids and Bases

- Arrhenius acids and bases are:
  - Acid: a substance that increases the concentration of protons (H^+) in water
  - Base: a substance that increases the concentration of hydroxide ions in water (OH^-)
- These definitions are limited to aqueous solutions

The Chemistry of Acids and Bases

- Brønsted and Lowry acids and bases as:
  - Acid: a substance that donates a proton to another substance
  - Base: a substance that accepts a proton
- These definitions can also apply to reactions that are not aqueous, so they are more accurate

The Chemistry of Acids and Bases

- Lewis acids and bases are:
  - Acid: a substance that accepts an electron pair
  - Base: a substance that donates an electron pair

The Chemistry of Acids and Bases

- Hydronium (H_3O^+): H^+ riding "piggyback" on a water molecule; water is polar, and the positive charge of the naked proton is greatly attracted to one of the negative electron pairs on adjacent oxygen
- Monoprotic: describes acids that can donate one H^+
- Diprotic: describes acids that can donate two H^+ ions
- Polyprotic: describes acids that can donate more than one H^+ ion
- Amphiprotic: describes a substance that can act as either an acid or a base. This means it can either lose a proton or gain one. Water is amphiprotic: it can form either a hydroxide ion or a hydronium ion. Other examples of amphiprotic substances are HCO_3^-, HSO_4^-, HPO_4^{2-}

Conjugate Acid-Base Pairs

- Brønsted and Lowry acids and bases as:
  - Acid: a substance that donates a proton to another substance
  - Base: a substance that accepts a proton
- Compounds that differ by the presence of one proton (H+)
- All acids have a conjugate base → formed when their proton has been donated
- All bases have a conjugate acid → formed after they have accepted a proton

Conjugate Acid-Base Pairs

- Generic acid-base reaction:

  $HX(aq) + H_2O(l) \leftrightarrow X^-(aq) + H_3O(aq)$
- Forward reaction: HX (as base) donates a proton to water to form hydronium
- Reverse reaction: hydronium (as acid) donates a proton to X-
- HX and X- are conjugate acid-base pairs

Conjugate Acid-Base Pairs

Example:

Apply the appropriate acid-base theory to first identify the acid and base reacting and then identify the conjugate acid-base pairs in the examples below:

1. $HNO_3 + H_20 \rightarrow H_3O^+ + NO_3^-$
2. $NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$

Relative Strengths of Acids and Bases

- Strong acids: dissociates or ionizes completely in aqueous solution; better at donating a proton
- More O atoms in polyatomic ion = stronger acid
- Hydrohalic acids: HCl, HBr, HI
- Nitric acid: HNO_3
- Sulfuric acid: H_2SO_4
- Perchloric acid: $HClO_4$

Relative Strengths of Acids and Bases

- Strong bases: dissociates or ionizes completely in aqueous solution; better at accepting a proton
- Hydroxides (—OH)
- Oxides of 1A and 2A metals (except Mg and Be)
- H^-
- CH_3^-

Relative Strengths of Acids and Bases

- Remember that the stronger the acid, the weaker its conjugate base
- The converse is also true
- See chart

The pH Scale

- Water can act as either a proton donor (in the form of the hydronium ion, H_3O^+) or a proton acceptor (as OH^-)
- Autoionization: when a water molecule donates a proton to or accepts a proton from another water molecule in solution

$$2H_2O \rightarrow H_3O^+ + OH^-$$

The pH Scale

- Autoionization takes place in equilibrium so the equilibrium expression, K_{eq}

$$\boxed{K_{eq} = [H_3O^+][OH^-]}$$

- K_{eq} for water (K_w) is 1×10^{-14}
- $[H_3O^+] = [OH^-] = 1 \times 10^{-7}$ (neutral)

The pH Scale

- In acidic solutions, the concentration of H^+ is higher than that of OH^-, and in basic solutions, the concentration of OH^- is greater than that of H^+
- pH of a solution:

$$\boxed{pH = -\log [H^+] \text{ or } -\log [H_3O^+]}$$

The pH Scale

- pOH of a solution:

$$\boxed{pOH = -\log [OH^-]}$$

- pH and pOH are related to each other:

$$\boxed{pH + pOH = 14}$$

The pH Scale

Example:

What is the pH of a solution at 25ºC in which $[OH^-] = 1.010^{-5}M$?

Acid–Base Reactions: Neutralization Reactions

- When a strong acid and a strong base solution are mixed, a neutralization reaction occurs, and the products do not have characteristics of either acids or base
- If a strong acid and a weak base are mixed, the resulting salt will be acidic
- If a strong base and a weak acid are mixed, the resulting salt will be basic

Acid–Base Reactions: Neutralization Reactions

Example:

Classify each of the salts listed below as acidic, basic, or neutral.

1. $Fe(NO_3)_3$
2. $MgSO_4$
3. $Ni(ClO_4)_2$

Oxidation-Reduction

- Oxidation-reduction reactions involve the transfer of electrons between substances
- Take place simultaneously, which makes sense because if one substance loses electrons, another must gain them
- All single-replacement reactions are redox reactions

Oxidation-Reduction

- Electrochemistry: the study of the interchange of chemical and electrical energy
- Oxidation: the loss of electrons; since electrons are negative, this will appear as an increase in the charge (e.g., Zn loses two electrons; its charge goes from 0 to +2); metals are oxidized
- Oxidizing agent (OA): the species that is reduced and thus causes oxidation
- Oxidation number: the assigned charge on an atom; used to balance formulas

Oxidation-Reduction

- Reduction: the gain of electrons; When an element gains electrons, the charge on the element appears to decrease, so we say it has a reduction of charge (e.g., Cl gains one electron and goes from an oxidation number of 0 to -1); nonmetals are reduced
- Reducing agent (RA): the species that is oxidized and thus causes reduction
- Half-reaction: an equation that shows either oxidation or reduction alone

Rules for Assigning Oxidation States

- A reaction is considered a redox reaction if the oxidation numbers of the elements in the reaction change in the course of the reaction
- Use the following rules to assign oxidation states to the components of oxidation-reduction reactions

1. The oxidation state of an element is *zero*, including all elemental forms of the elements (e.g., N_2, P_4, S_8, O_3).
2. The oxidation state of a monatomic ion is the same as its charge.
3. In compounds, fluorine is always assigned an oxidation state of -1.

Rules for Assigning Oxidation States

4. Oxygen is usually assigned an oxidation state of -2 in its covalent compounds. Exceptions to this rule include peroxides (compounds containing the group), where each oxygen is assigned an oxidation state of -1, as in hydrogen peroxide (H_2O_2).
5. Hydrogen is assigned an oxidation state of +1. Metal hydrides are an exception: in metal hydrides, H has an oxidation state of -1.
6. The sum of the oxidation states must be zero for an electrically neutral compound.
7. For a polyatomic ion, the sum of the oxidation states must equal the charge of the ion.

Rules for Assigning Oxidation States

Example:

Assign oxidation numbers to each element in the following:

1. H_2S
2. MgF_2
3. PO_4^{3-}

Rules for Assigning Oxidation States

Example:

When powdered zinc metal is mixed with iodine crystals and a drop of water is added, the resulting reaction produces a great deal of energy. The mixture bursts into flames, and a purple smoke made up of I_2 vapor is produced from the excess iodine. The equation for the reaction is

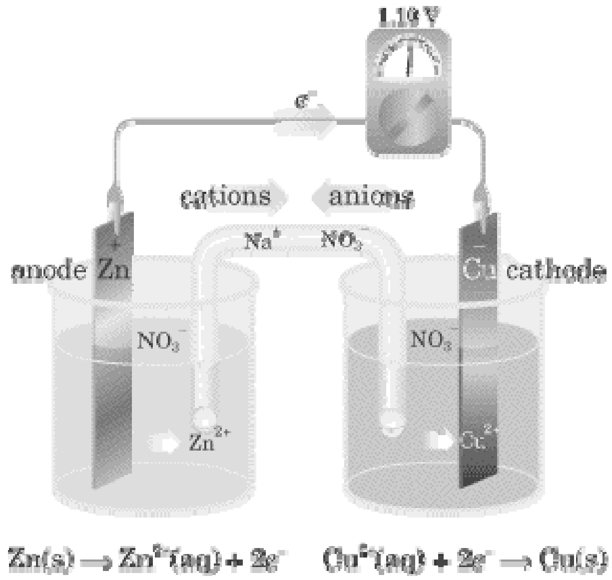
$$Zn_{(s)} + I_{2(s)} \rightarrow ZnI_{2(s)} + \text{energy}$$

Identify the elements that are oxidized and reduced, and determine the oxidizing and reducing agents.

Voltaic (or Galvanic) Cells

- Redox reactions release energy, and this energy can be used to do work if the reactions take place in a voltaic cell (galvanic cell)
- In a voltaic cell, the transfer of electrons occurs through an external pathway instead of directly between the two elements
- Electron flow always occurs from anode to cathode, through the wire that connects the two half-cells, and a voltmeter is used to measure the cell potential in volts
  - Zinc with copper
  - Lead with lead (IV) oxide

Voltaic (or Galvanic) Cells



Standard Reduction Potentials

- Potential of a voltaic cell as a whole will depend on the half-cells that are involved
- Standard reduction potential (Eº): Known potential of each half-cell has a known potential
- Cell potential is a measure of the difference between the two electrode potentials, and the potential at each electrode is calculated as the potential for *reduction* at the electrode
- Elements with *most positive* reduction potentials are easily reduced and would be good oxidizing agents (in general, the nonmetals)
- Elements with least positive reduction potentials are easily oxidized and would be good reducing agents (in general, metals)

Standard Reduction Potentials

Example:

Which of the following elements would be most easily oxidized: Ca, Cu, Fe, Li, or Au?

Standard Reduction Potentials

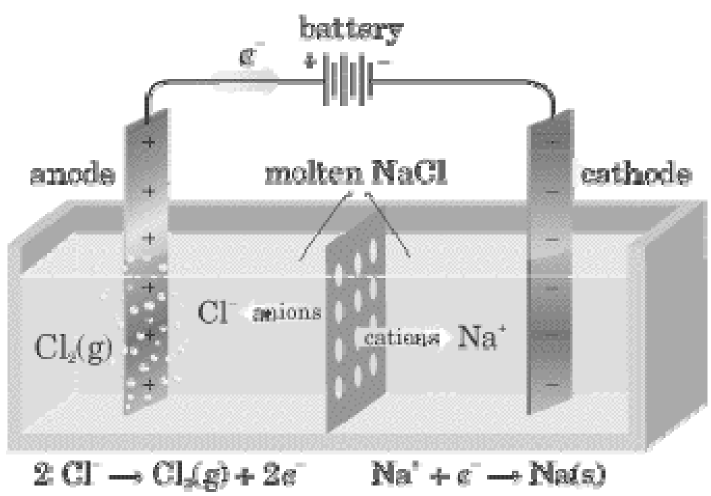
Example:

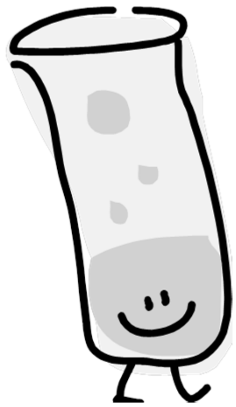
Which one of the following would be the best oxidizing agent: Ba, Na, Cl, F, or Br?

Electrolytic Cells

- Electrolytic cells can be used to drive nonspontaneous redox reactions (electrolysis reactions)
- Electrons still flow from the anode to the cathode
- Electrolytic cells are used to produce pure forms of an element
- Electrolytic cell *needs* a battery
- Electrolytic cells have just one container

Electrolytic Cells





Subject Test Chemistry

Stoichiometry

Strategy Notes

APPLYME

The Mole

- Matter is neither created nor destroyed in the course of a chemical reaction
- Atomic weight of a substance: the mass of one mole of a substance
- One mole of a substance is equal to 6.02 x 10^{23} atoms (Avogadro's number) or molecules of the substance
- Ex: carbon's atomic weight is roughly 12 amu; this means that 6.02 x 10^{23}carbon atoms, in a pile, weigh 12 grams

The Mole

Example:

Substance	Molar mass	Number of moles	Mass in grams	Number of particles
Carbon dioxide, CO_2		3.0		
Oxygen, O_2			64.0	
Methane, CH_4		0.279		
Nitrogen, N_2				9.50×10^{25}

Percent Composition of Compounds

Example:

Calculate the percent composition of each of the elements in C_2H_5OH.

Dimensional Analysis

Example:

What mass of oxygen will react completely with 96.1 grams of propane?

Dimensional Analysis

Example:

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the air; it reacts with carbon dioxide to form solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide will be consumed in a reaction with 1.00 kg of lithium hydroxide?

Dimensional Analysis

Example:

Baking soda ($NaHCO_3$) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach in the reaction below:

$$NaHCO_{3(s)} + HCl_{(aq)} NaCl_{(aq)} + H_2O_{(l)} + CO_{2(aq)}$$

How many grams of $NaHCO_3$ would be needed to completely react with 10.0 g of HCl?

Limiting Reagents

- Limiting reactant or reagent: consumed first in the chemical reaction, and its consumption halts the progress of the forward reaction

Limiting Reagent

Example:

In the Haber process, ammonia is created from the reaction of nitrogen and hydrogen gas. Suppose you have a total of 25.0 kg of nitrogen to react with a total of 5.00 kg of hydrogen. What mass of ammonia can be produced? Which reactant is the limiting reactant? What is the mass of the reactant that's in excess?

Chemical Yields

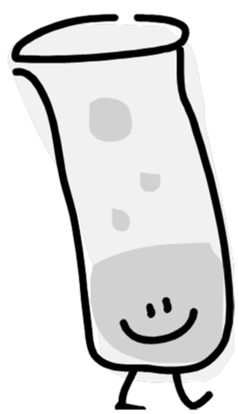
- Theoretical yield: the amount of product formed once the limiting reactant has been completely consumed (assumes perfect conditions and gives a maximum amount)
- Actual yield: what actually occurs in the course of the reaction—how much product is actually formed
- Percent yield is the ratio of the actual yield to the theoretical yield

$$\frac{actual\ yield}{theoretical\ yield} \times 100\% = \text{percent yield}$$

Chemical Yields

Example:

What would the percent yield be if you performed the reaction from the previous example and you actually collected 26,450 grams of ammonia?



Subject Test Chemistry

Equilibrium and Rates

Strategy Notes

APPLYME

Factors That Affect Reaction Rates

- Reaction rate: a measure of the change in the concentration of reactants or products over time in a chemical reaction
- 4 main external conditions affect reaction rate
  - Concentration of the reactants
  - Temperature
  - Presence of a catalyst
  - Physical state of the reactants

Factors That Affect Reaction Rates

- Concentration of reagents
- If we increase the concentration of one or more reactants, the reaction will go more quickly
- The more molecules, the more collisions between molecules, and the faster the reaction will go

Factors That Affect Reaction Rates

- Temperature
- The higher the temperature of the reaction, the more quickly it will proceed
- At higher temperatures, the molecules are moving around more quickly (they have more kinetic energy)
- More collisions with each other with more energy makes it more likely that they will overcome the activation energy needed to start the reaction
- A 10˚C increase in temperature will double the reaction rate

Factors That Affect Reaction Rates

- Presence of a catalyst
- A catalyst speeds up the rate of reaction by lowering the activation energy
- Biological catalysts are known as enzymes
- Catalysts is that they are not consumed in the course of the reaction

Factors That Affect Reaction Rates

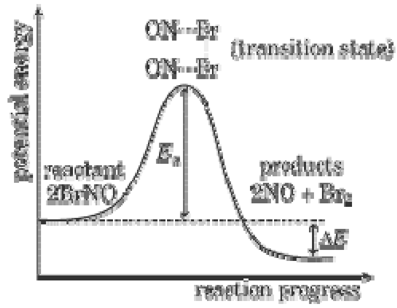
- Physical state of the reactants
- If one is a gas and one is a liquid, then the reaction area is limited to the area where they touch each other
- The larger this area, the faster the reaction will proceed (faster reactions if reactants are in same state)
- For example, consider a teaspoon of salt dissolving in water → it would take much longer for the salt to dissolve if you don't stir than if you stirred the solution

Energy Diagrams

- In order for a reaction to occur, reactant molecules must collide and that both an increase in the concentration of reactant molecules and an increase in the temperature of the system can cause an increase in reaction rate
- It takes more than just a regular collision to cause a chemical reaction to occur
  - The colliding molecules must be oriented in exactly the correct way: they must be oriented in suitable way for the product molecule bonds to be formed
  - The two molecules must collide with sufficient energy to overcome the activation energy of the reaction.
- Activation energy (E_a): the minimum energy needed to initiate a chemical reaction, and it is symbolized by

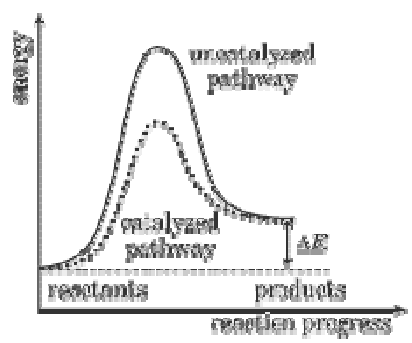
Energy Diagrams

- Energy diagram: a graph of the progress of a chemical reaction, versus the total energy of the system
- Exothermic reaction: energy is given off
- Endothermic reaction: energy required; energy of the products is higher than that of the reactants



Energy Diagrams

- Energy diagram for a reaction in the presence of a catalyst and in the absence of a catalyst
- The catalyst has decreased the activation energy of the reaction, which means that more molecules are able to surmount it and react



Equilibrium

- Chemical equilibrium has been reached in a reaction when the rate of the forward reaction is equal to the rate of the reverse reaction
- Collisions are still occurring: the reaction is now happening in each direction at the same rate.
- Reactants are being formed at the same rate as products are being formed

Equilibrium

- Concentration of the reactants and products in a reaction at equilibrium expressed by equilibrium constant, K_{eq}:
- Reaction:

$$aA + bB \leftrightarrow cC + dD$$

$$Equilibrium\ constant = K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Equilibrium

- K_c = equilibrium constant in an aqueous solution
- K_p = partial pressures of gases in equilibrium
- K_{sp} = solubility product of solids classified as insoluble
- *K* values have no units
- $K > 1$ means that the reaction favors the products at equilibrium
- $K < 1$ means that the reaction favors the reactants at equilibrium

Equilibrium Constant Rules

1. Pure solids and do not appear in the equilibrium expression.
2. Water, either as a liquid or solid, does not appear in the equilibrium expression.
3. When a reactant or product is preceded by a coefficient, its concentration is raised to the power of that coefficient in the K_{eq} expression.
4. When the K_{eq} of a reaction has been multiplied by a number, the *K* is raised to the power of the multiplication factor (K^n), so if it has been multiplied by 2, *K* is squared, if it has been multiplied by 3, *K* is cubed, and so on.
5. The K_{eq} of a reaction occurring in the reverse direction is simply the inverse of theK_{eq} of the reaction occurring in the forward direction ($1/K_{eq}$).
6. The K_{eq} of a net reaction that has two or more steps is found by the product of theK_{eq} s for each of the steps: $K_s = (K_1K_2K_3 \ldots)$.

Equilibrium

Example:

Write the equilibrium expression for the following equation:

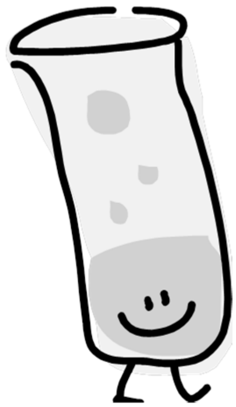
$$H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$$

If *K* is calculated to have a value of 2.5 for the reaction above, what is the value of the equilibrium constant for the following reaction?

$$4HI_{(g)} \leftrightarrow 2H_{2(g)} + 2I_{2(g)}$$

Le Chatelier's Principle

- If stress is applied to a system at equilibrium, the position of the equilibrium will shift in the direction that reduces the stress to reinstate equilibrium
  - If more reactants are added to the system, the reaction will shift in the forward direction
  - If more products are added, the reaction will shift in the reverse direction
  - If heat is added to the system and the reaction is exothermic, heat should be thought of as a product and the reaction will shift to the left
  - If the reaction is endothermic and heat is added, the reaction will shift to the right
  - The addition of pressure will cause a shift in the direction that results in the fewer number of moles of a gas
  - If pressure is relieved, the reaction will shift in the direction that produces more moles of a gas



Subject Test Chemistry

Thermodynamics

Strategy Notes

APPLYME

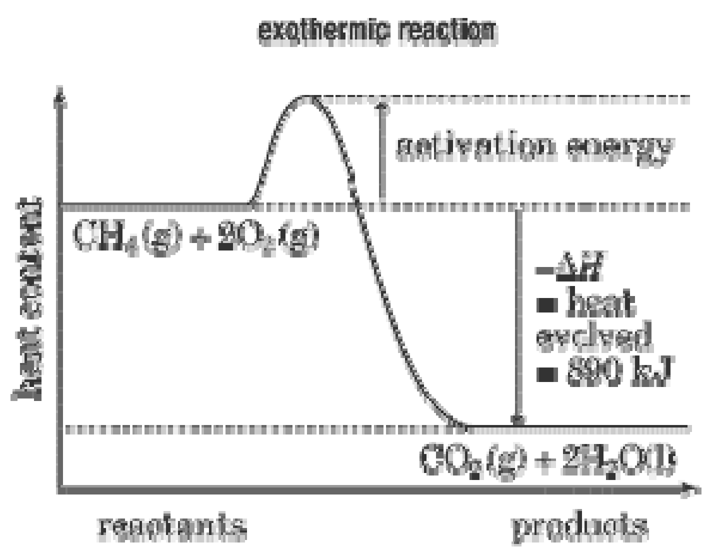
The Basics

- Law of conservation of energy (also known as the first law of thermodynamics): Energy can neither be created nor destroyed in the course of a chemical reaction
- Energy (E, joules): the ability to do work or produce heat
- Heat (q, joules): the transfer of energy in a physical or chemical process: heat always flows from a warmer object to a cooler one
- Internal energy of the system: sum of all of the potential and kinetic energy in a system
- Forms of energy
  - Potential energy: the energy stored in chemical bonds
  - Kinetic energy: exists in matter in motion
- The energy of particles is proportional to the temperature (K) of the system as well as the mass and the velocity of the object: $KE = \frac{1}{2}mv^2$

Enthalpy

- Enthalpy (H): the change in the heat in the system from chemical changes resulting in either the release or the absorption of heat
- Endothermic: a reaction in which there is a net absorption of heat energy; energy is a reactant, and the change in enthalpy of the system, ΔH, has a positive value
- Exothermic: a reaction in which there is a net production of heat by the system; energy is a product, and the change in enthalpy of the system, ΔH, has a negative value

Enthalpy



Enthalpy

- Different forms of enthalpy
  - Enthalpy of reaction (ΔH_{rxn}): The amount of heat absorbed or released by the chemical reaction
  - Enthalpy of combustion (ΔH_{comb}): The amount of heat absorbed or released by combustion (burning; usually in the presence of O_2)
  - Enthalpy of formation (ΔH_f): The amount of heat absorbed or released when 1 mole of a compound is formed from elements in their standard states
  - Enthalpy of fusion (ΔH_{fus}): The amount of heat that must be absorbed to melt 1 mole of solid to liquid at the normal melting point
  - Enthalpy of vaporization (ΔH_{vap}): The amount of heat that must be absorbed to change 1 mole of liquid to gas at the normal boiling point

Enthalpy

- Enthalpy is a state function, meaning that its value is fixed when temperature, pressure, composition, and physical form are specified
- At a constant pressure, $\Delta H = q$, meaning that at constant pressure, the enthalpy of a system is equal to the heat, in joules, of a system
- Enthalpy changes of a reaction can be calculated (by stoichiometry, calorimetry, tables of standard values, Hess's law, and the bond energies of the substances involved)

Spontaneous Reactions

- Spontaneous reaction occurs without being driven by some outside force
- Driving forces for all chemical reactions.
  - Enthalpy
  - Entropy (ΔS): a measure of the disorder of a system, and systems tend to favor a more disordered system: nature tends toward chaos.
- Spontaneous reactions occur without outside intervention
- Examples: combustion of hydrogen and graphite turning into diamond

Spontaneous Reactions

Example:

The addition of 14.0 g solid potassium hydroxide pellets to water causes the following reaction to take place:

$$KOH_{(s)} \rightarrow KOH_{(aq)} + 43\ kJ$$

1. Does the beaker get warmer or colder as the reaction takes place?
2. Is the reaction endothermic or exothermic?
3. What is the enthalpy change for the dissolution of the 14.0 grams of KOH?

Heat Capacity and Specific Heat

- Heat capacity (joules or calories): the amount of heat energy it must consume in order to raise its temperature by 1K or 1ºC
- Calorie: the amount of heat needed to raise the temperature of 1.00 gram of water by 1.00ºC
- 1 calorie = 4.184 joules
- Every pure substance involved in a chemical reaction has a unique heat capacity
- Molar heat capacity: the heat capacity of 1 mol of a pure substance (J/mol-K or J/mol-ºC)

Heat Capacity and Specific Heat

- Specific heat: heat capacity of 1 gram of a substance (J/g-K)

$$\boxed{q = mC_p\mathrm{D}T}$$

  - q = quantity of heat (joules or calories)
  - m = mass in grams
  - DT = T_f - T_i (final – initial)
  - C_p = specific heat capacity (J/g ºC)
- The specific heat of liquid water is 4.184 J/g ºC (or 1.00 cal/g ºC)

Heat Capacity and Specific Heat

Example:

100.0 mL of 1.0 M NaOH and 100.0 mL of 1.0 M HCl are mixed in a calorimeter. Both solutions were originally at 24.6ºC. After the reaction, the final temperature is 31.3ºC. Assuming that the solution has a density of 1.0 g/cm^3 and a specific heat capacity of 4.184 J/g ºC, calculate the enthalpy change for the neutralization of HCl by NaOH. Assume that no heat is lost to the surroundings or the calorimeter.

Heat Capacity and Specific Heat

Example:

How much energy would be needed to heat 450 grams of copper metal from a temperature of 25.0ºC to a temperature of 75.0ºC? The specific heat of copper at 25.0ºC is 0.385 J/g ºC.

Enthalpies of Reactions

- Overall enthalpy of reaction under the thermodynamic standard states 25C (298K), 1 atm, and 1 M

$$\Delta \boldsymbol{H} = \boldsymbol{H}_{\text{final}} - \boldsymbol{H}_{\text{initial}}$$

or

$$\Delta \boldsymbol{H} = H_{\text{products}} - H_{\text{reactants}}$$

Enthalpies of Reactions

Example:

Calculate the ΔH for the following:

$$3Al_{(s)} + 3NH_4ClO_{4(s)} \rightarrow Al_2O_{3(s)} + AlCl_{3(s)} + 3NO_{(g)} + 6H_2O_{(g)}$$

given the following values:

Substance	$H_f°$(kJ/mol)
$NH_4ClO_{4(s)}$	-295
$Al_2O_{3(s)}$	-1676
$AlCl_{3(s)}$	-704
$NO_{(g)}$	90
$H_2O_{(g)}$	-242
$Al_{(x)}$	0 (since it's an element)

Enthalpies of Reactions

Example:

Find the ΔH_f of $C_6H_{12}O_{6(s)}$ using the following information:

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)} + 2800\ kJ$$

Substance	$H_f°$(kJ/mol)
$CO_{2(g)}$	-393.5
$H_2O_{(l)}$	-285.8

Hess's Law

- The total enthalpy of a reaction is independent of the reaction pathway
- If a reaction is carried out in a series of steps, the enthalpy change (ΔH) for the overall reaction will be equal to the sum of the enthalpy changes for the individual steps
- Some rules:
  - Make sure to rearrange the given equations so that reactants and products are on the appropriate sides of the arrows
  - If you reverse equations, you must also reverse the sign of D*H*
  - If you multiply equations to obtain a correct coefficient, you must also multiply the D*H* by this coefficient

Hess's Law

Example:

Given the following equations

$H_3BO_{3(aq)} \rightarrow HBO_{2(aq)} + H_2O_{(l)}$ DH_{rxn} = -0.02 kJ

$H_2B_4O_{7(aq)} + H_2O_{(l)} \rightarrow 4HBO_{2(aq)}$ DH_{rxn} = -11.3 kJ

$H_2B_4O_{7(aq)} \rightarrow 2B_2O_{3(s)} + H_2O_{(l)}$ DH_{rxn} = 17.5 kJ

find the D*H* for this overall reaction:

$2H_3BO_{3(aq)} \rightarrow B_2O_{3(s)} + 3H_2O_{(l)}$

Bond Energies

- Energy must be added or absorbed to break bonds and that energy is released when bonds are formed

ΔH = bonds broken - bonds formed

Bond Energies

Example:

Using bond energies, calculate the change in energy that accompanies the following reaction:

$$H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$$

Bond type	Bond energy
H—H	432 kJ/mol
F—F	154 kJ/mol
H—F	565 kJ/mol

More About Entropy

- 1st law of thermodynamics: Energy can neither be created nor destroyed.
- 2nd law of thermodynamics: The disorder of the universe, its entropy, or ΔS, is constantly increasing.
- 3rd law of thermodynamics: The entropy of a perfect crystal at 0K is zero.

More About Entropy

1. The greater the disorder or randomness in a system, the larger the entropy.
2. The entropy of a substance always increases as it changes state from solid to liquid to gas.
3. When a pure solid or liquid dissolves in a solvent, the entropy of the substance increases.
4. When a gas molecule escapes from a solvent, there is an increase in entropy.
5. Entropy generally increases with increasing molecular complexity.
6. Reactions that increase the number of moles of particles often increase the entropy of the system.

More About Entropy

Example:

Which of the following reactions results in the largest increase in entropy?

(A) $CO_{2(s)} \rightarrow CO_{2(g)}$

(B) $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

(C) $KNO_{3(s)} \rightarrow KNO_{3(l)}$

(D) $C_{(diamond)} \rightarrow C_{(graphite)}$

More About Entropy

- Calculate entropy in J/K

$$\Delta S_{sys}^{\circ} = \sum \Delta S^{\circ}{}_{products} - \sum \Delta S^{\circ}{}_{reactants}$$

- The higher the *S* value, the more disordered the system
- A positive (+) *S* value is more disordered, and a –*S* value is less disordered

More About Entropy

Example:

Calculate the entropy change at 25ºC in J/K for

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

given the following data:

$SO_{2(g)}$: 248.1 J/mol-K

$O_{2(g)}$: 205.3 J/mol-K

$SO_{3(g)}$: 256.6 J/mol-K

Gibb's Free Energy

- Gibb's free energy (G) equation helps to determine if a reaction is spontaneous

$$\boxed{\Delta G = \Delta H - T\Delta S}$$

  - If G is negative, the reaction is spontaneous in the forward direction.
  - If G is equal to zero, the reaction is at equilibrium.
  - If G is positive, then the reaction is nonspontaneous in the forward direction, but the reverse reaction will be spontaneous.
  - $G_f^\circ = 0$ for elements at standard state (pure elements at 25ºC and 1 atm are assigned a value of zero).

Gibb's Free Energy

Example:

Find the thermodynamic boiling point of

$$H_2O_{(l)} \rightarrow H_2O_{(g)}$$

given the following information:

$$H_{vap} = +44 \text{ kJ} \qquad S_{vap} = 118.8 \text{ J/K}$$

Gibb's Free Energy

- Calculate change in Gibb's Free Energy in J/K

$$\Delta G_{sys}^{\circ} = \sum \Delta G^{\circ}{}_{products} - \sum \Delta G^{\circ}{}_{reactants}$$

ΔH	ΔS	**Result**
Negative	Positive	Spontaneous at all temperatures
Positive	Positive	Spontaneous at high temperatures
Negative	Negative	Spontaneous at low temperatures
Positive	Positive	*Never* spontaneous

Gibb's Free Energy

Example:

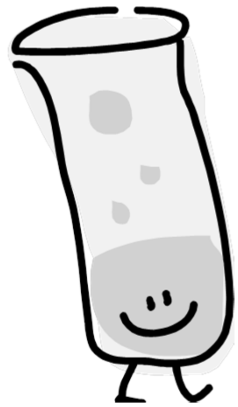
Find the free energy of formation for the oxidation of water to produce hydrogen peroxide.

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

given the following information:

	ΔG_f°
$H_2O_{(l)}$	-56.7 kcal/mol
$O_{2(g)}$	0 kcal/mol
$H_2O_{2(l)}$	-27.2 kcal/mol

Subject Test Chemistry



Descriptive Chemistry

Strategy Notes

APPLYME

Organic Chemistry

- Organic chemistry: the branch of chemistry that studies carbon compounds
- One reason that there are so many carbon compounds is because of carbon’s unique ability for bonding
- Carbon has four valence electrons, and these electrons can hybridize into *sp*^{3}, *sp*^{2}, and *sp* atomic orbitals
- Enables carbon to join with other elements and be involved in single, double, and triple bonds
- Hydrocarbons: organic compounds that contain only carbon and hydrogen

Organic Chemistry

- They usually have low melting points.
- They usually are nonpolar (unless they bear functional groups).
- They are usually nonconductors of electricity.
- They can exist in solid, liquid, and gaseous form.

Organic Chemistry

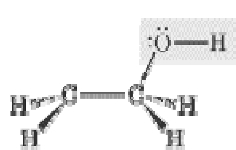
- Compounds with:
  - 1–4 carbons tend to be gases at room temperature; butane and propane are among the lightest hydrocarbons and are used for fuel
  - 5–10 carbons tend to be in the liquid state at room temperature; compounds that fall in this size range are used to make gasoline and solvents
  - 12–18 carbons make up jet fuels and kerosene
  - More than 18 carbons tend to be solids at room temperature

Organic Chemistry

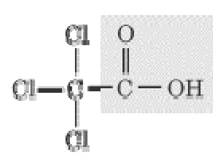
- Organic compounds can exist as polymers, in which many repeating units (called monomers) make up a larger molecule
- Amino acids: monomers of proteins when amino acids are bonded in a chain, they make a polypeptide or protein
- Starches: polymers of the monomer glucose
- Plastics: polymers of organic molecules extracted from crude oil
  - Polyethylene—Many ethenes strung together with covalent bonds (ethylene is another name for ethene); shopping bags and plastic bottles are made of polyethylene
  - Polypropylene—Many propenes strung together; glues and carpets are made of polypropylene
  - Polystyrene—A clear, hard, brittle polymer used in CD cases; if you blow carbon dioxide into it during manufacture and you get the soft, opaque, foamy polymer used in a coffee cup

Common Functional Groups

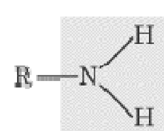
- Functional groups: atoms or groups of atoms attached to an organic compound that impart characteristic shapes and chemical properties to the compound



Hydroxyl (-OH) group alcohols



Carboxylic acid (-COOH) group



Amine ($-NH_2$) group

Isomerism

- Isomers: compounds that have the same number and kinds of atoms but have different structures—meaning that the atoms are arranged differently in the molecule
- Generally the number of isomers increases dramatically as the number of carbon atoms increases because there are more options for molecular structure
  - Different carbon skeletons (one or more bonds differ)
  - Different functional groups
  - Different positions of functional groups

Simple Organic Reactions

- Complete combustion

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

- Incomplete combustion

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$$

- Addition reaction

$$H_2C{=}CH_2 + H_2 \rightarrow H_3C\text{-}CH_3$$

- Substitution reaction

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

Group 1A (Alkali Metals)

- Most active metals on the periodic table
- React with water at room temperature to form bases
- React readily with acids to produce hydrogen gas
- Get even more reactive as you move down the family → more energy levels, more shielding, so it's harder for the nucleus to hold on to the lonely valence electron

Group 7A (Halogens)

- Most reactive nonmetals on the periodic table
- All of these elements are diatomic
- Fluorine is a gas, bromine is a liquid, and iodine is a solid → as the molecules get larger, there are more intermolecular forces to hold them together
- Fluorine is the most reactive of the halogens
- Chlorine is a very common antibacterial agent, found in bleach and muriatic acid (HCl), and is added to every city's water supply

Group 8A (Noble Gases)

- Most stable family on the periodic table
- Many of these gases appear in signs (such as neon signs)
- Helium is used to fill balloons because it is much less dense than air
- Argon is fairly abundant in our atmosphere

Metals

- Metals have a positive center surrounded by a sea of electrons → very good conductors of electricity
- Alloys: contain a mixture of elements that have metallic properties; often much stronger than the individual metal itself
- Some of the more common alloys include:
  - Brass: mixture of copper and zinc
  - Sterling silver: mixture of silver and copper
  - Steel: mixture of iron and carbon
  - Bronze: mixture of copper, zinc, and other metals
  - Pewter: mixture of tin, copper, bismuth, and antimony

Properties of Some Common Gases

- Hydrogen
  - Colorless, odorless gas
  - When hydrogen gas is collected in a test tube in the lab, a burning splint inserted into the test tube filled with hydrogen will “bark” as the hydrogen ignites
- Oxygen
  - Makes up about 21% of our atmosphere
  - Colorless, odorless gas that is necessary for life
  - Supports combustion reactions
  - When oxygen is collected in a test tube in the laboratory, a glowing wooden splint will reignite

Properties of Some Common Gases

- Carbon dioxide
  - Colorless, odorless gas that does not support combustion
  - When carbon dioxide gas is collected in a test tube in the laboratory, a burning wooden splint will go out when placed into the gas
  - Another common lab test for CO_2 is to bubble it into limewater, $Ca(OH)_2$. The clear solution will turn cloudy as calcium carbonate, $CaCO_3$, begins to precipitate
- Chlorine
  - Deadly yellow-green gas
  - Often been used as a weapon in warfare

Environmental Chemistry: Fuels

- Coal is solid and is composed of large hydrocarbons and other compounds that contain sulfur, oxygen, and nitrogen
- When it's combusted, the sulfur it contains is converted to SO_2, which is an air pollutant
- Petroleum is a liquid made up of hundreds of different components, but mostly hydrocarbons
- It also contains some compounds that have functional groups containing sulfur, nitrogen, or oxygen
- Natural gas consists of hydrocarbons in the gas phase, primarily methane (CH_4)

Air Pollution

- Carbon monoxide (CO): produced from incomplete combustion of all types of natural and synthetic products, including cigarette smoke. When it builds up in high concentrations, it can be very toxic. Cities with heavy traffic problems are known for dangerous CO levels.
- Carbon dioxide (CO_2): the principal greenhouse gas and is primarily responsible for the greenhouse effect. It can be formed from all types of common human activity, such as burning fuels and even breathing.

Air Pollution

- Chlorofluorocarbons (CFCs): used in great quantities in industry, for refrigeration and air-conditioning, and in consumer products. When released into the air, they rise into the stratosphere, where they readily react with the ozone that constitutes the ozone layer, effectively degrading it
- Ozone (O_3): occurs naturally in the upper atmosphere, where it shields the earth from the sun's dangerous ultraviolet rays. When found at ground level (vehicle exhaust and industry waste), however, it's a pollutant. It can cause damage to humans (especially our respiratory system), the environment, and a wide range of natural and artificial materials

Air Pollution

- Nitrogen oxide (NO_x) and sulfur dioxide (SO_x): major contributors to smog and acid rain. These gases both react with volatile organic compounds to form smog, which can cause respiratory problems in humans. Acid rain can harm vegetation, change the chemistry of river and lake water by lowering the pH so that it's harmful to animal life, and react with the marble of statues and buildings and decompose them.