Chapter 1– ATOMIC CONCEPTS

I. Subatomic Particles:

- Scientists found that you can breakdown atoms into smaller parts called subatomic particles. 3 most important are *electrons, protons, and neutrons*.

	Protons	**Neutrons**	**Electrons**
Charge	positive	none	electrons
Mass/Location	1 amu / Nucleus	1 amu / Nucleus	0 amu / Orbital

- **An atoms identity is defined entirely by the number of protons in the nucleus; the number of protons of any given element NEVER changes.**
- **ATOM = NEUTRAL**

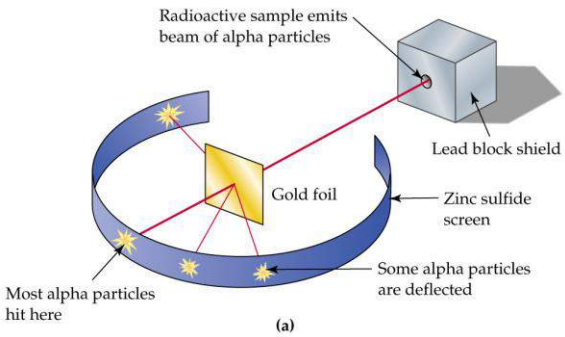
ATOMIC MODELS: helps scientists imagine on a macroscopic level what happens microscopically. Three contrasting models include:

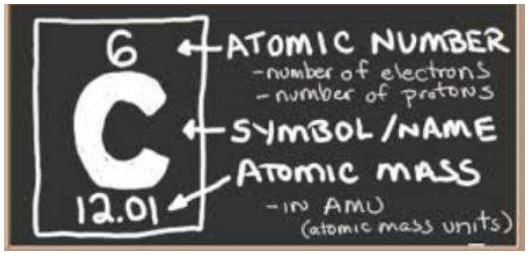
1. **Rutherford's Model:** Most of the mass of the atom is in the center (nucleus) is positive. Protons are in the nucleus, most of the atom is empty space. Electrons go around the nucleus
2. **Bohr's Model:** Protons are in the nucleus, which is positive. Electrons revolve around the nucleus in concentric orbits.
3. **Orbital Model (modern model):** Electron Cloud Model. Electrons located in ORBITALS which is the most probable location of an electron.

Rutherford's Gold Foil Experiment:

- When he bombarded the foil, most of the particles (alpha particles which are positively charged) went straight through the foil, BUT, some of the particles bounced back. Two very important results were concluded by this experiment.

1. **An atom is made up of mostly empty space**
2. **An atom has a nucleus that is positively charged.**





ATOMIC NUMBER: Located on the lower left hand in the box of the individual element on the Periodic Table. The atomic

MASS NUMBER: Located on the upper left corner in the box of the individual element on the Periodic Table is equal to the total number of particles in the NUCLEUS (PROTONS + NEUTRONS)

 # Neutrons = Mass Number – Atomic Number

ISOTOPES: atoms of the same element that have different numbers of neutrons

AVERAGE ATOMIC MASS: the average of all the NATURALLY occurring isotopes of a given element.

EXAMPLE:

1. Calculate the average atomic mass of potassium using the following data:

Isotope	Mass	% abundance
Potassium-39	38.964 amu	93.12%
Potassium-41	40.962 amu	6.88 %

Potassium-39 38.964 amu x 0.9312 = 36.28 amu

Potassium-41 40.962 amu x 0.0688 = 2.82 amu +

Average atomic mass for K = 39.10 amu

PRINCIPLE ENERGY LEVELS: The energy level shows how far the electron is from the nucleus the first energy level is closest to the nucleus and the others are further away. Electrons in the first level have the lowest energy and the energy of the electron increases as the levels increase.

- FIRST PRINCIPLE ENERGY LEVEL: holds only 2 electrons.
- SECOND PRINCIPLE ENERGY LEVEL: holds only 8 electrons.
- THIRD PRINCIPLE ENERGY LEVEL: holds only 18 electrons.
- FOURTH PRINCIPLE ENEGY LEVEL: holds only 32 electrons.

ELECTRON CONFIGURATION: Located on the lower left corner, below the atomic number. It shows how many electrons are in each principle energy level.

VALENCE ELECTRON: The number of electrons in the last principle energy level. According to the **octet rule,** there can be no more than 8 valence electrons. All other electrons in an atom other than the last level (valence) and called **non-valence electrons.** Example: Mg 2-8-2. There are 2 valence electrons and 10 non-valence electrons.

GROUND & EXCITED STATES: An atom is in the ground state when the electrons are filling the atom in the order 2-8-18-32 like it is written in the periodic table that shows electrons in the ground state

- When an electron goes from the ground to excited, energy is absorbed.
- When an electron goes from excited back down to ground, then energy is released in the form of the bright-line spectrum (color)
- Every element gives off a different amount of energy (like a fingerprint).

IONS: Charged particles.

- **ANION:** negatively charged ions (gains electrons)
- **CATION:** positively charged ions (loses electrons)

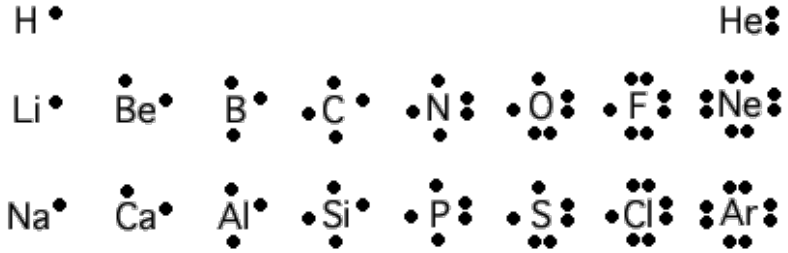
Metals

Atoms	Number of Protons = Number of Electrons	Ion	# Electons
Li	3	Li^{+}	2
Ca	20	Ca^{+2}	18

Non-Metals

Atoms	Number of Protons = Number of Electrons	Ion	# Electons
F	9	F^{-}	10
S	16	S^{2-}	18

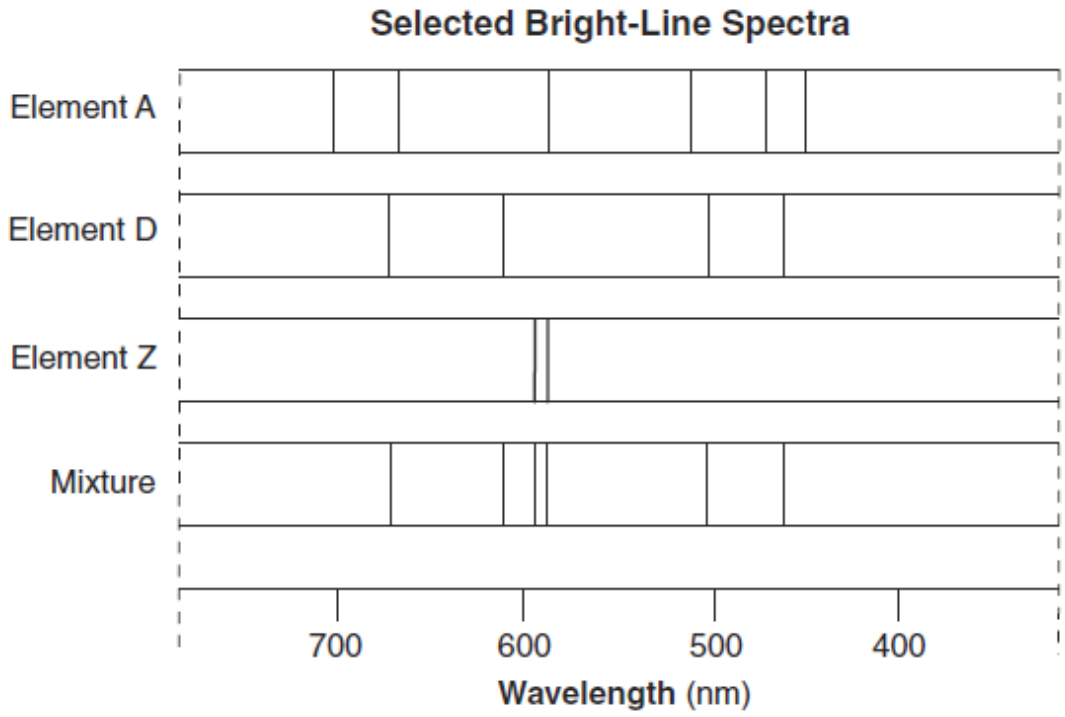
LEWIS ELECTRON DOT DIAGRAMS: Helps to show the number of **<u>valence</u>** electrons in the last principle energy level of an atom.



Name: **Chapter 1 Atomic Review** **Constructed Response**

Base your answers to questions **1** through **3** on the information below and on your knowledge of chemistry.

The bright-line spectra observed in a spectroscope for three elements and a mixture of two of these elements are represented in the diagram below.



1. Describe, in terms of *both* electrons and energy state, how the light represented by the spectral lines is produced.

2. Explain why the spectrum produced by a 1-gram sample of element Z would have the same spectral lines at the same wavelengths as the spectrum produced by a 2-gram sample of element Z.

3. State evidence from the bright-line spectra that indicates element *A* is *not* present in the mixture.

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

Illuminated **EXIT** signs are used in public buildings such as schools. If the word **EXIT** is green, the sign may contain the radioisotope tritium, hydrogen-3. The tritium is a gas sealed in glass tubes. The emissions from the decay of the tritium gas cause a coating on the inside of the tubes to glow.

State, in terms of neutrons, how an atom of tritium *differs* from an atom of hydrogen-1.

Base your answers to questions **5** through **8** on
the information below and on your knowledge of chemistry.

A student compares some models of the atom. These models are listed in the table below in order of development from top to bottom.

Models of the Atom

Model	Observation	Conclusion
Dalton model	Matter is conserved during a chemical reaction.	Atoms are hard, indivisible spheres of different sizes.
Thomson model	Cathode rays are deflected by magnetic/electric fields.	Atoms have small, negatively charged particles as part of their internal structure.
Rutherford model	Most alpha particles pass straight through gold foil but a few are deflected.	An atom is mostly empty space with a small,dense,positively as part of their nucleus.
Bohr model	Unique spectral lines are emitted by excited gaseous elements.	Packets of energy are absorbed or emitted by atoms when an electron changes shells.

5. State *one* way in which the Bohr model agrees with the Thomson model.

6. Using the conclusion from the Rutherford model, identify the charged subatomic particle that is located in the nucleus.

7. State *one* conclusion about the internal structure of the atom that resulted from the gold foil experiment.

8. State the model that first included electrons as subatomic particles.

Base your answers to questions **9** through **11** on
the information below

The element boron, a trace element in Earth's crust, is found in foods produced from plants. Boron has only two naturally occurring stable isotopes, boron-10 and boron-11.

9. State, in terms of subatomic particles, *one* difference between the nucleus of a carbon-11 atom and the nucleus of a boron-11 atom.

10. Write an isotopic notation of the heavier isotope of the element boron. Your response must include the atomic number, the mass number, and the symbol of this isotope.

11. Compare the abundance of the two naturally occurring isotopes of boron.

Answer Key

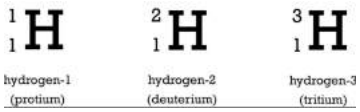
atomic Constructed response

1. —Different colors of light are produced when electrons return from higher energy states to lower energy states. —Light energy can be emitted when electrons in excited atoms return to lower shells. —Electrons release energy as they move toward the ground state.
2. —The wavelengths of the spectral lines for element Z are independent of the mass of the sample. —All atoms of element Z have the same electron configuration in the ground state. —The intensive properties of an element remain constant.
3. —Not all of the wavelengths of element A are shown in the wavelengths of the mixture. —The mixture has no spectral line at 700 nm.
4. –A tritium atom has two neutrons and an H-1 atom has no neutrons. –Only the tritium atom has neutrons. –H-1 has no neutrons.
5. –Atoms have electrons. –Atoms have small, negatively charged particles. –Both models show an internal structure. –Atoms are neutral.
6. –proton –p –p^+ –$_1$1p –$_1{}^1$H –H^+
7. –An atom is mainly empty space. –It has a nucleus. –The small, dense nucleus is positively charged.
8. –Thomson model –Thomson –plum pudding model
9. –The carbon-11 nucleus has one more proton than the nucleus of boron-11. –A B-11 atom has a different number of neutrons than a C-11 atom.
10. $^{11}_{5}\mathrm{B}$
11. –Boron-11 is about four times more abundant than boron-10. –The B-10 is less abundant.

Chapter 2 – NUCLEAR CHEMISTRY

- The ratio of neutrons to protons is what determines whether a nucleus is stable or unstable.
- For elements whose atomic numbers are small (1-20), if the ratio of neutrons to protons (neutrons/protons) is **about 1,** the nucleus of the **isotope are stable.** Remember that isotopes have the same number of protons but different numbers of neutrons, or the same atomic number but different atomic mass.

RADIOACTIVITY: in radioactivity, the nucleus of an unstable isotope or element decays spontaneously and gives off rays and particles, which is also known as decay. **The symbols for decay are listed below, the number on the top left indicates the mass of the decay, and the number on the bottom left indicates the charge of the decay.**



- **Transmutation:** when the nucleus of an atom decays and one element changes into another element. The mass and charge have to be equal on both sides. Table N has the decay modes for radioactive isotopes that undergo **Natural Transmutation**, which means the decay (breakdown of the nucleus) occurs spontaneously.

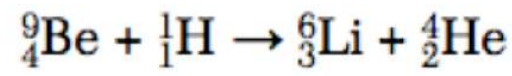
Table N
Selected Radioisotopes

Nuclide	Half-Life	Decay Mode	Nuclide Name
^{198}Au	2.695 d	β^-	gold-198
^{14}C	5715 y	β^-	carbon-14
^{37}Ca	182 ms	β^+	calcium-37
^{60}Co	5.271 y	β^-	cobalt-60
^{137}Cs	30.2 y	β^-	cesium-137
^{53}Fe	8.51 min	β^+	iron-53
^{220}Fr	27.4 s	α	francium-220
^{3}H	12.31 y	β^-	hydrogen-3
^{131}I	8.021 d	β^-	iodine-131
^{37}K	1.23 s	β^+	potassium-37
^{42}K	12.36 h	β^-	potassium-42
^{85}Kr	10.73 y	β^-	krypton-85
^{16}N	7.13 s	β^-	nitrogen-16
^{19}Ne	17.22 s	β^+	neon-19
^{32}P	14.28 d	β^-	phosphorus-32
^{239}Pu	2.410×10^4 y	α	plutonium-239
^{226}Ra	1599 y	α	radium-226
^{222}Rn	3.823 d	α	radon-222
^{90}Sr	29.1 y	β^-	strontium-90
^{99}Tc	2.13×10^5 y	β^-	technetium-99
^{232}Th	1.40×10^{10} y	α	thorium-232
^{233}U	1.592×10^5 y	α	uranium-233
^{235}U	7.04×10^8 y	α	uranium-235
^{238}U	4.47×10^9 y	α	uranium-238

Source: *CRC Handbook of Chemistry and Physics*, 91st ed., 2010–2011, CRC Press

$$^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$$

Natural Transmutation



$$^{9}_{4}Be + {}^{1}_{1}H \rightarrow {}^{6}_{3}Li + {}^{4}_{2}He$$

Artificial Transmutation

- **Artificial Transmutation:** elements can be made radioactive by bombarding their nucleus with high energy particles. In natural transmutation, the element will change into another element when the nucleus decays. In artificial transmutation, the same thing occurs except **not** spontaneously. Remember that the atomic # and the mass # have to equal the same thing on both sides.

Table O:

Table O
Symbols Used in Nuclear Chemistry

Name	Notation	Symbol
alpha particle	$^{4}_{2}He$ or $^{4}_{2}\alpha$	α
beta particle	$^{0}_{-1}e$ or $^{0}_{-1}\beta$	β^-
gamma radiation	$^{0}_{0}\gamma$	γ
neutron	$^{1}_{0}n$	n
proton	$^{1}_{1}H$ or $^{1}_{1}p$	p
positron	$^{0}_{+1}e$ or $^{0}_{+1}\beta$	β^+

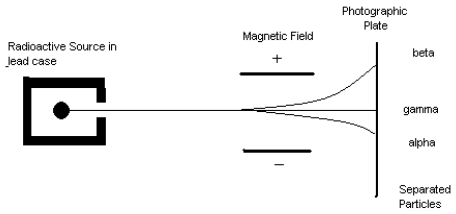
Weakest penetrating power-HEAVIEST

Strongest penetrating power-LIGHTEST

- Number on the upper left is the mass
- Number on the lower left is the charge

SEPARATING ALPHA, BETA, & GAMMA PARTICLES:

- Can be separated by using an electric or magnetic field. In an electric field, an alpha particle, which is positively charged (has 2 protons), is deflected toward the negative electrode. A beta particle is negatively charged, and will be deflected towards the positive electrode. Gamma rays have no charge, and therefore are not deflected, there is no bend.

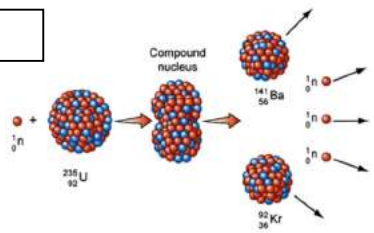


Nuclear Energy: in a nuclear reaction, mass is converted into energy. Two types are fission and fusion.

- **Fission:** Type of artificial transmutation. A neutron bombards an atom causing it to split into two or more pieces and gives off a lot of energy. This is the energy behind nuclear power plants (controlled chain reaction) and atomic bombs (uncontrolled chain reaction).

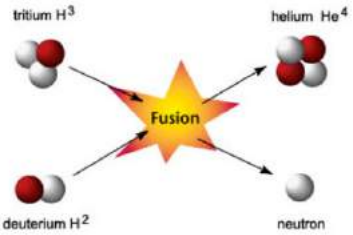
- **Fusion:** two nuclei unite to form a heavier nucleus ("u"- unit). High temperature and pressure are needed, which is why this occurs on the SUN. **Fusion creates more energy than fission.**

Fission



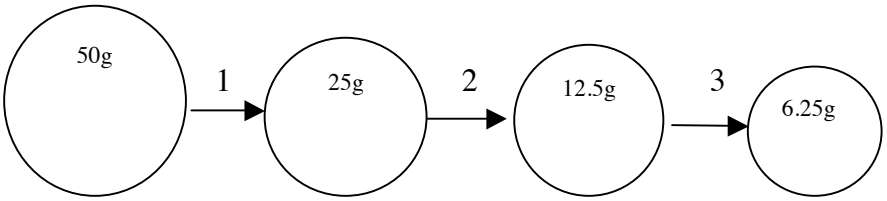
$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{93}_{36}Kr + ^{140}_{56}Ba + 3^{1}_{0}n$$

Fusion



$$^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n$$

HALF-LIFE: Each radioactive isotope has its own rate of decay. Half-life is the time it takes a sample to decay in half. Table N lists some common half-life rates for some common isotopes. For every problem it's important to determine how many half-lives occurred.



QUESTIONS:

Table N
Selected Radioisotopes

Nuclide	Half-Life	Decay Mode	Nuclide Name
^{198}Au	2.695 d	β^-	gold-198
^{14}C	5715 y	β^-	carbon-14
^{37}Ca	182 ms	β^+	calcium-37
^{60}Co	5.271 y	β^-	cobalt-60

1. What is the half-life?
   - $\frac{\text{Total time elapsed}}{\text{\# half-life series}}$
2. How much total time elapsed?
   - Half-life x # half-life series
3. How many grams will remain after …?
   - Must determine how many half-life series and then half the original amount by that many times (see example above)
4. What fraction remains?
   - $(1/2)^{\#\text{half-lives}}$ – knowing the fraction remaining will give you the number of half-life series.

$(1/2)^1 = ½$ $(1/2)^4 = 1/16$
$(1/2)^2 = ¼$ $(1/2)^5 = 1/32$
$(1/2)^3 = 1/8$ $(1/2)^6 = 1/64$

1. A radioactive isotope has a half-life of 2.5 years. Which fraction of the original mass remains unchanged after 10. years?

   A) 1/2 B) 1/4 C) 1/8 D) 1/16

2. After decaying for 48 hours, $\frac{1}{16}$ of the original mass of a radioisotope sample remains unchanged. What is the half-life of this radioisotope?

   A) 3.0 h B) 9.6 h C) 12 h D) 24 h

3. What is the half-life of a radioisotope if 25.0 grams of an original 200.-gram sample of the isotope remains unchanged after 11.46 days?

   A) 2.87 d B) 3.82 d
   C) 11.46 d D) 34.38 d

4. What is the half-life of sodium-25 if 1.00 gram of a 16.00-gram sample of sodium-25 remains unchanged after 237 seconds?

   A) 47.4 s B) 59.3 s C) 79.0 s D) 118 s

5. How many days are required for 200. grams of radon-222 to decay to 50.0 grams?

   A) 1.91 days B) 3.82 days
   C) 7.64 days D) 11.5 days

6. After 32 days, 5 milligrams of an 80-milligram sample of a radioactive isotope remains unchanged. What is the half-life of this element?

   A) 8 days B) 2 days
   C) 16 days D) 4 days

CHEMICAL VS. NUCLEAR ENERGY:

- Nuclear reactions release more energy than chemical reactions.
- **Benefits:** nuclear provides a lot of energy (produces the energy of the sun), less CO_2 is produced from nuclear power than fossil fuel combustion.
- **Risks:** Wastes from are nuclear power very radioactive, must be stored for more than 100,000 years without leaking into the ground (long half-lives), accidents can cause radioactive spills (mutation/death).

RADIOACTIVE ISOTOPES (RADIOISOTOPES):

- Tracers:
  - **Carbon-14 – Date LIVING THINGS**
  - **Uranium-238 & Lead 206 – Date NON-LIVING THINGS**
- Medical: isotopes with very short half-lives can be eliminated by the body quickly.
  - **Technetium-99 – brain tumors**
  - **Iodine-131 – thyroid disorders**
  - **Radium & Cobalt-60 – treatment of cancer**

Risks:

- Biological Damage: exposure can damage or destroy cells – mutation.
- Long-term storage: must be stored in special containers for a long period of time, is it safe?
- Accidents: 1986 – Chernobyl, 1979—3 Mile Island & Pollution: radioactive materials in air, water, food, and soil.

Name: **Chapter 2 Nuclear Review**

Base your answers to questions **1** through **4** on the information below and on your knowledge of chemistry.

A breeder reactor is one type of nuclear reactor. In a breeder reactor, uranium-238 is transformed in a series of nuclear reactions into plutonium-239.

The plutonium-239 can undergo fission as shown in the equation below. The *X* represents a missing product in the equation.

$$^{1}_{0}\mathrm{n} + {}^{239}_{94}\mathrm{Pu} \rightarrow X + {}^{94}_{36}\mathrm{Kr} + 2{}^{1}_{0}\mathrm{n}$$

1. Write a notation for the nuclide represented by missing product *X* in this equation.

2. Compare the amount of energy released by 1 mole of completely fissioned plutonium-239 to the amount of energy released by the complete combustion of 1 mole of methane.

3. Based on Table *N,* identify the decay mode of the plutonium radioisotope produced in the breeder reactor.

4. Determine the number of neutrons in an atom of the uranium isotope used in the breeder reactor.

Base your answers to questions **5** and **6** on the information below and on your knowledge of chemistry.

Illuminated **EXIT** signs are used in public buildings such as schools. If the word **EXIT** is green, the sign may contain the radioisotope tritium, hydrogen-3. The tritium is a gas sealed in glass tubes. The emissions from the decay of the tritium gas cause a coating on the inside of the tubes to glow.

5. Complete the nuclear equation for the radioactive decay of tritium, by writing a notation for the missing product.

$$^{3}_{1}\mathrm{H} \rightarrow {}^{0}_{-1}\mathrm{e} + \underline{\qquad\qquad}$$

6. Determine the fraction of an original sample of tritium that remains unchanged after 24.62 years.

Base your answers to questions **7** through **10** on the information below.

Nuclear radiation is harmful to living cells, particularly to fast-growing cells, such as cancer cells and blood cells. An external beam of the radiation emitted from a radioisotope can be directed on a small area of a person to destroy cancer cells within the body.
Cobalt-60 is an artificially produced radioisotope that emits gamma rays and beta particles. One hospital keeps a l00.0-gram sample of cobalt-60 in an appropriate, secure storage container for future cancer treatment.

7. Determine the total time that will have elapsed when 12.5 grams of the original Co-60 sample at the hospital remains unchanged.

8. Complete the nuclear equation below for the beta decay of the Co-60 by writing an isotopic notation for the missing product.

$$^{60}_{27}\mathrm{Co} \rightarrow ^{0}_{-1}\beta$$

9. Compare the penetrating power of the two emissions from the Co-60.

10. State *one* risk to human tissue associated with the use of radioisotopes to treat cancer.

Answer Key

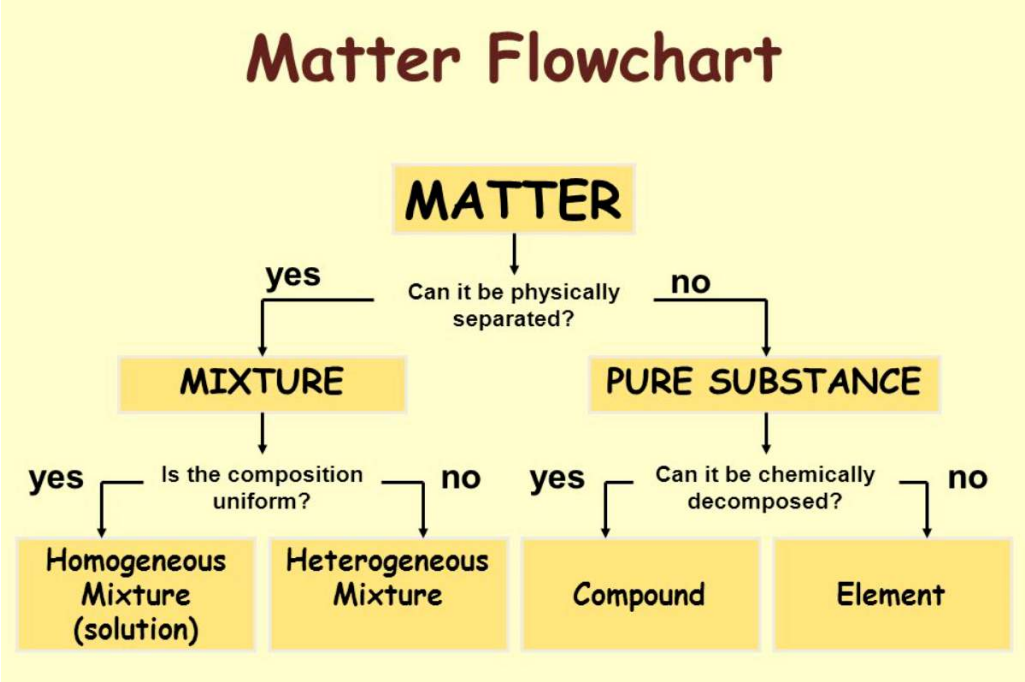
Nuclear Part 2

1. –$^{144}_{58}$Ce –^{144}Ce –cerium-144 –Ce-144
2. –The fission of one mole of Pu-239 releases much more energy than the combustion of one mole of CH_4. –The energy released during the chemical reaction is less than the energy released during the nuclear reaction. –greater for $^{239}_{94}\mathrm{Pu}$
3. –alpha –a –$^{4}_{2}$a –$^{4}_{2}$He
4. 146
5. –$^{3}_{2}\mathrm{He}$ –helium-3 –He-3 –$^{3}\mathrm{He}$
6. –$\frac{1}{4}$ –0.25 –25%
7. 15.813 y / 15.8 y
8. $^{60}_{28}$Ni/ ^{60}Ni/ nickel-60
9. —Gamma radiation has greater penetrating power. —Beta particles have weaker penetrating power.
10. —Nuclear radiation is harmful to all living cells. —Radioisotopes can cause gene mutations. —Treatments can cause stomach problems, such as nausea.

Chapter 3 – PHYSICAL BEHAVIOR OF MATTER/ENERGY

I. Law of Conservation of Energy:
   - Energy cannot be created or destroyed, only transferred from one type to another.
   - Examples of energy: heat, chemical, electrical, mechanical, nuclear, potential, kinetic.

II. Matter: Anything that has mass and takes up space



- Substances are ALWAYS homogeneous (pure)
- Mixtures can be either homogeneous or heterogeneous
- Elements **cannot** be broken down into anything simpler
- Diatomic molecules are elements (H_2, O_2, N_2, F_2, Br_2, Cl_2, I_2)
- Compounds **can** be broken down into something simpler
- Binary compounds are compounds composed of 2 elements.

Ways to separate mixtures:

1. **Magnet:** ex. Sulfur and iron – separate iron out with the magnet
2. **Distillation**: Separate by boiling – salt water. Water will evaporate and salt will be left
3. **Filtration**: use filter, liquid goes through and solids stay behind on the filter paper.

III. States of Matter:

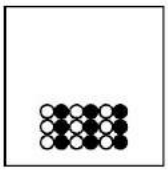
- Solids, liquids, gases (must know the general definition of each)
- MUST KNOW how to draw and interpret particle diagrams

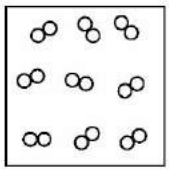
Example:

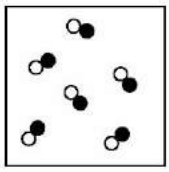
Given the key:

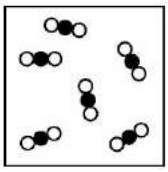
KEY:
○ = Atom of oxygen
● = Atom of carbon

Which particle diagram represents a sample containing the compound CO(g)?

A) 

B) 

C) 

D) 

- Physical vs. Chemical changes
- Energy is absorbed or given off in a chemical reaction. Reactions involving heat energy are classified as:

  EXOTHERMIC: energy is given off in a chemical reaction.
  A+B →C + Heat
  ENDOTHERMIC: energy is absorbed in a chemical reaction.
  A+ Heat →B+C

- **SOLID → LIQUID → GAS = <u>ENDOTHERMIC</u>**
  Solid → Liquid = melting
  Liquid → Gas = evaporation
  Solid → Gas = sublimation (CO_2 and I_2)

- **GAS → LIQUID → SOLID = <u>EXOTHERMIC</u>**
  Gas → Liquid = condensation
  Liquid → Solid = freezing (solidification)
  Gas → Solid = Deposition (CO_2 and I_2)

- **Entropy:** a measure of the disorder of a system. The more disorder, the more entropy. Gas molecules are more disorderly than liquid molecules and therefore entropy increases as you go from the liquid to the gas phase. As you go from a gas to a liquid phase, entropy decreases (becomes less disorderly).

MEASURING ENERGY:

- Energy is measured in joules. 1kilojoule = 1000 joules
- To solve heat energy problems (how much heat in joules is absorbed or released) the formula is: (Table T)

Heat	$q = mC\Delta T$ $q = mH_f$ $q = mH_v$	q = heat m = mass C = specific heat capacity ΔT = change in temperature	H_f = heat of fusion H_v = heat of vaporization

Table B
Physical Constants for Water

Heat of Fusion	334 J/g
Heat of Vaporization	2260 J/g
Specific Heat Capacity of $H_2O(\ell)$	4.18 J/g•K

- **If q is negative: exothermic reaction (release heat)**
- **If q is positive: endothermic reaction (absorb heat)**

Heat of vaporization plateau is longer than the heat of fusion plateau because it takes longer to gain or lose the required amount of heat during boiling and condensation than during melting or freezing.

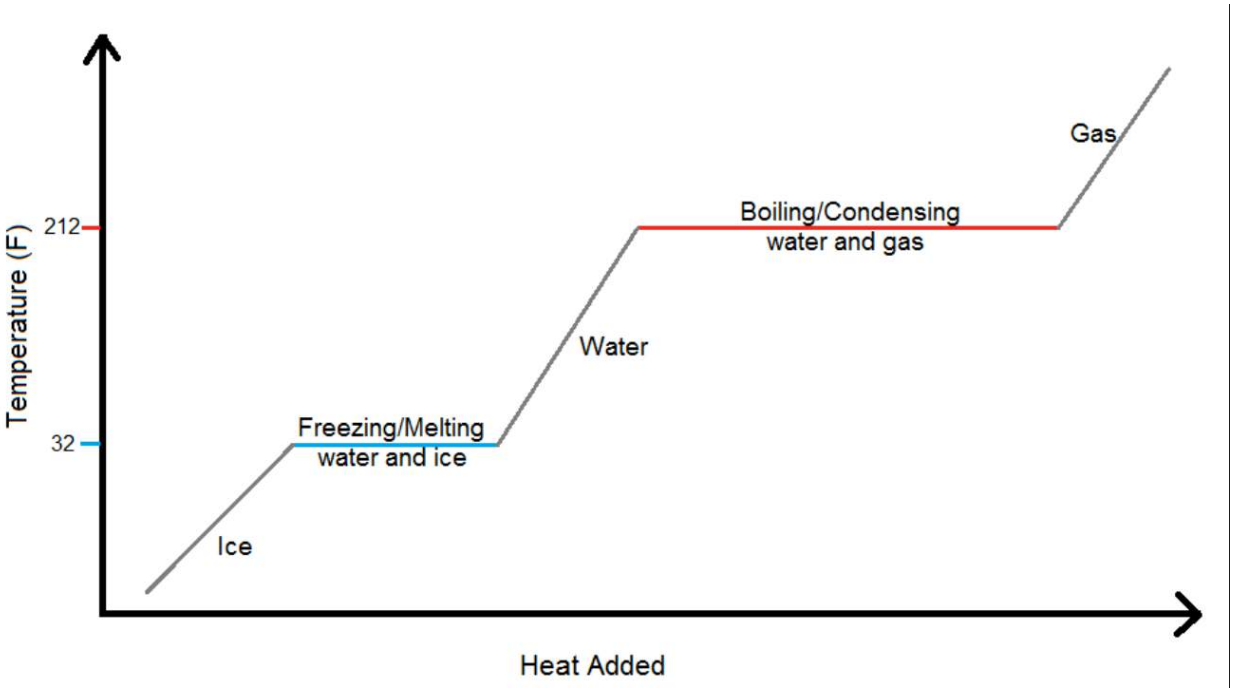
Temperature:

- a measure of the **average kinetic energy** of molecules. The higher the temperature, the more the kinetic energy; the lower the temperature, the lower the kinetic energy.

- **RULE**: heat flows from an area of higher temperature to an area of lower temperature until both temperatures are the same.

Temperature	$K = °C + 273$	K = kelvin °C = degree Celsius

The higher the temperature, the higher the kinetic energy, the higher the entropy
0 Kelvin = Absolute 0



When KE changes PE remains constant and vice versa. Know what happens to KE and PE at each interval.

Name: **Matter & Energy Review** **Constructed Response**

1. Base your answers to the following questions on the diagram of a molecule of nitrogen shown below:


represents one molecule of nitrogen

a Draw a particle model that shows at least six molecules of nitrogen gas.
b Draw a particle model that shows at least six molecules of liquid nitrogen.

Base your answers to questions **2** through **5** on the information below and on your knowledge of chemistry.

A student prepares two 141-gram mixtures, *A* and *B*. Each mixture consists of NH_4Cl, sand, and H_2O at 15°C. Both mixtures are thoroughly stirred and allowed to stand. The mass of each component used to make the mixtures is listed in the data table below.

Mass of the Components in Each Mixture

Component	Mixture A (g)	Mixture B (g)
NH_4Cl	40.	10.
sand	1	31
H_2O	100.	100.

2. Describe *one* property of sand that would enable the student to separate the sand from the other components in mixture *B*.

3. Determine the temperature at which all of the NH_4CL in mixture *A* dissolves to form a saturated solution.

4. Which type of mixture is mixture *B*?

5. State evidence from the table indicating that the proportion of the components in a mixture can vary.

Base your answers to questions **6** and **7** on the information below and on your knowledge of chemistry.

A few pieces of dry ice, $CO_2(s)$, at $-78^\circ C$ are placed in a flask that contains air at $21^\circ C$. The flask is sealed by placing an uninflated balloon over the mouth of the flask. As the balloon inflates, the dry ice disappears and no liquid is observed in the flask.

6. Write the name of the process that occurs as the dry ice undergoes a phase change in the flask.

7. State the direction of heat flow that occurs between the dry ice and the air in the flask.

Base your answers to questions **8** through **10** on the information below.

A student investigated heat transfer using a bottle of water. The student placed the bottle in a room at 20.5°C. The student measured the temperature of the water in the bottle at 7 a.m. and again at 3 p.m. The data from the investigation are shown in the table below.

Water Bottle Investigation Data

7 a.m.		3 p.m.	
Mass of Water (g)	**Temperature** (°C)	**Mass of Water** (g)	**Temperature** (°C)
800.	12.5	800.	20.5

8. Show a numerical setup for calculating the change in the thermal energy of the water in the bottle from 7 a.m. to 3 p.m.

9. State the direction of heat transfer between the surroundings and the water in the bottle from 7 a.m. to 3 p.m.

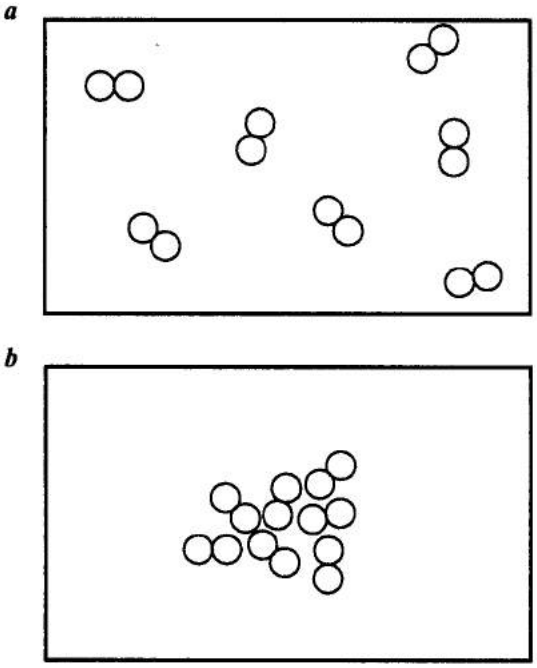
10. Compare the average kinetic energy of the water molecules in the bottle at 7 a.m. to the average kinetic energy of the water molecules in the bottle at 3 p.m.

Answer Key

Matter & Energy Review

1. (essay)
2. —Sand is insoluble in water. —Sand particles are too large to pass through filter paper. —Sand is more dense than $NH_4Cl(aq)$. —Sand remains a solid in the mixture.
3. $23°C$ to $26°C$
4. —heterogeneous —nonuniform mixture
5. —The ratio by mass of NH_4Cl to H_2O in mixture *A* is 40. g/100. g, and the ratio in mixture *B* is 10. g/100. g. —Both mixtures have the same total mass, but have different amounts of sand. —Mixture *B* has more sand. —The mixtures have different proportions of NH_4Cl.
6. –sublimation –subliming
7. –Heat flows from the air in the flask to the dry ice. –air to CO_2 –to dry ice –from air
8. q = (800. g)(4.18 J/g • ºC)(20.5ºC – 12.5ºC) (800)(4.18)(8)
9. –Heat was transferred from the surroundings to the water in the bottle. –The water absorbed energy from the surroundings.
10. –The average kinetic energy of the water molecules at 7 a.m. is less than the average kinetic energy of the water molecules at 3 p.m. –The average kinetic energy of the molecules is greater at 3 p.m.

1.



c Acceptable responses include, but are not limited to, these examples:
– The particles in nitrogen gas are farther away from each other than the partic les in the liquid nitrogen. *or*
– spacing of particles *or* – Gas particles have greater entropy (randomness) tha n the particles in the liquid.
d – Two dimensional models do not show geometric relationships. *or* – not 3-D *or* – Real particles are
three-dimensional. *or* – The model does not show momentary dipoles.

Chapter 5 – GASES/GAS LAWS

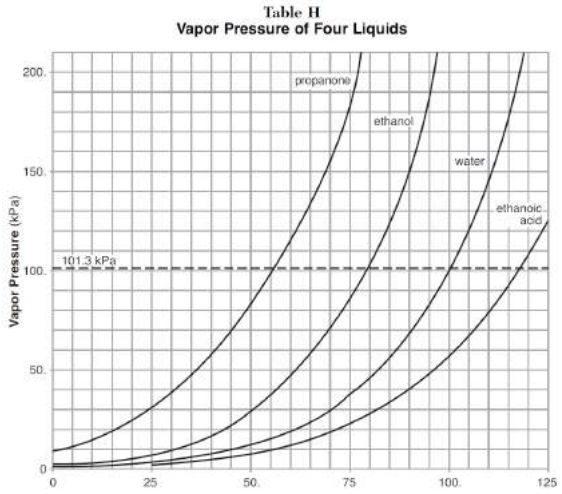
- **STP = Standard Temperature & Pressure (Table A)**

Table A
Standard Temperature and Pressure

Name	Value	Unit
Standard Pressure	101.3 kPa 1 atm	kilopascal atmosphere
Standard Temperature	273 K 0°C	kelvin degree Celsius

Boiling Point: Water boils when vapor pressure equals atmospheric pressure (Table H: when water is at 100ºC, atmospheric pressure is 101.3kPa). Water boils when vapor pressure = atmospheric pressure. Must know which has the strongest and weakest IMF.

Stronger IMF = Higher MP/BP



<u>COMBINED GAS LAW (TABLE T):</u> An equation that can be used to determine for pressure, volume, or temperature. Nothing remains constant. **Remember that by stating <u>STP</u>, numerical values are being given!!!!!**

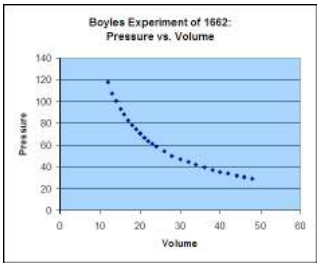
Combined Gas Law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	P = pressure V = volume T = temperature

Temperature MUST be in Kelvin

BOYLE'S LAW: At **constant temperature**, the volume of a gas is **inversely** proportional to pressure. This means that the more pressure you have on a gas, the smaller the volume of the gas. Doubling the pressure will half the volume.

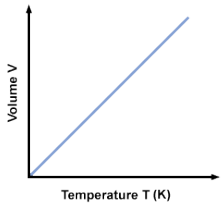
2 times as much pressure = ½ volume of gas
3 times as much pressure = 1/3 volume of gas

$$P_1V_1 = P_2V_2$$



CHARLES'S LAW: At constant pressure, volume is directly proportional to Kelvin (absolute) temperature. Temperature **must be in degrees Kelvin**, and therefore, conversion from Celsius to Kelvin may have to occur.

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



****BOTH BOYLE'S & CHARLES'S LAW CAN BE SOLVED BY USING THE COMBINED GAS LAW and just removing either pressure or temperature from the equation****

KINETIC MOLECULAR THEORY: A model that tells how gases should behave; also called **"ideal gas laws".** There are 4 key points:

1. A gas is composed of particles that are in continuous, random, straight line motion.
2. There is a transfer of energy between colliding particles; the total energy remains constant.
3. The volume of gas particles is negligible in comparison with the volume of space they are in. There is a lot of space between the particles.
4. Gas particles are considered as having no force of attraction for each other.

- "Real gases" deviate from the "ideal gas laws". **The deviations from the ideal gas laws mean how gases are different from the kinetic molecular theory.**
- **Point 3 Deviation**: The volume of gas particles is significant. Gas particles do have some volume.
- **Point 4 Deviation**: Gas particles do have a force of attraction.

IDEAL GASES: low pressure & high temperature (particles far apart from each other, and moving very fast)

REAL GASES: high pressure & low temperature (particles close together, and moving slowly by one another)

- Hydrogen and Helium (H_2 and He) are two REAL gases that act most like IDEAL gases.

- **Avogadro's Hypothesis or Law:** Equal volumes of all gases under the same conditions of temperature and pressure have equal numbers of molecules.

REMEMBER: 6.02×10^{23} particles

Base your answers to questions **1** and **2** on the information below and on your knowledge of chemistry.

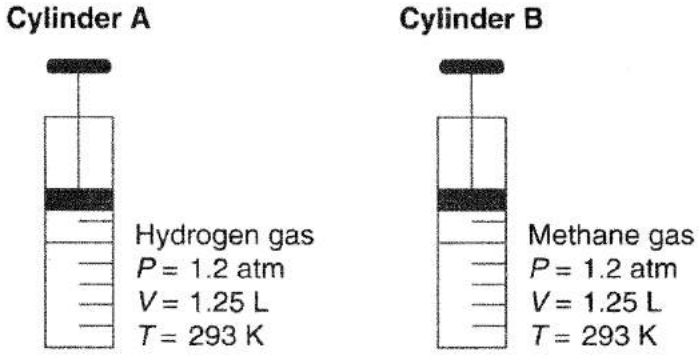
Ethane, C_2H_6, has a boiling point of $-89^\circ C$ at standard pressure. Ethanol, C_2H_5OH, has a much higher boiling point than ethane at standard pressure. At STP, ethane is a gas and ethanol is a liquid.

1. Compare the intermolecular forces of the two substances at STP.

2. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Based on Table *H*, what is the boiling point of ethanol at standard pressure?

Base your answers to questions **3** through **5** on the information below and on your knowledge of chemistry.

Cylinder *A* has a movable piston and contains hydrogen gas. An identical cylinder, *B*, contains methane gas. The diagram below represents these cylinders and the conditions of pressure, volume, and temperature of the gas in each cylinder.



3. Show a numerical setup for calculating the volume of the gas in cylinder *B* at STP.

4. State a change in temperature and a change in pressure that will cause the gas in cylinder *A* to behave more like an ideal gas.

5. Compare the total number of gas molecules in cylinder *A* to the total number of gas molecules in cylinder *B*.

Base your answers to questions **6** through **8** on the information below.

A sample of helium gas is in a closed system with a movable piston. The volume of the gas sample is changed when both the temperature and the pressure of the sample are increased. The table below shows the initial temperature, pressure, and volume of the gas sample, as well as the final temperature and pressure of the sample.

Helium Gas in a Closed System

Condition	Temperature (K)	Pressure (atm)	Volume (mL)
Initial	200.	2.0	500.
final	300.	7.0	?

6. Compare the total number of gas particles in the sample under the initial conditions to the total number of gas particles in the sample under the final conditions.

7. Convert the final temperature of the helium gas sample to degrees Celsius.

8. In the space below show a correct numerical setup for calculating the final volume of the helium gas sample.

Base your answers to questions **9** and **10** on the information below.

Air bags are an important safety feature in modern automobiles. An air bag is inflated in milliseconds by the explosive decomposition of $NaN_3(s)$. The decomposition reaction produces $N_2(g)$, as well as Na(s), according to the unbalanced equation below.

$$NaN_3(s) \rightarrow Na(s) + N_2(g)$$

9. When the air bag inflates, the nitrogen gas is at a pressure of 1.30 atmospheres, a temperature of 301 K, and has a volume of 40.0 liters. Calculate the volume of the nitrogen gas at STP. Your response must include *both* a correct numerical setup and the calculated volume

10. Balance the equation for the decomposition of NaN_3, using the smallest whole-number coefficients.

Answer Key
Gases Review

1. –Ethane has weaker intermolecular forces (IMF) than ethanol.
–Ethanol has hydrogen bonding.
–Van der Waals forces are weaker in C_2H_6.

2. any value from 78°C to 80.°C

3. $\frac{(1.2\,\text{atm})(1.25\,\text{L})}{293\,\text{K}} = \frac{(1.0\text{atm})(V_2)}{273\,\text{K}}$

$\frac{(273)(1.2)(1.25)}{293}$

4. Temperature: above 293 K Pressure : below 1.2 atm
Temperature: higher
Pressure: lower

5. –The number of gas molecules in cylinder *A* is the same as the number of gas molecules in cylinder *B*.

6. —The total number of gas particles is the same under the initial and final conditions.
—The total number of particles before and after is the same.

7. 27°C

8. $\frac{(2.0\,\text{atm}(500.\,\text{mL})}{200.\,\text{K}} = \frac{7.0\,\text{atm}V_2}{300.\,\text{K}}$

$\frac{(2)(500)(300)}{200(7)}$

9. 47.2 L
V_2 = (273 K)(1.30 atm)(40.0 L)
(301 K)(1.00 atm)
(273)(1.30)(40.0)
(301)(1.00)

10. $2\,NaN_3(s) \rightarrow 2\,Na(s) + 3\,N_2(g)$

Chapter 6 – PERIODIC TABLE

I. Location and arrangement of elements on the PT:
- **Periods:** Horizontal rows on the PT (elements have the same # of PEL's)
- **Groups:** Vertical columns (elements have the same # of valence electrons)
- **Periodic Law:** The properties of the elements are a periodic function of their atomic number

Elements in the same groups have more similar chemical properties than elements in the same period because they have the same number of valence electrons.

- Representative groups follow the trends – Groups 1, 2, and 13-18.
- Transition Metals do not follow the same type of trends – Groups 3-12
- **Group 1 – Alkali Metals**
- **Group 2 – Alkaline Earth Metals**
- **Groups 3-12 – Transition Metals:** transition metals that have more than one oxidation number form ions that are **colored** in solution.
- **Group 17 – Halogen Group**
- **Group 18 – Noble Gases**

II. Classifying Elements:
- Left of the zig-zag is a METAL
- Right of the zig-zag is a NON-METAL
- On the zig-zag is a METALLOID (exceptions Al and Po)
- Most elements are solid with the following exceptions
  - **i. Gases**: 11, 5 are reactive and 6 are non-reactive (reactive gases are H_2, O_2, N_2, F_2, Cl_2; the non-reactive are the Noble Gases)
  - **ii. Liquids**: Bromine is the only non-metal liquid; Mercury is the only metal liquid.

- Francium is the most reactive metal
- Fluorine is the most reactive non-metal

III. Allotropes: Different … Different … Different!!
- Two or more forms of the same element that differ in their molecules (O_2 and O_3) or crystalline structure (forms of carbon).
- Oxygen has 2 allotropes: O_2 and O_3
- Carbon has many different allotropes which differ in arrangement of atoms
  - **i.** Diamond: every carbon bonded to 4 other carbons = very hard
  - **ii.** Graphite: arranged in sheets or layers = "lead" pencils.
  - **iii.** Coal: no definite pattern.
  - **iv.** Buckminsterfullerence: rings of 5 and 6 carbon atoms, looks like the outside of a soccer ball (60-70 carbons)

IV. Properties of Metals:
- Luster, malleable, are good conductors of heat and electricity, are very soluble (metals in group 1 are more soluble than metals in groups 2 and so on).
- Transition metals are much harder than metals in group 1 or 2.
- Elements in group 1 are more reactive than elements in group 2.

V. Properties of Non-metals:
- Brittle, lack luster, poor conductors of heat and electricity
- Non-metals are usually gases, molecular solids, or network solids.

PERIODIC TRENDS

Table S
Properties of Selected Elements

	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)

IONIZATION ENERGY: The amount of energy needed to remove an electron. The smaller the amount of ionization energy, the easier it is to lose an electron.

Ionization energy decreases as you go down a group
- More PEL's; farther away, less attraction so requires less energy to lose electrons.

Ionization energy increases as you go across a period
- More protons and more valence electrons as you go across a period and therefore more attraction so requires more energy to lose electrons.

ELECTONEGATIVITY: The attraction for electrons; the larger the electronegativity, the more the atom attracts electrons.

Electronegativity decreases as you go down any group
- More PEL's; farther away, less attraction so requires less energy to lose electrons.

Electronegativity increases as you go across a period
- More valence electrons as you go across a period and therefore more attraction so requires more energy to lose electrons.

ATOMIC RADIUS: Is the distance from the nucleus to the outer valence electrons.

Atomic radius increases as you go down a group

- As you go down a group, each element has an extra PEL (shell) and therefore the atomic radius increases

Atomic radius decreases as you go across a period.

- More attraction as you go across a period and so the distance from the nucleus to the outer PEL decreases slightly

IONIC RADIUS: The distance of the nucleus to the outermost valence electron in an ion. You will be asked to compare an atom to its own ion.

GAIN = GREATER
LOSE = LESS

Name: **Periodic Table Review** **Constructed Response**

Base your answers to questions **1** through **5** on the information below and on your knowledge of chemistry.

The Lewis electron-dot diagrams for three substances are shown below.

Diagram 1

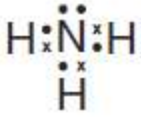


Diagram 2

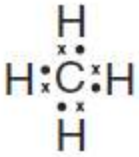


Diagram 3

1. Identify the noble gas that has atoms with the same electron configuration as the positive ion represented in diagram 1, when both the atoms and the ion are in the ground state.

2. Draw a Lewis electron-dot diagram for a molecule of $\mathrm{Br_2}$.

3. Explain, in terms of distribution of charge, why a molecule of the substance represented in diagram 3 is nonpolar.

4. Determine the total number of electrons in the bonds between the nitrogen atom and the three hydrogen atoms represented in diagram 2.

5. Describe, in terms of valence electrons, how the chemical bonds form in the substance represented in diagram 1.

---

6. Explain, in terms of element classification, why $\mathrm{K_2O}$ is an ionic compound.

Base your answers to questions **7** through **9** on the information below and on your knowledge of chemistry.

There are six elements in Group 14 on the Periodic Table. One of these elements has the symbol Uuq, which is a temporary, systematic symbol. This element is now known as flerovium.

7. State the expected number of valence electrons in an atom of the element flerovium in the ground state.

8. Explain, in terms of electron shells, why each successive element in Group 14 has a larger atomic radius, as the elements are considered in order of increasing atomic number.

9. Identify an element in Group 14 that is classified as a metalloid.

10. Explain, in terms of electrons, why the radius of a potassium atom is larger than the radius of a potassium ion in the ground state.

11. Explain, in terms of atomic structure, why Group 18 elements on the Periodic Table rarely form compounds.

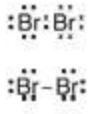
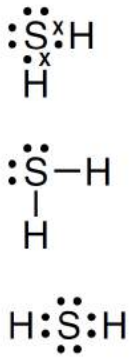
12. Base your answer to the following question on the information below and on your knowledge of chemistry.

    Silver-plated utensils were popular before stainless steel became widely used to make eating utensils. Silver tarnishes when it comes in contact with hydrogen sulfide, H_2S, which is found in the air and in some foods. However, stainless steel does not tarnish when it comes in contact with hydrogen sulfide.

    Draw a Lewis electron-dot diagram for the compound that tarnishes silver.

Answer Key

Periodic Table Review 1

1. argon and Ar
2. 
3. —Charge is symmetrically distributed. —The molecule has uniform charge distribution. —The centers of positive charge and negative charge coincide.
4. 6 *or* six
5. —Valence electrons are lost by potassium and gained by bromine. —The ions form as a result of a transfer of electrons between the atoms.
6. —A metal reacts with a nonmetal to produce an ionic compound. —Potassium is a metal and oxygen is a nonmetal.
7. –4 –four –4e$^-$ –four valence electrons
8. –The atomic radius of these elements increases down the group because each successive element has one more electron shell. –The number of shells per atom increases.
9. –Si –germanium –element 32
10. –A potassium atom has four electron shells and a potassium ion has three electron shells. –A potassium atom has one more electron shell than a potassium ion. –A K^+ ion has one fewer electron than a K atom.
11. Group 18 elements rarely form compounds because their atoms have stable electron configurations. –Their valence shells are completely filled. –All the elements have maximum numbers of valence electrons. –Atoms of Group 18 have a stable octet except He, which is stable with two electrons.
12. 

Chapter 7 – NOMENCLATURE/BALANCING/REACTIONS

REACTANTS → PRODUCTS

- **BALANCING EQUATIONS**: THE ELEMENTS ON ONE SIDE OF THE EQUATION MUST EQUAL THE ELEMENTS ON THE OTHER SIDE OF THE EQUATION.
- **Molecular formulas:** indicates the total number of atoms of each element needed to form a molecule. Ex: C_2H_6 (2 carbons and 6 hydrogens).
- **Empirical formula:** is the simplest ratio in which atoms combine to form a compound. If the formula is C_2H_6 divide the C and the H by the largest number that all of the elements can be divided by, in this case 2, and the empirical formula would be CH_3. **Try $C_6H_{12}O_6$.**

NAMING TYPE 1 COMPOUNDS: IONIC COMPOUNDS

- May occur between the following:
  - Metal (+) & Non-metal (-)
  - Metal (+) & Polyatomic ion (-)
  - Polyatomic ion (+) & Non-metal (-)
  - Polyatomic ion (+) & Polyatomic ion (-)

- NEVER change the name of the polyatomic ions
- The metal keeps its same name
- The non-metal you drop the last few letters and add –ide.
- **Criss-cross method:** if the oxidation numbers of the elements don't equal zero, then you must cross the number of the oxidation number, not the charge, in order for the charge to equal 0.
- If you can reduce the subscripts then reduce.
- **EX**: $CaCl_2$ = calcium chloride, NH_4Cl = ammonium chloride, $Li_2(SO_4)$ = lithium sulfate, $(NH_4)_2(SO_4)$ = ammonium sulfate

NAMING TYPE 2 COMPOUNDS: (Stock System) IONIC COMPOUNDS

- USE THIS ONLY WHEN THE METAL HAS MORE THAT 1 OXIDATION #
- THE ROMAN NUMERAL INDICATES WHICH OXIDATION # TO USE; Example: I = +1, III = +3, VI = +6
- **EX:** $Fe_2(O)_3$ **=** iron (III) oxide, Au_2O = gold (I) oxide.

NAMING TYPE 3 COMPOUNDS: COVALENT COMPOUNDS

- Use this when you name covalent/molecular compounds (2 NM's)
- Use the prefixes to name the compounds
- First non-metal keeps its same name and only gets a prefix if there are more than 1

- Second non-metal **always** gets a prefix, and the ending is –ide.

Mono - 1	Hexa - 6
Di - 2	Hepta - 7
Tri - 3	Octa - 8
Tetra - 4	Nona - 9
Penta - 5	Deca - 10

CLASSIFYING CHEMICAL REACTIONS:

- When two or more chemicals are brought together, a chemical change (reaction) is likely to take place. Some evidence that a chemical reaction has occurred include the following:
    1. A color change occurs
    2. A solid forms (precipitate) – INSOLUBLE (Table F)
    3. A gas is released (bubbles form)
    4. Heat and/or a flame are produced (exothermic)
    5. Heat is absorbed (endothermic)

- **SYNTHESIS REACTIONS:** Two or more elements or simpler compounds unite to form a compound (BARF)

$$\mathbf{2Na + Cl_2 \rightarrow 2NaCl}$$

- **DECOMPOSITION REACTIONS:** A compound is broken down into two or more simpler compounds (BARF)

$$\mathbf{2NaCl \rightarrow 2\ Na + Cl_2}$$

- **SINGLE REPLACEMENT REACTIONS:** A free element (an element alone like Fe) replaces an element that is part of a compound. Free element must be more reactive than the element it's replacing (Table J)

$$\mathbf{Fe + CuSO_4 \rightarrow FeSO_4 + Cu} \qquad \mathbf{F_2 + CuI_2 \rightarrow CuF_2 + I_2}$$

- **DOUBLE REPLACEMENT REACTIONS:** Two elements replace each other or switch partners. In the example given, the Na and Ag replace each other or switch partners. Two new compounds, $NaNO_3$ (sodium nitrite) and AgCl (silver chloride) are formed.

$$\mathbf{NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl}$$

- **COMBUSTION REACTIONS:** the reaction of a carbon-based compound with oxygen; the products are carbon dioxide and water (first 6 reactions on Table I)

$$\mathbf{C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O}$$

Chapter 8 – ORGANIC CHEMISTRY

- **Organic Chemistry:** study of carbon and carbon compounds; there are a lot of carbon compounds. The C atoms bond together to form chains or rings.
- **Common Characteristics:**

1. Generally non-polar (won't dissolve in water)
2. Soluble in non-polar solvents
3. Non-electrolytes.
4. Low melting points
5. Reactions generally slower than inorganic compounds
6. Reactions have higher activations energies (therefore slower reactions)

- Carbon has 4 valence electrons and therefore can form 4 covalent bonds around them (remember single, double or triple)
- **Hydrocarbons:** contain carbon and hydrogen atoms.

Table P &Q: Table P indicates the prefix used based on the # of carbons.

ALKANE: (-ane)

- Hydrocarbons are made up of only **single bonds**, and are therefore considered **saturated.**

ALKENE: (-ene)

- Alkenes have one **double bond** and classified as **unsaturated.**

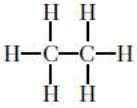
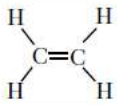
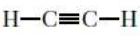
ALKYNE : (-yne)

- Alkynes have one **triple bond** and are considered **unsaturated.**

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 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

- **Condensed Formulas:** Taking an organic compound, and shortening the chemical formula. Same compound, just written differently.

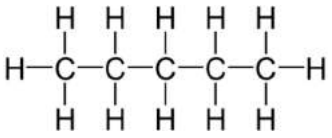
Alkyl Group: (-yl)

- Also hydrocarbons, known as side chains.
- Have one less hydrogen then a corresponding alkane.

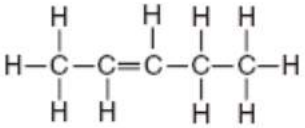
NUMBER OF CARBON ATOMS	STEM NAME	SIDE CHAIN (ALKYL GROUP) NAME	
1	meth	methyl	CH_3
2	eth	ethyl	C_2H_5
3	prop	propyl	C_3H_7
4	but	butyl	C_4H_9
5	pent	pentyl	C_5H_{11}
6	hex	hexyl	C_6H_{13}
7	hep	heptyl	C_7H_{15}
8	oct	octyl	C_8H_{17}
9	non	nonyl	C_9H_{19}
10	dec	decyl	$C_{10}H_{21}$

SUMMARY ON HOW TO DRAW CARBON COMPOUNDS:

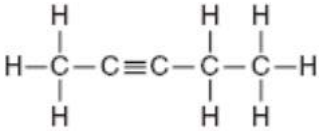
1. Look at the prefix (table P), which will tell you the number of carbons.
2. Put the bonds between the carbons (-ane = single bonds; -ene = double bond; and –yne means a triple bond)
3. The number before –ene and –yne tells you where the double or triple bond is after that carbon atom, Ex: 2-pentene (db after 2nd carbon), or 1-butyne (tb after 1st carbon).
4. When needed, if you have an alkyl group like methyl (CH_3), the number before tells you which carbon atom to put it on; Ex: 2-methyl (put the methyl on the 2nd carbon).
5. Put all of the appropriate hydrogen atoms around the carbon atoms if they belong there.



Pentane



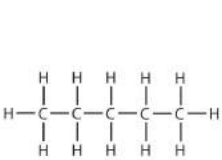
2-Pentene



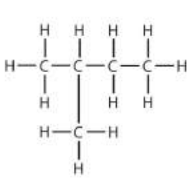
3-Pentyne

ISOMERS: compounds that have the same molecular formula but different structural formula.

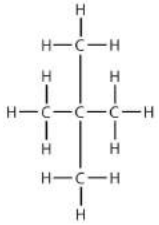
- The rule to naming these structures is to count the number of carbon atoms in the longest unbroken chain. You want to use the lowest possible number, so this means you might have to count from the left to right or from right to left.



Pentane



2-methyl butane



2,2 dimethyl propane

OTHER ORGANIC COMPOUNDS & FUNCTIONAL GROUPS (Table R):

Alcohols: (OH)

- The OH in an alcohol is a hydroxyl, and the OH^- in a base is a hydroxide ion; **ALCOHOLS ARE NOT BASES.**
- To name an alcohol, drop the –e from the corresponding alkane and all **–ol.**
- OH can be placed on any carbon and therefore a number is required to indicate on which carbon it's on when there are 3 or more carbons in a chain.
- Diol (2 OH are present); triol (3 OH are present)

Ether:

- You name this compound by using the word **–ether** at the end and use the name of the alkyl groups attached at both ends.
- Alcohols and Ethers make isomers of each other when there are the same number of carbons atoms.

Aldehyde: (CHO)

- Named by dropping the final –e in an alkane and adding **–al.**
- Found at the end, no number required.

Ketone:

- Ketones are named by dropping the final –e from the corresponding alkane and adding **–one.**
- The double bonded oxygen MUST be on an inside carbon, and therefore, requires a number to indicate on which carbon it's attached to on the chain.
- Aldehydes and Ketones make isomers of each other when there are the same number of carbon atoms.

Organic Acids: (COOH)

- Named by dropping the final -e from the corresponding alkane and adding **– oic acid.**
- Functional group at the end, and therefore, no number is required.

Halides: (F, Cl, Br, I)

- Name it by using a number to state which atom the halogen is being attached to, then use the prefix for that halogen (fluoro-, chloro-, bromo-, iodo-) and end with the appropriate alkane.
- There can be multiple halides so it's necessary to use a number to indicate which carbon(s) they are on. Use prefixes like di, tri, etc … if there are more than one of the same kind of halogen.

Amine:

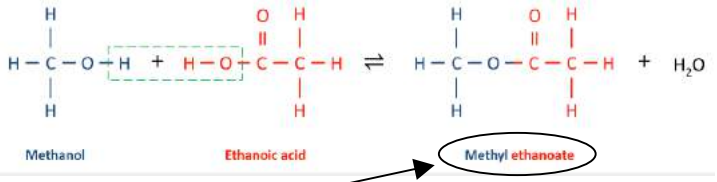
- Named by dropping the final –e in the hydrocarbon and adding **–amine.**
- Can be placed on any carbon, so a number is required.

Amide:

- Named by dropping the final –e in the hydrocarbon and adding **–amide.**
- Functional group at the end, therefore no number is required.

Ester: (COO)

- Esterfication – Alcohol + Acid → Ester + Water



- How to name:
  1. Name the part that came from the alcohol first using an alkyl name depending on how many carbon atoms there were.
  2. Name the part that came from the acid second by naming it like an alkane depending on the number of carbons, dropping the –ane, and adding the ending –oate.

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	$-\overset{\vert}{N}-$	$R-\overset{R'}{\overset{\vert}{N}}-R''$	$CH_3CH_2CH_2NH_2$ 1-propanamine
amide	$-C(=O)-NH$	$R-C(=O)-\overset{R'}{N}H$	$CH_3CH_2C(=O)-NH_2$ propanamide

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

Effects of Functional Groups on Boiling Point

1.Alcohols and Acids: Have the highest BP because they are held together with hydrogen bonding.

2.All other functional groups: Next highest BP.

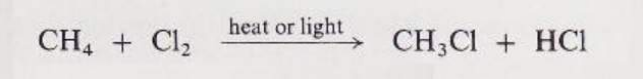
3.Hydrocarbons: Weakest BP.

Examples: Methanol → Methanal → Methane (highest to lowest BP)

ORGANIC REACTIONS:

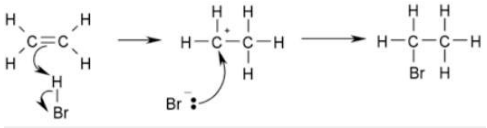
Substitution: Occurs in **ALKANES** only.

- 2 reactants and 2 products.



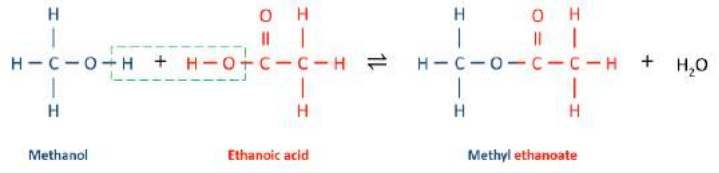
Addition: Happen only to **ALKENES & ALKYNES.** Breaking a double to a single or a triple to a double.

- 2 reactants and 1 product.



Fermentation: Glucose is broken down into **ethanol and carbon dioxide** (Anaerobic Respiration).

Esterfication: Alcohol + Acid → Ester + Water



Saponification: Reverse of esterification. This process produces SOAP.

Combustion: First 6 reactions on Table I.

- Organic Compound + O_2 → CO_2 + H_2O

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
$2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(\ell)$	−10943
$2CH_3OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(\ell)$	−1452
$C_2H_5OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(\ell)$	−1367

Polymerization: involves smaller molecules joining together to form one big molecule. A **polymer** is a large molecule made up of multiple monomers.

- **condensation:** polymerization by dehydration synthesis (removing water) to form a polymer. Ex: nylons, polyester. Naturally occurring polymers include starches and protein.
- **addition:** polymerization joining together by breaking a double or triple bond.

Name _____ Chap. 8 Organic Review Constructed Response

Base your answers to questions **1** and **2** on the information below and on your knowledge of chemistry.

Natural gas and coal are two fuels burned to produce energy. Natural gas consists of approximately 80% methane, 10% ethane, 4% propane, 2% butane, and other components.

The burning of coal usually produces sulfur dioxide, $SO_2(g)$ and sulfur trioxide, $SO_3(g)$, which are major air pollutants. Both $SO_2(g)$ and $SO_3(g)$ react with water in the air to form acids.

1. Draw a structural formula for the hydrocarbon that is approximately 2% of natural gas.

2. Write the general formula for the homologous series that includes the components of the natural gas listed in this passage.

---

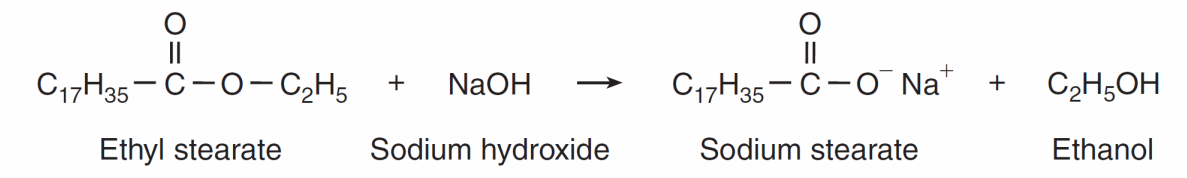
3. Base your answer to the following question on the information below and on your knowledge of chemistry.

Ethane, C_2H_6, has a boiling point of $-89^\circ C$ at standard pressure. Ethanol, C_2H_5OH, has a much higher boiling point than ethane at standard pressure. At STP, ethane is a gas and ethanol is a liquid.

Identify the class of organic compounds to which ethanol belongs.

Base your answers to questions **4** and **5** on information below.

One type of soap is produced when ethyl stearate and sodium hydroxide react. The soap produced by this reaction is called sodium stearate. The other product of the reaction is ethanol. This reaction is represented by the balanced equation below.



4. To which class of organic compounds does ethyl stearate belong?

5. Identify the type of organic reaction used to make soap.

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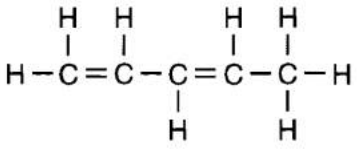
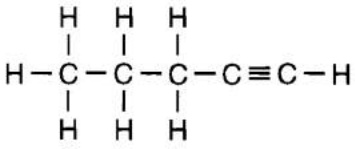
6. Two hydrocarbons that are isomers of each other are represented by the structural formulas and molecular formulas below.

Hydrocarbon 1



C_5H_8

Hydrocarbon 2

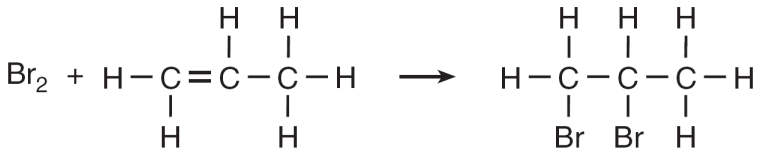


C_5H_8

Explain, in terms of structural formulas and molecular formulas, why these hydrocarbons are isomers of each other.

Base your answers to questions **7** and **8** on the information below.

A reaction between bromine and a hydrocarbon is represented by the balanced equation below.



7. Write the name of the homologous series to which the hydrocarbon belongs.

8. Identify the type of organic reaction.

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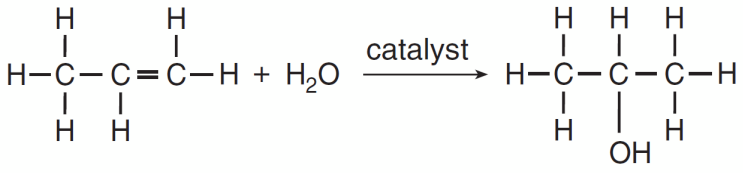
9. Base your answer to the following question on the information below.

In one industrial organic reaction, C_3H_6 reacts with water in the presence of a catalyst. This reaction is represented by the balanced equation below.



Explain, in terms of bonding, why C_3H_6 is classified as an unsaturated hydrocarbon.

10. Base your answer to the following question 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.

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

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

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

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

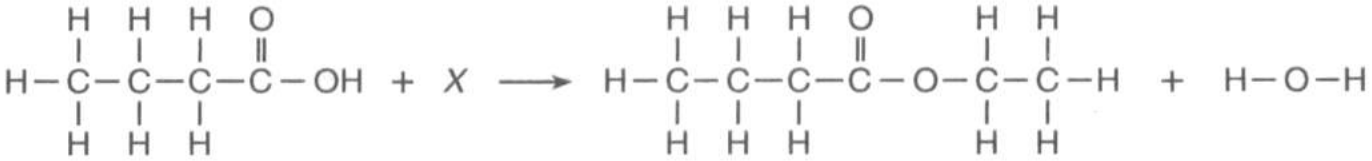
Gasoline is assigned a number called an octane rating. Gasoline with an octane rating of 87 performs the same as a mixture that consists of 87% isooctane and 13% heptane.

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

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

Base your answers to questions **12** and **13** on the following information.

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



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

13. Identify the type of organic reaction represented by the equation.

Base your answers to questions **14** and **15** on the information below.

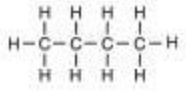
Given the balanced equation for an organic reaction between butane and chlorine that takes place at 300.ºC and 101.3 kilopascals:

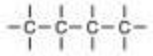
$$C_4H_{10} + Cl_2 \rightarrow C_4H_9Cl + HCl$$

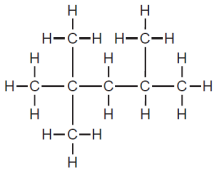
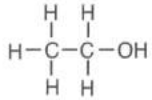
14. Draw a structural formula for the organic product.

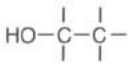
15. Identify the type of organic reaction shown.

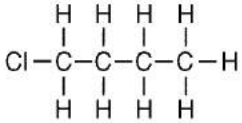
Answer Key
Organic Chemistry

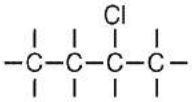
1. 



2. C_nH_{2n+2}
3. –alcohol –alcohols –primary alcohol –monohydroxy alcohols
4. ester *or* esters
5. saponification
6. The molecular formulas of the two hydrocarbons are the same, but the structural formulas are different.
7. – alkene *or alkenes.*
8. – addition – halogenation – bromination
9. Acceptable responses include, but are not limited to: The C_3H_6 is unsaturated because each molecule has a double covalent bond between two of its carbon atoms. There is a carbon-carbon double bond in each molecule
10. Example: combustion
11. 
12. 



13. *Examples:* – esterification – dehydration synthesis
14. 



15. *Examples*: – substitution – chlorination – halogenation

Chapter 9 - CHEMICAL BONDING

REMEMBER:

- **Electronegativity:** an atom's attraction for electrons in a bond.
- Metals tend to have lower EN and Non-metals have higher EN.

IONIC BONDS: formed when a metal transfers one or more electrons to a non-metal to form ions. Opposite charges attract. Ionic bonds are **ALWAYS POLAR BONDS.**

- May occur between the following:
  - Metal (+) & Non-metal (-)
  - Metal (+) & Polyatomic ion (-)
  - Polyatomic ion (+) & Non-metal (-)
  - Polyatomic ion (+) & Polyatomic ion (-)

- If electronegativity difference is >1.7, then it is an IONIC BOND.

PROPERTIES OF IONIC BONDS:

1. Hard
2. Good conductors of electricity – ONLY IN LIQUID OR AQUEOUS PHASE.
3. High melting and boiling points
4. Dissolve in polar substances: "Likes Dissolve in Likes"

Ionic:

*Metals & PI – charge and possible subscript

*NM 

Ionic compounds that have at least one Polyatomic ion in it is said to have BOTH Ionic and Covalent Bonds.

Na_3N

$Zn(OH)_2$

$(NH_4)_2O$

COVALENT BONDS: Formed when two atoms (both non-metals) share electrons.

- Also called **molecular bonds.**

Covalent Bonds (Molecular):

2 Non-Metal

Sharing of electrons

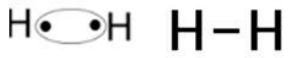
single, double, or triple bond

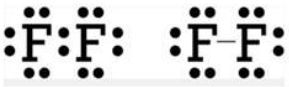
POLAR BOND	NON-POLAR
(different EN)	(same EN)
H_2O, NH_3, HCl, CH_4, CO_2	7 DIATOMIC
POLAR MOLECULE	NP MOLECULE
(ASYMMETRICAL)	(SYMMETRICAL)
H_2O, NH_3, HCl	CH_4, CO_2
δ^+ or δ^-	NO δ^+ or δ^-
	NP MOLECULE
	(SYMMETRICAL)
	NO δ^+ or δ^-

Higher EN (-), Lower EN (+)

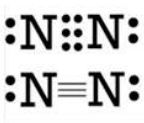
PROPERTIES OF COVALENT BONDS:

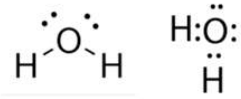
1. Soft
2. Poor conductors of heat/electricity
3. Low melting and boiling points

H_2 

F_2, Cl_2, Br_2, I_2 

O_2 

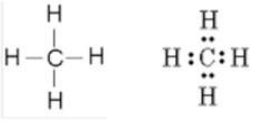
N_2 

H_2O (H_2S) 

NH_3 (PH_3, NF_3, NCl_3, NBr_3, NI_3, PF_3, PCl_3, PBr_3, PI_3)

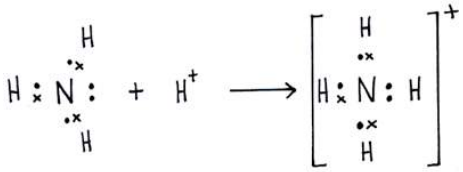
HCl (HF, HBr, HI)

CH_4 (CF_4, CCl_4, CBr_4, CI_4)



CO_2 (CS_2)

COORDINATE COVALENT: Formed when one atom donates a pair of electrons. NH_4^+ and H_3O^+.



NETWORK SOLIDS: Solids that have covalent bonds between atoms linked in one big network or one big **macromolecule** with no discrete particles. This gives them some different properties from most covalent compounds
Examples: Diamond (C), silicon carbide (SiC), and silicon dioxide (SiO_2)

Properties of network solid substances that have covalent bonds:

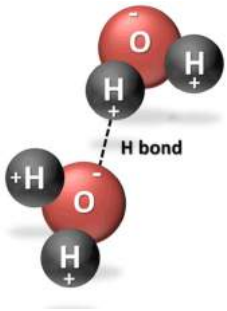
1. Hard
2. Poor conductors of heat and electricity
3. High melting points

METALLIC BOND: "Mobile Electrons". Why metals are good conductors in the solid phase.

INTERMOLECULAR FORCES: forces of attraction between molecules

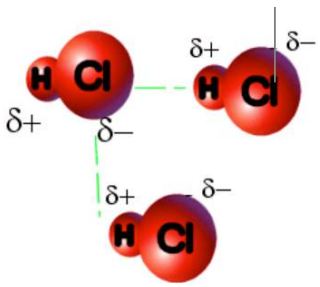
HYDROGEN BONDING:

- Strong intermolecular force that connects one water molecule with another water molecule (also other examples)
- Hydrogen bonds are also formed when hydrogen is covalently bonded to a small, highly electronegative atom such as F, O, or N. Examples: H_2O, NH_3, and HF.



DIPOLE ATTRACTION:

- Same concept as Hydrogen Bond but for other polar molecules (asymmetrical with partially (+) and (-) ends. Ex: H_2S, PCl_3, HCl

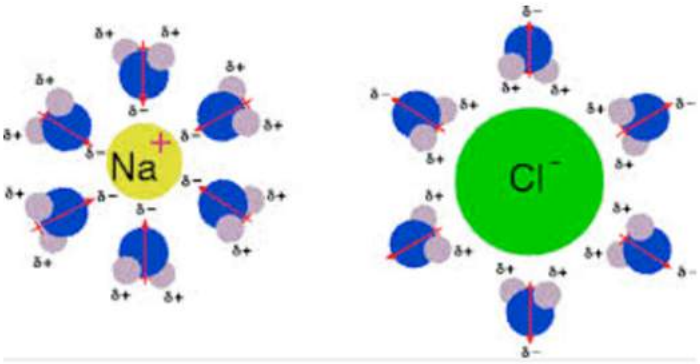


DISPERSION FORCES: (AKA: van der Waals forces)

- Weak intermolecular forces between SYMETRICAL NON-POLAR MOLECULES.
- No δ+ or δ-
- Increases with GFM. As you go down group 17 these forces increase and BP increase. This is why F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid.

MOLECULE-ION ATTRACTION:

- Attraction between the ions of an ionic compound such as salt, and molecules of water (or other polar liquids).
- "Likes Dissolve in Likes"



Greater EN Δ = More Polar = Stronger IMF = Higher MP/BP

Name _____ Chap. 9 Chemical Bonding Constructed Response

1. The volume of 1.00 mole of hydrogen bromide at STP is 22.4 liters. The gram-formula mass of hydrogen bromide is 80.9 grams per mole. What is the density of hydrogen bromide at STP?

2. Identify the type of bonding in solid potassium.

Base your answers to questions **3** and **4** 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.

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

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

4. Write the chemical formula for *one compound in the equation that contains both ionic bonds and covalent bonds.*

Base your answers to questions **5** and **6** on the information below.

Ozone, $O_3(g)$, is produced from oxygen, $O_2(g)$ by electrical discharge during thunderstorms. The unbalanced equation below represents the reaction that forms ozone.

$$O_2(g) \xrightarrow{\text{electricity}} O_3(g)$$

5. Explain, in terms of electron configuration, why an oxygen molecule is more stable than an oxygen atom.

6. Identify the type of bonding between the atoms in an oxygen molecule.

7. Base your answer to the following question on the balanced equation below.

$2Na(s) + Cl_2 \rightarrow 2NaCl(s)$

Draw a Lewis electron-dot diagram for a molecule of chlorine, Cl_2.

Base your answers to questions **8** through **11** on the table below.

Physical Properties of Four Gasses

Name of Gas	hydrogen	hydrogen chloride	hydrogen bromide	hydrogen iodide
Molecular Structure	H-H	H-Cl	H-Br	H-I
Boiling Point (K) at 1 Atm	20.	188	207	237
Density (g/L) at STP	0.0899	1.64	?	5.66

8. Explain, in terms of molecular polarity, why hydrogen chloride is more soluble than hydrogen in water under the same conditions of temperature and pressure.

9. Explain, in terms of electronegativity difference, why the bond in H–Cl is more polar than the bond in H–I.

10. The density of hydrogen at STP is 0.0899 gram per liter. Express this density to *two significant figures.*

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

Physical Properties of CF_4 and NH_3 at Standard Pressure

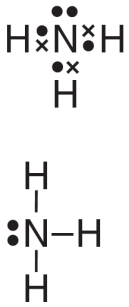
Compound	Melting Point (°C)	Boiling Point (°C)	Solubility in Water at 20.0°C
CF_4	−183.6	−127.8	insoluble
NH_3	−77.7	−33.3	soluble

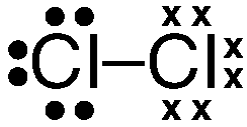
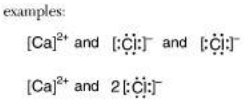
State evidence that indicates NH_3 has stronger intermolecular forces than CF_4.

12 Explain, in terms of molecular structure or distribution of charge, why a molecule of methane is nonpolar.

13. Draw the electron-dot (Lewis) structure of calcium chloride.

Answer Key
Bonding Review

1. 3.61 g/L
2. –metallic bonding –metallic
3. 

4. $NaHCO_3$ *or* NH_4Cl.
5. – Both atoms in an O2 molecule have achieved a noble gas electron configuration. – An oxygen atom does not have a stable octet of valence electrons.
6. – nonpolar covalent – covalent – double covalent
7. 
8. Examples: – HCl's molecular polarity is more similar to water's polarity than H_2's polarity compared to water's – HCl and water both polar, H_2 nonpolar, like dissolves like – HCl polarity is more similar to water's polarity
9. Examples: – The electronegativity difference for HCl is 1.1, which is higher than the 0.6 for HI. – The difference for HCl is greater.
10. 0.090 g/L or 9.0×10^{-2} g/L
11. Acceptable responses include, but are not limited to: • At standard pressure, NH_3 has a higher boiling point than CF_4. • The melting point of CF_4 is lower.
12. *Examples:* — Methane is nonpolar because its molecules are symmetrical. — Charges are evenly distributed throughout the molecule.
13. 

Chapter 10– MOLES/STOICHIOMETRY

- GRAM FORMULA MASS (GFM): the total mass of any substance.

GFM = 1 mole = 6.02×10^{23} particles, atoms, molecules = 22.4 L

- Use dimensional analysis set up.
- EX: How many moles of $H_2(g)$ do you have if you have 34 grams of $H_2(g)$?

Table T has an equation to convert moles to grams, grams to moles only.

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

Determining the molecular mass when given the empirical mass:

$$\frac{\textbf{Mass of the Molecular Formula}}{\textbf{Mass of the Empirical Formula}}$$

- EX: The empirical formula is CH and the molecular mass is 26, what is the molecular formula?

$$26/13 = 2$$
$$2\ (CH) = C_2H_2$$

Mole – Mole:

- Answers how many moles of one element or compounds react with a given number of moles of another element or compound.

- **Example:** How many moles of Ca are needed to react completely with 6 moles of H_2O in the following reaction: $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$.

  1. Cross out anything in the equation that the problem is not asking for or about, in this case cross out $Ca(OH)_2$ and H_2. Now we are left with Ca and $2H_2O$. Write the coefficient of each under the element or compound.

  2. On top of the element or compound write the number of moles given in the problem.

  3. Set up a proportion:

Percent Composition (Table T)

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

Hydrate: a compound that incorporates water molecules into its fundamental solid structure. The compound has a dot after it followed by the number of water molecules attached. Ex: $CaSO_4 \cdot 2H_2O$. To calculate the percent of water in a hydrate, you must determine the mass of the whole, and the mass of the part.

- Remember that the water molecule remains as a compound (1 mole H_2O = 18g)
- **The symbol • means add, not multiply.**

% Composition of a Hydrate from a lab:

$$\frac{\textbf{Hydrate – Anhydrate}}{\textbf{Hydrate x 100}}$$

Name _____ Chap. 10 Chemical Calculations Review Constructed Response

1. Show a numerical setup for calculating the percent composition by mass of silicon in SiO_2.

2. 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.

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

   A 2.50-liter aqueous solution contains 1.25 moles of dissolved sodium chloride. The dissolving of NaCl(s) in water is represented by the equation below.

   $$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

   Determine the molarity of this solution.

Base your answers to questions **4** through **6** 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.

4. Write the empirical formula for ascorbic acid.

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

6. Determine the number of moles of vitamin C in an orange that contains 0.071 gram of vitamin C.

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

Glycine, NH_2CH_2COOH, is an organic compound found in proteins. Acetamide, CH_3CONH_2, is an organic compound that is an excellent solvent. Both glycine and acetamide consist of the same four elements, but the compounds have different functional groups.

In the space below, calculate the gram-formula mass of glycine. Your response must include *both* a numerical setup and the calculated result.

Base your answers to questions **8** and **9** 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.

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

8. 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.

9. Identify the type of chemical reaction represented by the balanced equation.

_____

10. A hydrated compound contains water molecules within its crystal structure. The percent composition by mass of water in the hydrated compound $CaSO_4 \cdot 2H_2O$ has an accepted value of 20.9%. A student did an experiment and determined that the percent composition by mass of water in $CaSO_4 \cdot 2H_2O$ was 21.4%.

Calculate the percent error of the student's experimental result. Your response must include *both* a correct numerical setup and the calculated result.

_____ %

11. Given the balanced equation representing a reaction:

$$2\ C_2H_6 + 7O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

Determine the total number of moles of oxygen that react completely with 8.0 moles of C_2H_6.

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

Gypsum is a mineral that is used in the construction industry to make drywall (sheetrock). The chemical formula for this hydrated compound is $CaSO_4 \cdot 2\ H_2O$. A hydrated compound contains water molecules within its crystalline structure. Gypsum contains 2 moles of water for each 1 mole of calcium sulfate.

What is the gram formula mass of $CaSO_4 \cdot 2\ H_2O$?

13. Given the compound $C_4H_{10}O_8$,

a Calculate the molar mass of the compound.

b Calculate the number of moles in 17.7 grams of the compound.

c What is the empirical formula for this compound?

Answer Key
Moles/Stoichiometry Review

1. $\frac{28.0855\ \text{g}}{28.0855\ \text{g}+2(15.9994\ \text{g})} \times 100$

   $\frac{28.1\ \text{u}}{60.1\ \text{u}} \times 100$

   $\frac{28(100)}{60}$

2. 138 g *or* for any value from 137.8 g to 138.3 g, inclusive

3. –0.500 M –0.50 M –.5 M

4. $C_3H_4O_3$

5. $\frac{6(16\ \text{g/mol})}{176\ \text{g/mol}} \times 100$

   $\frac{(96)(100)}{176}$

6. 4.0×10^{-4} mol *or* 0.00040 mol

7. • (1)(14.0 g/mol) + (2)(12.0 g/mol) + (2)(16.0 g/mol) + (5)(1.0 g/mol) = 75.0 g/mol • (1)(14) + (5)(1) + (2)(12) + (2)(16) • 75.0 g/mol

8. A correct numerical setup is shown: $3 = \frac{x}{20.0g} \times 100$ *or* (20)(0.03)

9. –decomposition –redox

10. • *Correct Numerical Setup Examples:* — $\frac{21.4\% - 20.9\%}{20.9\%} \times 100$ — $\frac{21.4 - 20.9}{20.9} \times 100$ • *Calculated Result Examples:* — 2% ; — 2.4% ; — 2.39%

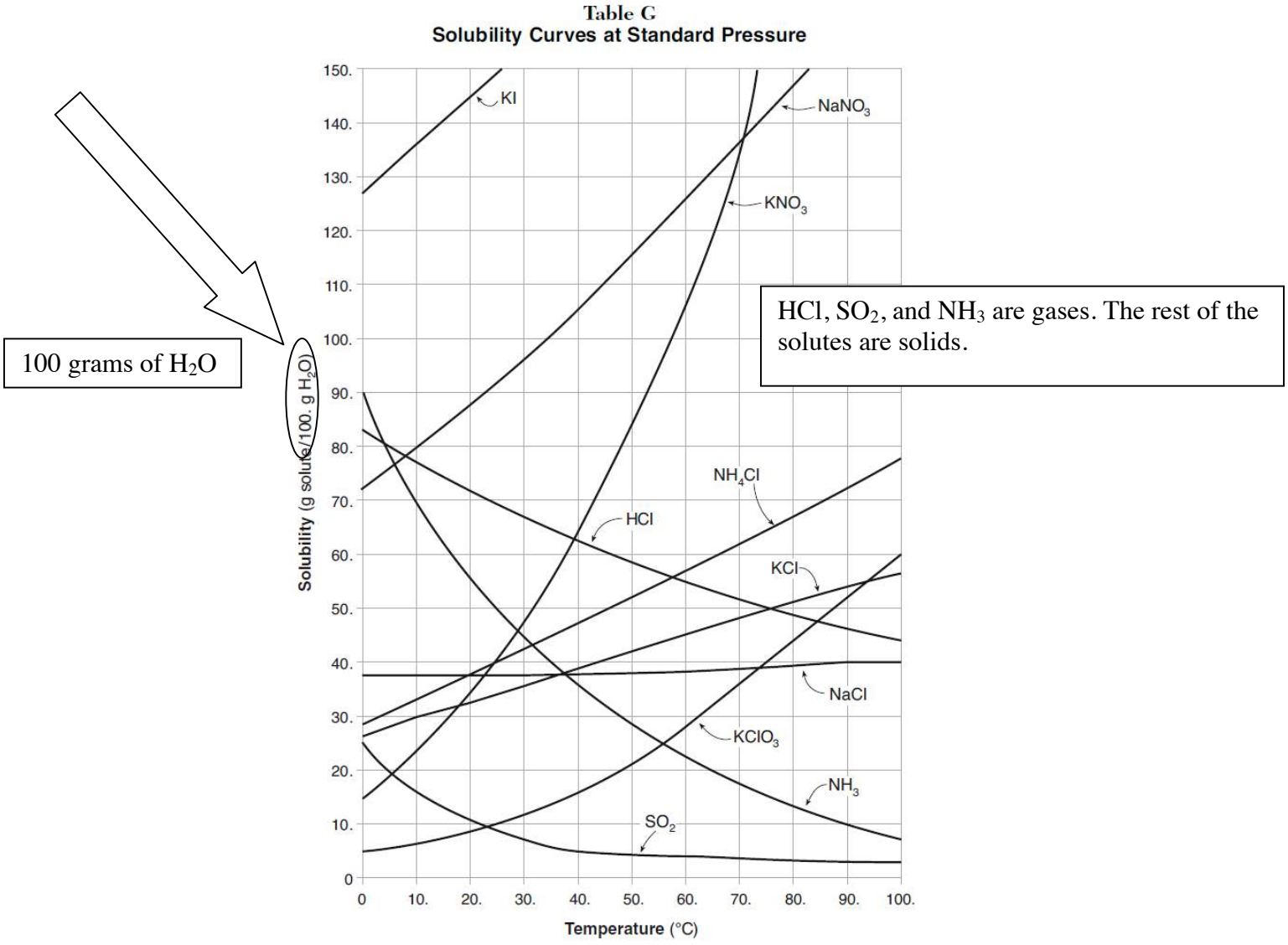
11. Examples: 28 mol

12. Acceptable responses: 172, 172.2.

13. a) 186 grams b) .095 moles c) $C_2H_5O_4$

Chapter 11 – SOLUTIONS/CONCENTRATION/MOLARITY

- **Solution:** a homogenous mixture made up of two or more substances. Example: when you mix salt and water, the salt is known as the **solute,** and the water is known as the **solvent.**
- **Solute:** a substance, like salt or sugar, that dissolves in water.
- **Solvent:** usually a liquid, for example water. If the solvent is water, then the solution is called an **aqueous solution.**
- **Precipitate:** INSOLUBLE (settle out, re-crystalize)
- **Solubility:** shows the most salt that water can hold, or the most salt that can dissolve in the water at a specific temperature.
    - **Saturated – On the line**
    - **Unsaturated – Below the line**
    - **Supersaturated – Above the line**
- **Dissociation:** When an ionic compound separates in solution to a (+) and (-) ion.



TEMPERATURE AFFECTS SOLUBILITY:

- As temperature increases, solubility of solids and liquids in solution also increases.
- As temperature increases, solubility of gases in solution decreases (HCl, NH_3, SO_2).

PRESSURE AFFECTS SOLUBILITY:

- No effect on solids & liquids
- As pressure increases, solubility of gases in solution also increases.

NATURE OF SOLUTE/SOLVENT AFFECTS SOLUBILITY:

- Likes dissolve in likes (polar in polar, and non-polar in non-polar)

DETERMINING IF COMPOUNDS ARE SOLUBLE OR INSOLUBLE

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

Soluble	Insoluble
Most soluble	Least soluble
Electrolyte	Poor/Non Electrolyte
High concentration of dissolved ions	Low concentration of dissolved ions
	PRECIPITATE

DESCRIBING CONCENTRATIONS OF A SOLUTION:

- You can describe the concentration of a solution by **molarity, percent by mass, or parts per million.**

Concentration	$\text{parts per million} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1\,000\,000$
	$\text{molarity} = \frac{\text{moles of solute}}{\text{liter of solution}}$

Remember some ways they can try and trick you:
1. May give you information to calculate moles (convert grams to moles)
2. May give you volume in ml (convert to liters)
3. May give you solute and solvent separately (must add them to get the liters or grams of solution)

- The higher the Molarity or ppm, the better conductor of electricity it will be.

COLLIGATIVE PROPERTIES:

Boiling Point Elevation & Freezing Point Depression:

- The presence of a solute (salt or sugar) raises the boiling point of the solvent.
- The presence of any solute (salt or sugar) lowers the freezing point of the solvent.

The Cold get Colder, and the Hot get Hotter.

Name: **Chapter 11 - Solutions** **Constructed Response**

Base your answers to questions **1** through **4** on the information below and on your knowledge of chemistry.

A student prepares two 141-gram mixtures, *A* and *B*. Each mixture consists of NH_4Cl, sand, and H_2O at 15°C. Both mixtures are thoroughly stirred and allowed to stand. The mass of each component used to make the mixtures is listed in the data table below.

Mass of the Components in Each Mixture

Component	Mixture A (g)	Mixture B (g)
NH_4Cl	40.	10.
sand	1	31
H_2O	100.	100.

1. Describe *one* property of sand that would enable the student to separate the sand from the other components in mixture *B*.

2. Determine the temperature at which all of the NH_4CL in mixture *A* dissolves to form a saturated solution.

3. Which type of mixture is mixture *B*?

4. State evidence from the table indicating that the proportion of the components in a mixture can vary.

---

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

A 2.50-liter aqueous solution contains 1.25 moles of dissolved sodium chloride. The dissolving of NaCl(s) in water is represented by the equation below.

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

Compare the freezing point of this solution to the freezing point of a solution containing 0.75 mole NaCl per 2.50 liters of solution.

6. 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.

7. What is the mass of $KNO_3(s)$ that must dissolve in 100. grams of water to form a saturated solution at 50.°C?

8. 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.

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

Explain, in terms of the distribution of particles, why the solution is a homogeneous mixture.

9. 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 **10** and **11** on the information below.

A 2.0-liter aqueous solution contains a total of 3.0 moles of dissolved NH_4Cl at 25°C and standard pressure.

10. Identify the *two* ions present in the solute.

11. Determine the molarity of the solution.

Answer Key

Solutions Constructed Response

1. —Sand is insoluble in water. —Sand particles are too large to pass through filter paper. —Sand is more dense than $NH_4Cl(aq)$. —Sand remains a solid in the mixture.
2. $23^\circ C$ to $26^\circ C$
3. —heterogeneous —nonuniform mixture
4. —The ratio by mass of NH_4Cl to H_2O in mixture *A* is 40. g/100. g, and the ratio in mixture *B* is 10. g/100. g. —Both mixtures have the same total mass, but have different amounts of sand. —Mixture *B* has more sand. —The mixtures have different proportions of NH_4Cl.
5. –The solution that contains 1.25 moles of NaCl has a lower freezing point. –lower for the first one –higher for the solution with 0.75 mol –The 0.30 M solution has a higher freezing point than the 0.50 M solution. –This solution has a lower f.p.
6. 0.70 M
7. 84g $\pm$ 2g
8. —The H^+ ions and the Cl^- ions are distributed uniformly throughout the solution. —There is an even distribution of $H^+(aq)$ and $Cl^-(aq)$.
9. unsaturated solution
10. Acceptable responses include, but are not limited to: • NH_4^+ and Cl^- • ammonium and chloride
11. 1.5 M

Chapter 12 - KINETICS & EQUILIBRIUM REVIEW SHEET

1. Collision Theory: in order for a chemical reaction to occur, effective collision of molecules must occur. Both the energy of the collision and the angle of the collision are important. The more collisions, the faster the rate of reaction.
    - Concentration: an increase will increase rate of reaction
    - Temperature: an increase will increase rate of reaction
    - Surface area: increasing SA will increase rate of reaction
    - Nature of compound: Ionic will react faster than covalent
    - Catalyst: increases rate of reaction by decreasing activation energy which is the energy required for a reaction to start.

2. Potential Energy Diagram: (Must know how to label and draw in a catalyst)

$$\Delta H = H_{product} - H_{reactant}$$

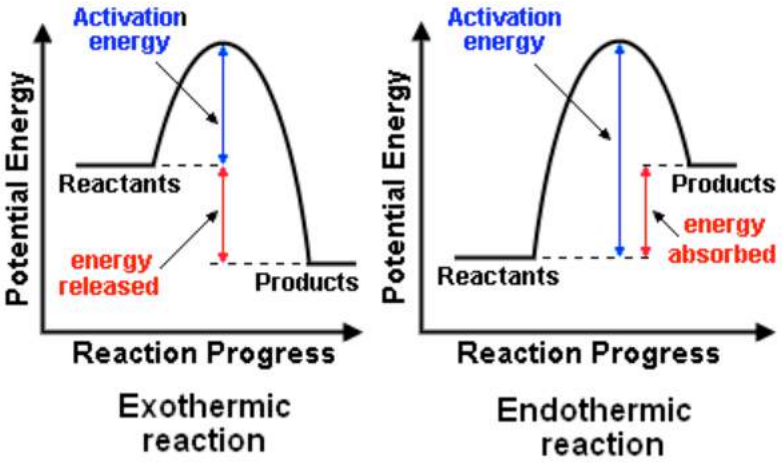
- ΔH = Heat of Reaction (Enthalpy)
- If ΔH is (+) then it is an endothermic reaction
- If ΔH is (-) then it is an exothermic reaction

Table I:

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
$2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(\ell)$	–10943
$2CH_3OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(\ell)$	–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
$H_2(g) + I_2(g) \longrightarrow 2HI(g)$	+53.0
$KNO_3(s) \xrightarrow{H_2O} K^+(aq) + NO_3^-(aq)$	+34.89
$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$	–44.51
$NH_4Cl(s) \xrightarrow{H_2O} NH_4^+(aq) + Cl^-(aq)$	+14.78
$NH_4NO_3(s) \xrightarrow{H_2O} NH_4^+(aq) + NO_3^-(aq)$	+25.69
$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$	+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.



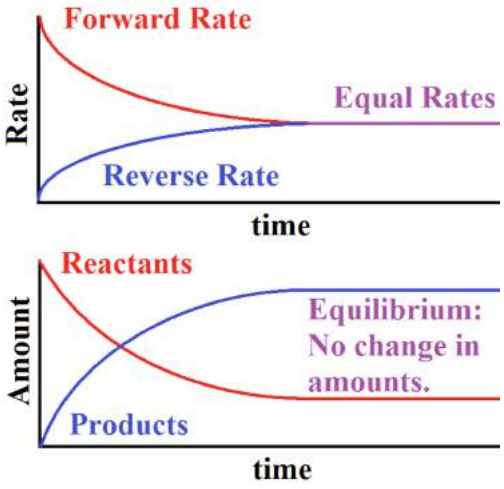
3. Equilibrium:

1. **Phase Equilibrium**: Melting/Freezing & Evaporation/Condensation (plateaus on the heating curve)

2. **Solution Equilibrium**: This is a SATURATED SOLUTION. Means that the rate of dissolving is equal to the rate of settling out.

3. **Chemical Equilibrium:** $R \rightleftharpoons P$

- This is a reversible reaction, can go in the forward and reverse direction.
- When the rate of the forward and the reverse reactions are EQUAL!!
- The concentrations are CONSTANT!!



4. LeChatelier's Principle: When a reaction is stressed, it will move in the direction to relieve the stress. Must be able to determine which side the stress is on, what direction the reaction will shift (left or right), and what will occur to the reactants and products of the reaction.

Factors that can stress a reaction:

- **Concentration** (may increase or decrease)
- **Temperature** (may increase or decrease): notice which side the heat is being released, and that will be the side that is affected by the temperature increase or decrease.
- **Pressure** (ONLY HAS AN EFFECT ON GASES): must look at the number of moles on each side of the reaction and if there is a difference is # of moles, than pressure will have an effect on the side with more moles. If the number of moles is the same, then pressure has NO EFFECT on the reaction.
- **Catalyst:** DOES NOT STRESS THE REACTION; a catalyst will increase the rate of the forward and reverse reaction equally by decreasing the activation energy of the reaction.

$$2\ SO_3(g) \rightleftharpoons 2\ SO_2\ (g) + O_2\ (g)\ \Delta H^o = 197.78\ kJ$$

+ΔH = Endothermic

4. Spontaneous Reactions: Two conditions are necessary:
    - Favor EXOTHERMIC REACTIONS – no energy required
    - High Entropy, Low Energy

5. Reactions go to completion: Reactions that only go in one direction and not the reverse. Some products that may occur to show that a reaction has gone to completion are:
    - The formation of a GAS
    - The formation of WATER
    - The formation of a PRECIPITATE (may have to refer to Table F remember that a precipitate is INSOLUBLE).

Name _____ **Chap. 12 Kinetics & Equilibrium Review** **Constructed Response**

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

   A few pieces of dry ice, $CO_2(s)$, at $-78^\circ C$ are placed in a flask that contains air at $21^\circ C$. The flask is sealed by placing an uninflated balloon over the mouth of the flask. As the balloon inflates, the dry ice disappears and no liquid is observed in the flask.

   Compare the entropy of the CO_2 molecules in the dry ice to the entropy of the CO_2 molecules in the inflated balloon.

2. 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.

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

   Describe how the catalyst, zymase, speeds up this reaction.

   Base your answers to questions **3** through **5** 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.

   $$Cl_2(g) + 2OH^-(aq) \rightleftharpoons ClO^-(aq) + Cl^-(aq) + H_2O(\ell)$$

3. State the effect on the concentration of the ClO^- ion when there is a *decrease* in the concentration of the OH^- ion.

4. Explain why the container must be closed to maintain equilibrium.

5. Compare the rate of the forward reaction to the rate of the reverse reaction for this system.

Base your answers to questions **6** and **7** on the information below.

At standard pressure, hydrogen peroxide, H_2O_2, melts at –0.4°C, boils at 151°C, and is very soluble in water. A bottle of aqueous hydrogen peroxide, $H_2O_2(aq)$, purchased from a pharmacy has a pressure-releasing cap. Aqueous hydrogen peroxide decomposes at room temperature, as represented by the balanced equation below.

$$2H_2O_2(aq) \longrightarrow 2H_2O(\ell) + O_2(g) + 196.0\ kJ$$

6. Explain why a hydrogen peroxide bottle needs a pressure-releasing cap.

7. State evidence that indicates the decomposition of $H_2O_2(aq)$ is exothermic.

---

8. State *two* methods to increase the rate of a chemical reaction and explain, in terms of particle behavior, how each method increases the reaction rate.

---

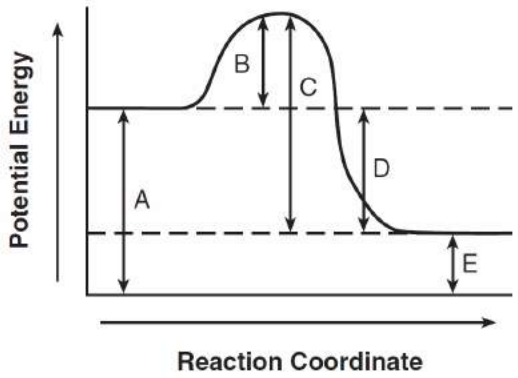
9. Several steps are involved in the industrial production of sulfuric acid. One step involves the oxidation of sulfur dioxide gas to form sulfur trioxide gas. A catalyst is used to increase the rate of production of sulfur trioxide gas. In a rigid cylinder with a movable piston, this reaction reaches equilibrium, as represented by the equation below.

$2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g) + 392\ kJ$

Determine the amount of heat released by the production of 1.0 mole of $SO_3(g)$.

Base your answers to questions **10** and **11** on the information below.

The chemical reaction between methane and oxygen is represented by the potential energy diagram and balanced equation below.



$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 890.4\ kJ$

10 Explain, in terms of collision theory, why a lower concentration of oxygen gas *decreases* the rate of this reaction.

11. Which potential energy interval in the diagram represents the activation energy of the forward reaction?

---

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

A 1.0-gram strip of zinc is reacted with hydrochloric acid in a test tube. The unbalanced equation below represents the reaction.

____ $Zn(s)$ + ____ $HCl(aq) \rightarrow$ ____ $H_2(g)$ + ____ $ZnCl_2(aq)$

Explain, in terms of collision theory, why using 1.0 grams of powdered zinc, instead of the 1.0-gram strip of zinc, would have increased the rate of the reaction.

13. Base your answer to the following question on the information below.
Given the reaction at equilibrium:

$2NO_2(g) \leftrightarrow N_2O_4(g) + 55.3\ kJ$

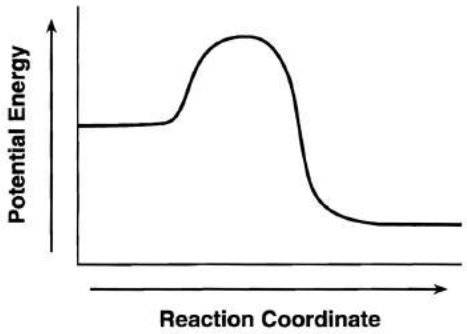
Explain, in terms of Le Chatelier's principle, why the equilibrium shifts to the right to relieve the stress when the pressure on the system is increased at constant temperature.

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

    The catalytic converter in an automobile changes harmful gases produced during fuel combustion to less harmful exhaust gases. In the catalytic converter, nitrogen dioxide reacts with carbon monoxide to produce nitrogen and carbon dioxide. In addition, some carbon monoxide reacts with oxygen, producing carbon dioxide in the converter. These reactions
    are represented by the balanced equations below.

    Reaction 1: $2NO_2(g) + 4CO(g) \longrightarrow N_2(g) + 4CO_2(g) +$ 1198.4 kJ
    Reaction 2: $2CO(g) + O_2(g) \longrightarrow 2CO_2(g) +$ 566.0 kJ

    The potential energy diagram below represents reaction 1 without a catalyst. On the same diagram, draw a dashed line to indicate how potential energy changes when the reaction is catalyzed in the converter.



Answer Key

Kinetics & Equilibrium Review

1. The CO_2 molecules in the dry ice have less entropy than the CO_2 molecules in the inflated balloon. –The CO_2 gas in the balloon is more disordered. –less for $CO_2(s)$

2. –Zymase is a catalyst that provides an alternative pathway, which requires less energy. –decreases the activation energy –changes the reaction mechanism

3. – The concentration of the ClO^- ion decreases. –$[ClO^-]$ decreases. – lower ClO^- concentration – less ClO^-

4. – The container must be closed so that no matter can enter or leave, thus distributing the equilbrium. – If the container is open, Cl_2 gas escapes. – to keep the concentration of the reactants and products constant

5. – The rate of the forward reaction is equal to the rate of the reverse reaction. – They are the same. – equal

6. —The excess pressure due to the production of oxygen gas in the bottle needs to be gradually released. —As $O_2(g)$ is produced, the pressure inside of the bottle might increase and the bottle might burst without the pressure-releasing cap.

7. —More energy is released than absorbed. —Heat is a product of the reaction.

8. *Examples:* – Increasing the temperature of the reaction causes the reacting particles to move faster and collide more frequently.– Increasing the concentration increases the number of particle collisions.– Increasing the surface area (solid reactant) allows a greater number of particles to collide.– Adding a catalyst provides an alternate way for the particles to react.

9. 196 kJ

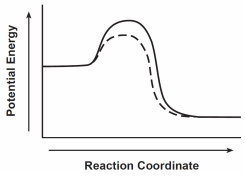
10. Acceptable responses include, but are not limited to: • A lower concentration of oxygen gas decreases the number of effective collisions between O_2 molecules and CH_4 molecules.

11. *B*

12. Example: The greater surface area in powdered zinc would have resulted in more frequent collisions between the zinc atoms and the hydrogen ions in the HCl(aq).

13. *Examples*: – Equilibrium shifts towards the fewer number of moles of gas – The reaction shifts to the side that would result in a reduction of pressure – fewer moles of gas, less pressure

14. An appropriate line is drawn.



Chapter 13 – ACIDS, BASES, & SALTS

- **Electrolyte:** a substance that dissolves in water and forms a solution that conducts an electric current. It's the IONS in solution that conduct electricity.
- Remember the higher the concentration of dissolved ions, the more soluble a solution and the better conductor of electricity (the higher the Molarity, the better conductor of electricity it is)

CHARACTERISTICS OF AN ACID: (Table K)

- Acids are always aqueous and conduct electricity (the ions make them electrolytes).
- Acids (ex. HCl) react with certain metals to produce H_2 - **TABLE J**
- Acids cause color changes in acid-base indicators: **Blue Litmus (turns red in an acid); Phenolphthalein (colorless in an acid).**
- Acids react with bases to form a salt and water. This is called a **Neutralization Reaction:** Acid Base Water Salt

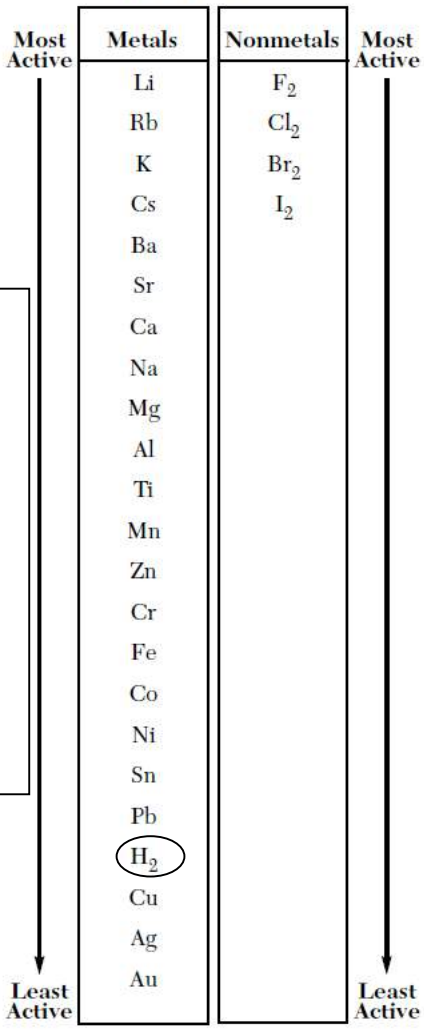
$$HCl + NaOH \rightarrow H_2O + NaCl$$

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)

COOH indicates an organic acid

Table J
Activity Series**



All metals above H_2 will react with acids in a single replacement reaction to produce H_2 gas. All metals below do not react with acids.

**Activity Series is based on the hydrogen standard. H_2 is *not* a metal.

CHARACTERISTICS OF A BASE: (Table L)

- Bases in an aqueous solution (in water) conduct electricity (bases are electrolytes).
- Bases cause color changes in acid-base indicators: **Red Litmus (turns blue in a base); Phenolphthalein (pink in a base).**
- Bases react with acids to form salts and water – **Neutralization Reaction (see example above).**

Table L
Common Bases

Formula	Name
NaOH(aq)	sodium hydroxide
KOH(aq)	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

Be careful of TRICKS!!
Organic alcohols also have OH, but are attached to carbon and hydrogen.
Example: CH_3OH
NOT A BASE!!!!

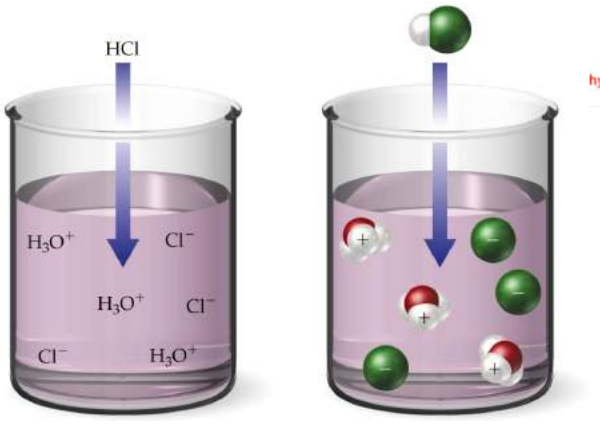
SALT:

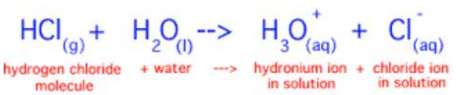
- An ionic compound that has positive ions other than hydrogen (H), and negative ions other than hydroxide (OH^-). Example: NaCl (Na^+ and Cl^-).
- Salts conduct electricity (salts are electrolytes).

****ACIDS, BASES & SALTS ARE ELECTROLYTES****

ARRHENIUS THEORY OF ACIDS & BASES:

- An **Arrhenius acid** has hydrogen and releases hydrogen ions in aqueous solutions. Examples of Arrhenius acids on Table K.
- When these acids are dissolved in H_2O, the H^+ is the only positive ion in the solution; when the H^+ is released in the water solution attaches to the water to produce H_3O^+ (hydronium ions).

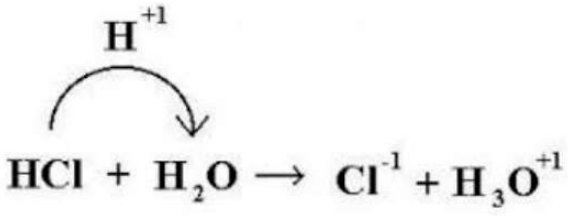




- An **Arrhenius base** has OH^- (hydroxide), and releases hydroxide ions (OH^-) in an aqueous solution. Examples of Arrhenius bases are on Table L.
- In these bases, the OH^- is the only negative ion in the solution.

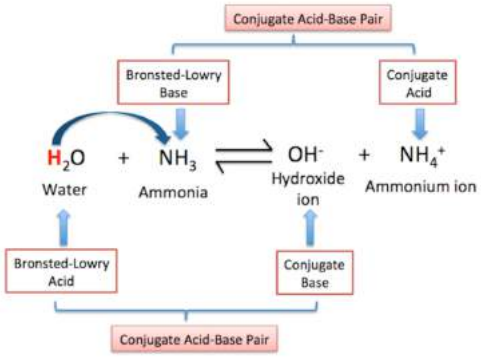
BRONSTED-LOWERY THEORY (ALTERNATE THEORY): BAAD

- An acid is an H^+ donor (proton donor); it gives away an H^+, a proton.
- A base is an H^+ acceptor (proton acceptor); it accepts H^+, a proton



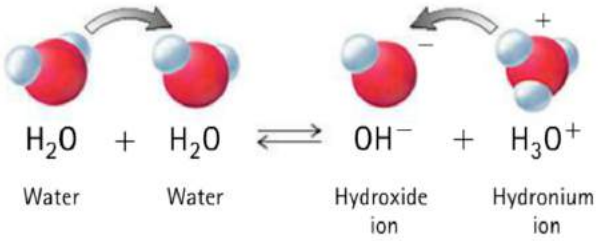
CONJUGATE ACIDS/BASES:

- A conjugate base is what remains after the acid donates a proton
- A conjugate acid is what is formed when a base accepts a proton.



AMPHOTERIC SUBSTANCES:

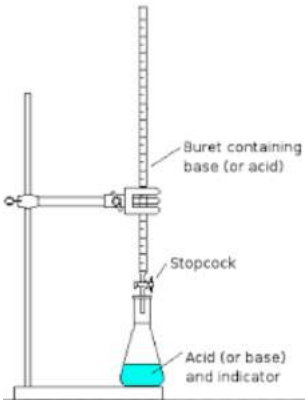
- An amphoteric substance can sometimes act like an acid and sometimes act like a base. Examples are **H_2O & HSO_4^-.**



TITRATION:

- A technique where a solution of known concentration is used to determine the concentration of an unknown solution.
- The set up requires a Burette. The endpoint of titration is the point when the indicator changes color. At neutralization, moles of acid = moles of base ($H^+ = OH^-$). Basic form of the titration equation on **TABLE T.**

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



INDICATORS (TABLE M):

- Indicators change color when pH changes. They show whether a solution is an acid or base and how strong.
- The first number and less corresponds with the first color
- The second number and greater corresponds with the second color
- Any pH value between means that a color between will appear like yellow and blue make green.

Table M
Common Acid–Base Indicators

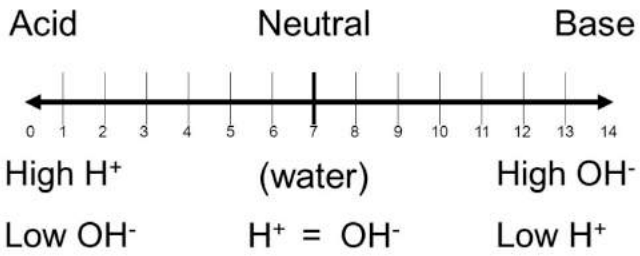
Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.1–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8–9	colorless to pink
litmus	4.5–8.3	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

Source: *The Merck Index*, 14th ed., 2006, Merck Publishing Group

MEANING OF pH:

$$pH = -log[H^+]$$

pH Scale – a measure of the concentration of Hydrogen Ions



- If the H^+ concentration is 1.0×10^{-3}, then the pH is 3 (acidic)
- If the H^+ concentration is 1.0×10^{-12}, then the pH is 12 (basic)
- As we go from 0-14 the H^+ concentration decreases
- As we go from 14-0 the H^+ concentration increases
- Every step on the pH scale is 10x
- Going from a pH of 3 to 4 means the H^+ concentration decreases 10x
- Going from as pH of 5 to 3 means the H^+ concentration increases 100x

Name _____ Chap. 13 Acids & Bases Review Constructed Response

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Base your answers to questions **1** through **4** on the information below and on your knowledge of chemistry.

A $NaOH(aq)$ solution and an acid-base indicator are used to determine the molarity of an $HCl(aq)$ solution. A 25.0-milliliter sample of the $HCl(aq)$ is exactly neutralized by 15.0 milliliters of 0.20 M $NaOH(aq)$.

1. Using the data, determine the concentration of the $HCl(aq)$.

2. Based on the data, the calculated molarity of the $HCl(aq)$ should be expressed to what number of significant figures?

3. Complete the equation for the neutralization reaction that occurs, by writing a formula for *each* product.

$HCl(aq) + NaOH(aq) \rightarrow$ _____ + _____

4. Identify the laboratory process described in this passage.

Base your answers to questions **5** and **6** on the information below and on your knowledge of chemistry.

The pH of various aqueous solutions are shown in the table below.

pH of Various Aqueous Solutions

Aqueous Solution	pH
$HCl(aq)$	2
$HC_2H_3O_2(aq)$	3
$NaCl(aq)$	7
$NaOH(aq)$	12

5. State how many times greater the hydronium ion concentration in the HCl(aq) is than the hydronium ion concentration in the $HC_2H_3O_2(aq)$.

6. Complete the table by writing the color of thymol blue in the NaCl(aq) and in the NaOH(aq) solutions.

Aqueous Solution	Color of Thymol Blue
$NaCl(aq)$	
$NaOH(aq)$	

Base your answers to questions **7** and **8** on the information below.

In a titration, 20.0 milliliters of 0.15 M HCl(aq) is exactly neutralized by 18.0 milliliters of KOH(aq).

7. Determine the concentration of the KOH(aq).

8. 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).

Base your answers to questions **9** and **10** on the information below.

Some carbonated beverages are made by forcing carbon dioxide gas into a beverage solution. When a bottle of one kind of carbonated beverage is first opened, the beverage has a pH value of 3.

9. After the beverage bottle is left open for several hours, the hydronium ion concentration in the beverage solution decreases to $\frac{1}{1000}$ of the original concentration. Determine the new pH of the beverage solution.

10. State, in terms of the pH scale, why this beverage is classified as acidic.

Base your answers to questions **11** and **12** 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.

11. Identify the negative ion found in milk of magnesia.

12. Compare the hydrogen ion concentration in tomato juice to the hydrogen ion concentration in milk of magnesia.

Answer Key
Acids & Bases Review

1. 0.12 M
2. 2 *or* two
3. $HCl(aq) + NaOH(aq) \rightarrow H_2O(\ell) + NaCl(aq)$

   $HCl(aq) + NaOH(aq) \rightarrow NaCl + HOH$
4. —titration
   —volumetric analysis
5. 10
6.

Aqueous Solution	Color of Thymol Blue
NaCl(aq)	yellow
NaOH(aq)	blue

7. 0.17 M
8. —The number of moles of $H^+(aq)$ ions equals the number of moles of $OH^-(aq)$ ions. — The number of hydrogen ions is the same as the number of hydroxide ions.
9. – 6
10. – The beverage is acidic because its pH value is below 7. – A pH of 3 is in the acid range on the pH scale.
11. $OH^-(aq)$ *or* OH^- *or* hydroxide ion
12. –The H^+ ion concentration in tomato juice is 10^6 times greater.
    –The hydrogen ion concentration in tomato juice is greater than that in milk of magnesia.
    –Milk of magnesia has a lower concentration of H_3O^+ ions.

Chapter 14 – REDOX

- **Oxidation:** The loss of electrons by a molecule, atom, or ion.
  $Na \rightarrow Na^{+} + e^{-}$ (electon(s) on the right means lost)

- **Reduction:** The gain of electrons by a molecule, atom, or ion.
  $P + 3e^{-} \rightarrow P^{-3}$ (electron(s) on the left means gained)

"OIL RIG"

Oxidation is Loss, Reduction is Gain

- **Reducing Agent:** An electron donor. Something that is oxidized is considered a reducing agent.
- **Oxidizing Agent:** An electron acceptor. Something that is reduced is considered an oxidizing agent

$Na \rightarrow Na^{+} + e^{-}$
$P + 3e^{-} \rightarrow P^{-3}$
$Na^{+} + P^{-3} \rightarrow Na_3P$ (sodium phosphide)
Na is being oxidized; it is the reducing agent
P is being reduced; it is the oxidizing agent

- **Redox:** Short-hand for an oxidation/reduction equation. In a single reaction there is both oxidation and reduction.
- Redox reactions have conservation of matter, charge, and energy just like any other chemical reaction. You know that a reaction is a REDOX reaction because some oxidation numbers will change from the reactant side to the product side.
- **Half-reaction:** shows either oxidation or reduction. A redox reaction is made up of two half reactions (one oxidation, and one reduction).

THE FIRST STEP ALWAYS IS TO ASSIGN OXIDATION NUMBERS TO ALL THE IONS!!

RULES FOR ASSIGNING OXIDATION NUMBERS:

1) Free elements (not combined with any other element) have an oxidation number of zero. Ex: Na, O_2, H_2

2) All metals in Group 1 have an oxidation number of +1.

3) All metals in Group 2 have an oxidation number of +2.

4) F (fluorine) always has an oxidation of $^{-}1$.

5) The oxidation of simple ions is equal to the charge on the ion. Ex: Mg^{+2} has an oxidation number of +2.

6) The sum of the oxidation numbers **must equal 0.** Examples: sodium chloride: Na^{+1} + Cl^{-1} = NaCl (sum of the oxidation numbers equals 0); magnesium chloride: Mg^{+2} + Cl^{-1} = $MgCl_2$ (sum of the oxidation numbers equals 0).

7) In ions (charged particles), the sum of the oxidation numbers of all the atoms must equal the charge of the ion. **Example:** sulfate ion SO_4^{-2} . O has an oxidation number of ⁻2, and therefore (⁻2) x (4) = ⁻8. Remember that the overall charge of this ion has to be ⁻2, so what must the oxidation number of S be?

8) Oxygen has an oxidation number of ⁻2 in all its compounds **except in peroxides (Ex: H_2O_2),** when oxygen has an oxidation number of ⁻1, and **in compounds with F (Ex OF_2),** when oxygen has an oxidation number of +2.

9) Hydrogen has an oxidation number of +1 in all compounds combined with a nonmetal. The exception is **in metal hydrides (metal and hydrogen, LiH, and CaH_2),** when hydrogen has an oxidation number of ⁻1.

Activity Series – Table J:

Table J
Activity Series**

Most Active	Metals	Nonmetals	Most Active
	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		
Least Active	Au		Least Active

**Activity Series is based on the hydrogen standard. H_2 is *not* a metal.

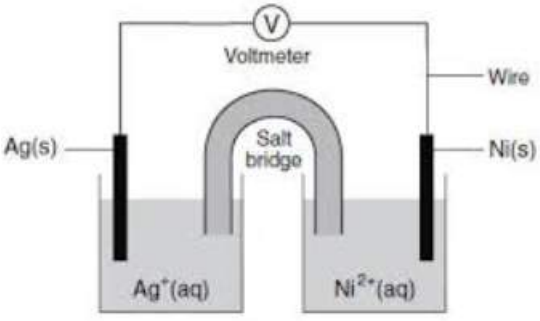
- SINGLE REPLACEMENT REACTIONS ARE **ALWAYS** REDOX
- DOUBLE REPLACMENT REACTIONS ARE **NEVER** REDOX
- Metals are easily oxidized (they lose electrons), Non-metals are easily reduced (gain electrons)
- A spontaneous redox reaction will occur when during a single replacement reaction the free element is more reactive than the element in solution. If the free element is not more reactive, then the reaction will not occur.

Voltaic Cell

- Spontaneous
- Converts chemical → electrical energy Anode is **always** the site of oxidation (in voltaic cell it is negatively charged)
- Cathode is **always** the site of reduction (in voltaic cell it is positively charged)
- Two half-cells each with an electrode (metal strip)
- The half-cell with the metal that is more reactive (Table J) will be the anode, and will therefore be oxidized.
- Wire is to allow the electrons to travel.
- Salt bridge allows ions to travel freely between the two half-cells.

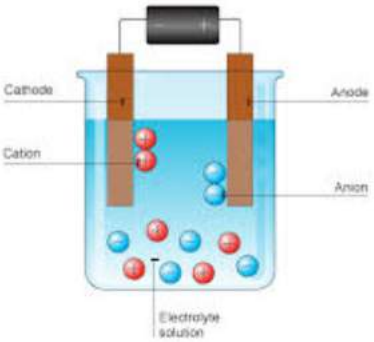
Electrolytic Cell

- Non-spontaneous
- converts electrical energy → chemical energy
- Requires an energy source (battery)
- Anode is the site of oxidation – the charge is positive
- Cathode is the site of reduction – the charge is negative.
- **Electroplating** is an electrolytic cell. The object that is going to be plated is the **cathode.**



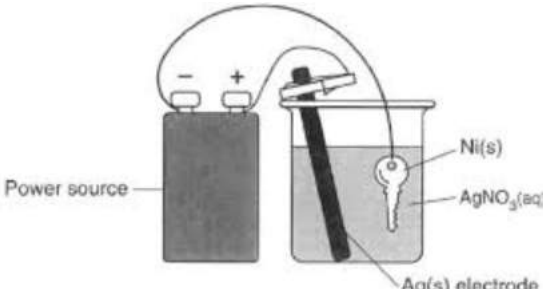
$2Ag^+(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

VOLTAIC



ELECTOLYTIC

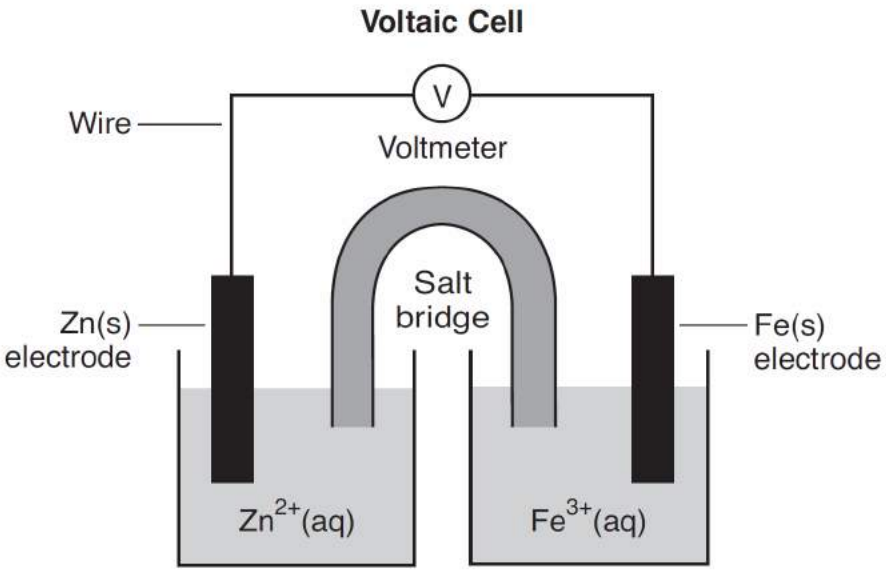
ELECTROPLATING



Name _____ Chap. 14 Oxidation & Reduction Constructed Response

Base your answers to questions **1** and **2** on the information below and on your knowledge of chemistry.

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.



$$Zn(s) + Fe^{3+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$$

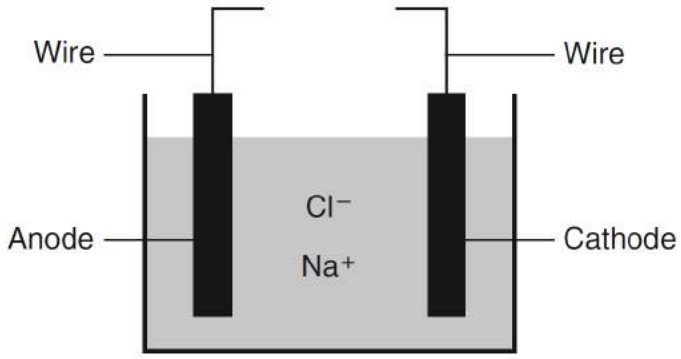
1. Explain, in terms of Zn atoms and Zn ions, why the mass of the Zn electrode *decreases as the cell operates.*

2. Identify the subatomic particles that flow through the wire as the cell operates.

Base your answers to questions **3** through **5** on the information below.

Metallic elements are obtained from their ores by reduction. Some metals, such as zinc, lead, iron, and copper, can be obtained by heating their oxides with carbon.

More active metals, such as aluminum, magnesium, and sodium, can not be reduced by carbon. These metals can be obtained by the electrolysis of their molten (melted) ores. The diagram below represents an incomplete cell for the electrolysis of molten NaCl. The equation below represents the reaction that occurs when the completed cell operates.



$$2NaCl(\ell) \rightarrow 2Na(\ell) + Cl_2(g)$$

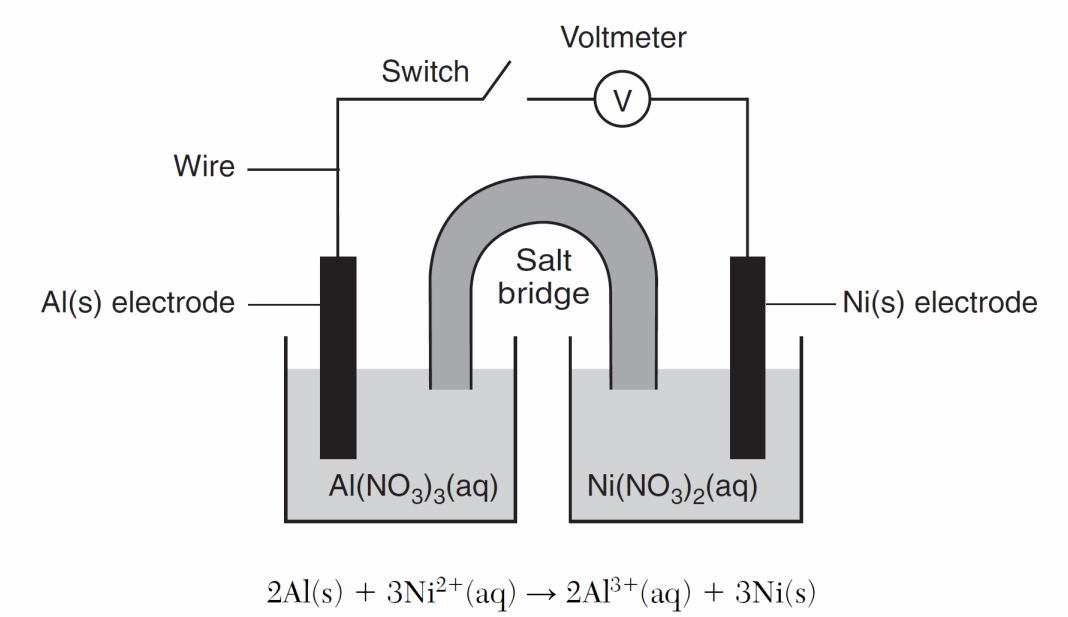
3. Write a balanced half-reaction equation for the reduction of the iron ions in iron(III) oxide to iron atoms.

4. Identify *one* metal from the passage that is more active than carbon and *one metal from the passage that is less active than carbon.*

5. Identify the component required for the electrolysis of molten NaCl that is missing from the cell diagram.

Base your answers to questions **6** through **9** on the information below.

A student constructs an electrochemical cell during a laboratory investigation. When the switch is closed, electrons flow through the external circuit. The diagram and equation below represent this cell and the reaction that occurs.



$$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$$

6. State, in terms of energy, why this cell is a voltaic cell.

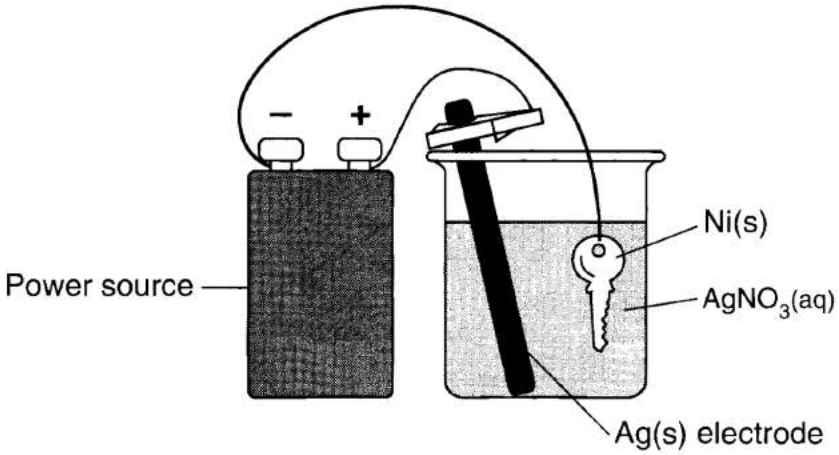
7. Determine the number of moles of Al(s) needed to completely react with 9.0 moles of $Ni^{2+}(aq)$ ions.

8. Write a balanced half-reaction equation for the oxidation that occurs when the switch is closed.

9. State the direction of electron flow through the wire when the switch is closed.

Base your answers to questions **10** through **12** on the information below.

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.



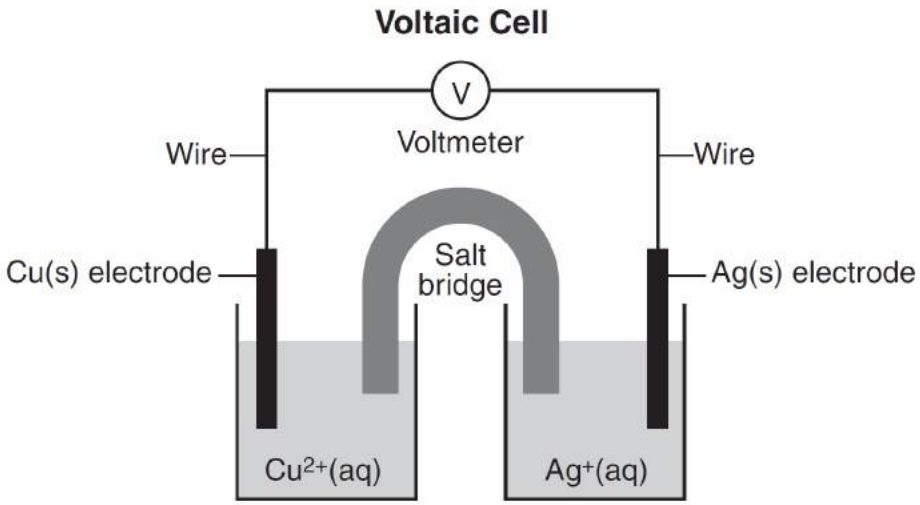
10. Explain, in terms of Ag atoms and $Ag^+(aq)$ ions, why the mass of the silver electrode *decreases* as the cell operates.

11. State the purpose of the power source in the cell.

12. Identify the cathode in the cell.

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

The diagram and balanced ionic equation below represent a voltaic cell with copper and silver electrodes and the reaction that occurs when the cell is operating.



$$Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$

State the purpose of the salt bridge in this voltaic cell.

Answer Key
Redox Review

1. –Zinc atoms from the electrode are oxidized to zinc ions in the solution, decreasing the mass of the electrode. –Zinc atoms become $Zn^{2+}(aq)$. –The atoms become ions dissolved in the water. –Zn atoms lose electrons, producing ions in solution.
2. –electrons $-e^{-}$ $-e$
3. $Fe^{3+} + 3e^{-} \rightarrow Fe$
4. –More active than carbon: aluminum, Mg, or Na –Less active than carbon: zinc, Pb, Fe, copper
5. –source of electrical energy –battery
6. —A spontaneous reaction converts chemical energy to electrical energy. —A battery is not required to provide energy for the cell to operate.
7. 6.0 mol.
8. —$Al \rightarrow 3e^{-} + Al^{3+}$ — $2Al \rightarrow 2Al^{3+} + 6e^{-}$
9. —Electrons flow from the Al electrode to the Ni electrode. —Electrons move left to right through the wire.
10. —Silver atoms lose electrons and become silver ions in the solution. —Some of the Ag atoms become Ag^{+} ions. —Silver atoms are oxidized to silver ions.
11. —The cell requires electrical energy for the non-spontaneous reaction to occur. —The power source causes some Ag(s) atoms to oxidize.
12. Ni(s) key / nickel
13. Acceptable responses include, but are not limited to: • The salt bridge allows for the migration of ions between the half-cells. • The salt bridge prevents polarization of the half-cells maintains electrical neutrality

How the heck am I supposed to know this????

1) Which liquid has the *highest* vapor pressure at 75° C?

A) propanone B) water C) ethanol D) ethanoic acid

2) Which liquid has the *lowest* vapor pressure at 65° C?

A) propanone B) water C) ethanoic acid D) ethanol

3) Which of the following compounds has the *lowest* vapor pressure at 50° C?

A) ethanol B) ethanoic acid C) propanone D) water

4) At 65° C, which of the following compounds has a vapor pressure of 58 kilopascals?

A) propanone B) ethanoic acid C) ethanol D) water

5) What is the name of the polyatomic ion in the compound Na_2O_2?

A) peroxide B) oxide C) oxalate D) hydroxide

6) As atomic number increases within Group 15 on the Periodic Table, atomic radius

A) decreases, then increases
B) increases, then decreases
C) decreases, only
D) increases, only

7) How do the atomic radius and metallic properties of sodium compare to the atomic radius and metallic properties of phosphorus?

A) Sodium has a larger atomic radius and is less metallic.
B) Sodium has a larger atomic radius and is more metallic.
C) Sodium has a smaller atomic radius and is more metallic.
D) Sodium has a smaller atomic radius and is less metallic.

8) Which two characteristics are associated with metals?

A) high first ionization energy and low electronegativity
B) low first ionization energy and high electronegativity
C) high first ionization energy and high electronegativity
D) low first ionization energy and low electronegativity

9) Which of the following elements has the *greatest* density at STP?

A) barium B) magnesium C) radium D) beryllium

10) At which Celsius temperature does lead change from a solid to a liquid?

A) 328° C B) 0° C C) 601° C D) 874° C

11) Which of the following elements has the *greatest* density at STP?

A) scandium B) selenium C) sodium D) silicon

12) Which of the following phrases describes the molarity of a solution?

A) moles of solution per liter of solution
B) liters of solution per mole of solution
C) liters of solute per mole of solution
D) moles of solute per liter of solution

13) What is the mass of NH_4Cl that must dissolve in 200. grams of water at 50.° C to make a saturated solution?

A) 104 g B) 42 g C) 26 g D) 84 g

14) An unsaturated aqueous solution of NH_3 is at 90.° C in 100. grams of water. According to the *Solubility Curves* chemistry reference table, how many grams of NH_3 could this unsaturated solution contain?

A) 10. g B) 20. g C) 5 g D) 15 g

15) Which of the following compounds is *least* soluble in water at 60.° C?

A) KNO_3 B) $KClO_3$ C) NH_4Cl D) NaCl

16) Which one of the following compounds is insoluble in water?

A) potassium bromide B) calcium bromide C) sodium bromide D) silver bromide

17) Which one of the following compounds is insoluble in water?

A) $KClO_3$ B) $BaSO_4$ C) Na_2S D) $CaCrO_4$

18) For which compound is the process of dissolving in water exothermic?

A) NaCl B) NH_4NO_3 C) NH_4Cl D) NaOH

19) Which equation represents an exothermic reaction at 298 K?

A) $C(s) + O_2(g) \rightarrow CO_2(g)$

B) $KNO_3(s) \xrightarrow{H_2O} K^+(aq) + NO^{3-}(aq)$

C) $N_2(g) + O_2(g) \rightarrow 2NO(g)$

D) $NH_4Cl(s) \xrightarrow{H_2O} NH^{4+}(aq) + Cl^-(aq)$

20) Which of the following substances is an Arrhenius acid?

A) CH_3COOCH_3 B) H_3PO_4 C) NaCl D) $Ba(OH)_2$

21) Which compound is an Arrhenius acid?

A) NaOH B) KCl C) H_2SO_4 D) NH_3

22) Which of the following substances is an Arrhenius base?

A) CH_3OH B) CH_3Cl C) LiCl D) LiOH

23) What 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) litmus B) bromthymol blue C) bromcresol green D) methyl orange

24) Which of the following indicators would *best* distinguish between a solution with a pH of 3.5 and a solution with a pH of 5.5?

A) thymol blue B) bromcresol green C) bromthymol blue D) litmus

25) Which one of the following metals is *more* active than H_2?

A) Au B) Ag C) Pb D) Cu

26) Which metal is *more* active than Ni and *less* active than Zn?

A) Cr B) Mg C) Pb D) Cu

27) Which of the following reactions occurs spontaneously?

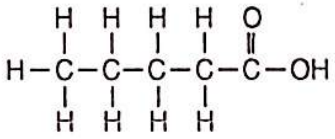
A) $I_2(s) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaI(aq)$

B) $Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaCl(aq)$

C) $I_2(s) + 2NaF(aq) \rightarrow F_2(g) + 2NaI(aq)$

D) $Cl_2(g) + 2NaF(aq) \rightarrow F_2(g) + 2NaCl(aq)$

28) Given the formula for an organic compound:



This compound is classified as an

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

29) What is the mass number of an alpha particle?

A) 1 B) 2 C) 0 D) 4

30) Which one of the following radioisotopes is matched with its decay mode?

A) H-3 and γ B) N-16 and α C) P-32 and β^- D) K-42 and β^-

31) Which of the following notations of a radioisotope is correctly paired with the notation of its emission particle?

A) ^{37}Ca and $^{4}_{2}He$ B) ^{16}N and $^{1}_{1}P$ C) ^{3}H and $^{0}_{-1}e$ D) ^{235}U and $^{0}_{+1}e$

32) Which radioisotope has an atom that emits a particle with a mass number of 0 and a charge of +1?

A) ^{19}Ne B) ^{239}Pu C) ^{3}H D) ^{16}N

33) What is the decay mode of ^{37}K?

A) γ B) α C) β^+ D) β^-

100 questions that you should never get wrong....

1) Which term is defined as a measure of the average kinetic energy of the particles in a sample?

A) temperature B) chemical energy C) thermal energy D) pressure

2) The particles in which sample of LiCl(s) have the same average kinetic energy as the particles in a 2.0-mole sample of $H_2O(\ell)$ at 25°C?

A) 1.0 mol at 75°C B) 2.0 mol at 50.°C C) 3.0 mol at 25°C D) 4.0 mol at 0°C

3) Which sample of ethanol has particles with the *highest* average kinetic energy?

A) 100.0 mL of ethanol at 45°C
B) 10.0 mL of ethanol at 25°C
C) 100.0 mL of ethanol at 35°C
D) 10.0 mL of ethanol at 55°C

4) Which substance can *not* be broken down by a chemical reaction?

A) ammonia B) water C) argon D) methane

5) Which one of the following substances can *not* be broken down by a chemical change?

A) water B) tungsten C) methane D) propanal

6) Which substance can *not* be decomposed by a chemical change?

A) HI B) H_2O C) Cu D) $AlCl_3$

7) Which statement describes a chemical property of the element magnesium?

A) Magnesium reacts with an acid.
B) Magnesium conducts electricity.
C) Magnesium is malleable.
D) Magnesium has a high boiling point.

8) Which statement describes a chemical property of hydrogen gas?

A) Hydrogen gas has a boiling point of 20. K at standard pressure.
B) Hydrogen gas burns in air.
C) Hydrogen gas is colorless.
D) Hydrogen gas has a density of 0.00009 g/cm^3 at STP.

9) Which of the following statements describes a chemical property of bromine?

A) Bromine has a reddish-brown color.
B) Bromine changes from a liquid to a gas at 332 K and 1 atm.
C) Bromine is soluble in water.
D) Bromine combines with aluminum to produce $AlBr_3$.

10) A 50.0-gram block of copper at 10.0°C is carefully lowered into 100.0 grams of water at 90.0°C in an insulated container. Which of the following statements describes the transfer of heat in this system?

A) The water gains heat and the block loses heat until both are at the same temperature that is between 10.0°C and 90.0°C.
B) The water loses heat to the block until both are at 10.0°C.
C) The block gains heat from the water until both are at 90.0°C.
D) The water loses heat and the block gains heat until both are at the same temperature that is between 10.0°C and 90.0°C.

11) What occurs when a 35-gram aluminum cube at 100.°C is placed in 90. grams of water at 25°C in an insulated cup?

A) Heat is transferred from the water to the aluminum, and the temperature of the water increases.
B) Heat is transferred from the aluminum to the water, and the temperature of the water increases.
C) Heat is transferred from the aluminum to the water, and the temperature of the water decreases.
D) Heat is transferred from the water to the aluminum, and the temperature of the water decreases.

12) A person with a body temperature of 37°C holds an ice cube with a temperature of 0°C in a room where the air temperature is 20.°C. The direction of heat flow is

A) from the ice to the person, only
B) from the person to the ice, only
C) from the ice to the person and air, and from the air to the person
D) from the person to the ice and air, and from the air to the ice

13) According to the wave-mechanical model of the atom, electrons in an atom

A) are most likely found in an excited state
B) have a positive charge
C) are located in orbitals outside the nucleus
D) travel in defined circles

14) In the wave-mechanical model of the atom, orbitals are regions of the most probable locations of

A) positrons B) electrons C) neutrons D) protons

15) According to the electron-cloud model of the atom, an orbital is a

A) region of the most probable electron location
B) circular path traveled by an electron around the nucleus
C) region of the most probable proton location
D) spiral path traveled by an electron toward the nucleus

16) What was concluded about the structure of the atom as the result of the gold foil experiment?

A) A negatively charged nucleus is surrounded by positively charged particles.
B) A positively charged nucleus is surrounded by positively charged particles.
C) A positively charged nucleus is surrounded by mostly empty space.
D) A negatively charged nucleus is surrounded by mostly empty space.

17) Which conclusion was a direct result of the gold foil experiment?

A) An atom is mostly empty space with a dense, positively charged nucleus.
B) An atom is composed of at least three types of subatomic particles.
C) An electron has properties of both waves and particles.
D) An electron has a positive charge and is located inside the nucleus.

18) Which statement describes how an atom in the ground state becomes excited?

A) The atom releases energy, and one or more electrons move to a lower electron shell.
B) The atom absorbs energy, and one or more electrons move to a higher electron shell.
C) The atom releases energy, and one or more electrons move to a higher electron shell.
D) The atom absorbs energy, and one or more electrons move to a lower electron shell.

19) During a flame test, ions of a specific metal are heated in the flame of a gas burner. A characteristic color of light is emitted by these ions in the flame when the electrons

A) emit energy as they move to higher energy levels
B) gain energy as they return to lower energy levels
C) gain energy as they move to higher energy levels
D) emit energy as they return to lower energy levels

20) The light emitted from a flame is produced when electrons in an excited state

A) release energy as they move to higher energy states
B) absorb energy as they move to higher energy states
C) absorb energy as they move to lower energy states
D) release energy as they move to lower energy states

21) What is the chemical formula for iron(III) oxide?

A) Fe_3O B) Fe_2O_3 C) FeO D) Fe_3O_2

22) Which formula represents copper(I) oxide?

A) CuO_2 B) CuO C) Cu_2O D) Cu_2O_2

23) Which of the following formulas represents strontium phosphate?

A) $Sr_3(PO_4)_2$ B) Sr_3PO_8 C) $SrPO_4$ D) $Sr_2(PO_4)_3$

24) Given the balanced equation representing a reaction:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) + 55.8\ kJ$$

In this reaction there is conservation of

A) mass, charge, and energy
B) mass and energy, only
C) mass and charge, only
D) mass, only

25) Given the balanced equation representing a reaction:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) + \text{energy}$$

In this reaction there is conservation of

A) charge and energy, only
B) charge, energy, and mass
C) mass, only
D) mass and charge, only

26) 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?

A) 36 g B) 18 g C) 56 g D) 72 g

27) The elements on the Periodic Table are arranged in order of increasing

A) atomic number B) atomic mass C) electronegativity D) boiling point

28) Elements on the modern Periodic Table are arranged in order of increasing

A) number of neutrons
B) atomic number
C) number of valence electrons
D) atomic mass

29) Which of the following elements requires the *least* amount of energy to remove the most loosely held electron from a gaseous atom in the ground state?

A) sodium B) silver C) calcium D) bromine

30) The amount of energy required to remove the outermost electron from a gaseous atom in the ground state is known as

A) activation energy
B) conductivity
C) electronegativity
D) first ionization energy

31) From which of these atoms in the ground state can a valence electron be removed using the *least* amount of energy?

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

32) Based on the *Properties of Selected Elements* chemistry reference table, which of the following atoms requires the *least* energy for the removal of the most loosely bound electron?

A) Sr B) Br C) Be D) Sn

33) The strength of an atom's attraction for the electrons in a chemical bond is the atom's

A) heat of reaction B) electronegativity C) heat of formation D) ionization energy

34) Which element has an atom with the *greatest* attraction for electrons in a chemical bond?

A) P B) As C) N D) Bi

35) Which element has an atom with the *greatest* tendency to attract electrons in a chemical bond?

A) sulfur B) chlorine C) carbon D) silicon

36) An atom of which element has the *greatest* attraction for the electrons in a bond with a hydrogen atom?

A) sulfur B) silicon C) phosphorus D) chlorine

37) Which statement explains why ozone gas, O_3, and oxygen gas, O_2, have different properties?

A) They have different molecular structures.
B) They have different oxidation numbers.
C) They have different electronegativities.
D) They are formed from different elements.

39) Which one of the following statements describes oxygen gas, $O_2(g)$, and ozone gas, $O_3(g)$?

A) They have the same molecular structure and the same properties.
B) They have different properties, only.
C) They have different molecular structures, only.
D) They have different molecular structures and different properties.

40) Which statement describes what occurs as two atoms of bromine combine to become a molecule of bromine?

A) Energy is released as a bond is broken.
B) Energy is absorbed as a bond is broken.
C) Energy is absorbed as a bond is formed.
D) Energy is released as a bond is formed.

41) Which phase change represents a decrease in entropy?

A) solid to liquid B) solid to gas C) gas to liquid D) liquid to gas

42) Which sample has the *lowest* entropy?

A) 1 mole of $H_2O(g)$ B) 1 mole of $KNO_3(s)$ C) 1 mole of $H_2O(\ell)$ D) 1 mole of $KNO_3(\ell)$

43) Given the balanced equation representing a reaction:

$$Br_2 + \text{energy} \rightarrow Br + Br$$

Which of the following statements describes the energy change and bonds in this reaction?

A) Energy is released as bonds are broken.
B) Energy is released as bonds are formed.
C) Energy is absorbed as bonds are formed.
D) Energy is absorbed as bonds are broken.

44) Given the balanced equation:

$$I + I \rightarrow I_2$$

Which statement describes the process represented by this equation?

A) A bond is broken and energy is released.
B) A bond is formed as energy is absorbed.
C) A bond is formed and energy is released.
D) A bond is broken as energy is absorbed.

45) Which compound contains *both* ionic and covalent bonds?

A) ammonia B) sodium nitrate C) potassium chloride D) methane

46) Which compound has *both* ionic and covalent bonding?

A) $C_6H_{12}O_6$ B) CH_3OH C) $CaCO_3$ D) CH_2Cl_2

47) The chemical bonding in sodium phosphate, Na_3PO_4, is classified as

A) both covalent and ionic
B) both covalent and metallic
C) metallic, only
D) ionic, only

48) Which of the following formulas represents a molecular compound?

A) LiOH B) Kr C) N_2O_4 D) NaI

49) Which of the following formulas represents a molecular compound?

A) HI B) LiCl C) KI D) KCl

50) The gram-formula mass of NO_2 is defined as the mass of

A) one mole of NO_2
B) two moles of NO
C) one molecule of NO_2
D) two molecules of NO

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

A) 263 g/mol B) 248 g/mol C) 279 g/mol D) 310. g/mol

52) Given the balanced equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$$

What is the total number of moles of CO_2 formed when 20. moles of HCl is completely consumed?

A) 20. mol B) 5.0 mol C) 10. mol D) 40. mol

53) 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?

A) 80. mol B) 20. mol C) 16 mol D) 64 mol

54) 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 $C_3H_8(g)$?

A) 0.30 mol B) 4.5 mol C) 7.5 mol D) 1.5 mol

55) According to the kinetic molecular theory, which of the following statements describes the particles in a sample of an ideal gas?

A) The force of attraction between the gas particles is strong.
B) The motion of the gas particles is random and straight-line.
C) The collisions between the gas particles cannot result in a transfer of energy between the particles.
D) The separation between the gas particles is smaller than the size of the gas particles themselves.

56) Which statement describes the particles of an ideal gas based on the kinetic molecular theory?

A) The gas particles have attractive forces between them.
B) The gas particles are relatively far apart and have negligible volume.
C) The gas particles have collisions without transferring energy.
D) The gas particles are in constant, nonlinear motion.

57) Which of the following statements describes the particles of an ideal gas?

A) When the particles collide, energy is lost.
B) There are forces of attraction between the particles.
C) The particles move in well-defined, circular paths.
D) The volume of the particles is negligible.

58) Under which conditions of temperature and pressure does a sample of neon behave *most* like an ideal gas?

A) 400 K and 0.25 atm B) 100 K and 0.25 atm C) 100 K and 25 atm D) 400 K and 25 atm

59) Under which conditions of temperature and pressure would a real gas behave *most* like an ideal gas?

A) 200. K and 50.0 kPa
B) 200. K and 200.0 kPa
C) 600. K and 200.0 kPa
D) 600. K and 50.0 kPa

60) Under which conditions of temperature and pressure does oxygen gas behave *least* like an ideal gas?

A) high temperature and low pressure
B) high temperature and high pressure
C) low temperature and high pressure
D) low temperature and low pressure

61) Which sample at STP has the same number of molecules as 5 liters of $NO_2(g)$ at STP?

A) 5×10^{23} molecules of $CO_2(g)$
B) 5 liters of $CH_4(g)$
C) 5 moles of $O_2(g)$
D) 5 grams of $H_2(g)$

62) At STP, 1.0 liter of helium contains the same total number of atoms as

A) 1.0 L of Ne B) 0.5 L of Rn C) 1.5 L of Ar D) 2.0 L of Kr

63) Which of the following gas samples at STP has the same total number of molecules as 2.0 liters of $CO_2(g)$ at STP?

A) 5.0 L of $CO_2(g)$ B) 2.0 L of $Cl_2(g)$ C) 6.0 L of He(g) D) 3.0 L of $H_2S(g)$

64) The table below shows data for the temperature, pressure, and volume of four gas samples.

Data for Four Gas Samples

Gas Sample	Temperature (K)	Pressure (atm)	Volume (mL)
A	100.	2	400.
B	200.	2	200.
C	100.	2	400.
D	200.	4	200.

Which two gas samples have the same total number of molecules?

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

65) Which solution has the *lowest* freezing point?

A) 10. g of KI dissolved in 100. g of water
B) 40. g of KI dissolved in 200. g of water
C) 20. g of KI dissolved in 200. g of water
D) 30. g of KI dissolved in 100. g of water

66) Which aqueous solution of KI freezes at the *lowest* temperature?

A) 1 mol of KI in 500. g of water
B) 1 mol of KI in 1,000. g of water
C) 2 mol of KI in 1,000. g of water
D) 2 mol of KI in 500. g of water

67) Which sample, when dissolved in 1.0 liter of water, produces a solution with the *lowest* freezing point?

A) 0.2 mol of $CaCl_2$ B) 0.1 mol of C_2H_5OH C) 0.2 mol of $C_6H_{12}O_6$ D) 0.1 mol of LiBr

68) Compared to the freezing point and boiling point of water at 1 atmosphere, a solution of a salt and water at 1 atmosphere has a

A) lower freezing point and a higher boiling point
B) lower freezing point and a lower boiling point
C) higher freezing point and a lower boiling point
D) higher freezing point and a higher boiling point

69) 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) entropy of the system
B) activation energy
C) heat of reaction
D) heat of fusion

70) What term refers to the difference between the potential energy of the products and the potential energy of the reactants for any chemical change?

A) heat of reaction
B) heat of deposition
C) heat of vaporization
D) heat of fusion

71) The net energy released or absorbed during a reversible chemical reaction is equal to

A) the activation energy of the endothermic reaction
B) the activation energy of the exothermic reaction
C) the sum of the potential energy of the products and the potential energy of the reactants
D) the difference between the potential energy of the products and the potential energy of the reactants

72) Which of the following 1-mole samples has the *least* entropy?

A) $Br_2(\ell)$ at 332 K B) $Br_2(s)$ at 266 K C) $Br_2(\ell)$ at 266 K D) $Br_2(g)$ at 332 K

73) In terms of energy and entropy, systems in nature tend to undergo changes toward

A) lower energy and lower entropy
B) higher energy and higher entropy
C) higher energy and lower entropy
D) lower energy and higher entropy

74) Systems in nature tend to undergo changes toward

A) higher energy and more disorder
B) lower energy and more disorder
C) higher energy and less disorder
D) lower energy and less disorder

75) Given the equation representing a phase change at equilibrium:

$C_2H_5OH(\ell) \longrightarrow C_2H_5OH(g)$

Which statement is true?

A) The reverse process proceeds faster than the forward process.
B) The forward process proceeds faster than the reverse process.
C) The forward and reverse processes proceed at the same rate.
D) The forward and reverse processes both stop.

76) Which quantities must be equal for a chemical reaction at equilibrium?

A) the potential energies of the reactants and products
B) the rates of the forward and reverse reactions
C) the concentrations of the reactants and products
D) the activation energies of the forward and reverse reactions

77) Which statement must be true for any chemical reaction at equilibrium?

A) The concentration of the products is less than the concentration of the reactants.
B) The concentration of the products is greater than the concentration of the reactants.
C) The concentration of the products and the concentration of the reactants are constant.
D) The concentration of the products and the concentration of the reactants are equal.

78) The entropy of a sample of CO_2 increases as the CO_2 changes from

A) solid to gas B) gas to liquid C) liquid to solid D) gas to solid

79) The entropy of a sample of H_2O increases as the sample changes from a

A) gas to a solid B) gas to a liquid C) liquid to a gas D) liquid to a solid

80) A substance is classified as an electrolyte because

A) its aqueous solution has a pH value of 7
B) its aqueous solution conducts an electric current
C) it has a high melting point
D) it contains covalent bonds

81) Which of the following two compounds are electrolytes?

A) NaOH and HCl
B) NaOH and CH_3CH_2OH
C) $C_6H_{12}O_6$ and CH_3CH_2OH
D) $C_6H_{12}O_6$ and HCl

82) Which of the following substances is an electrolyte?

A) CCl_4 B) C_2H_6 C) H_2O D) HCl

83) 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

84) Which one of the following compounds when dissolved in water is an Arrhenius acid?

A) CH_3OH B) NaOH C) NaCl D) HCl

85) Which of the following substances is an Arrhenius acid?

A) $Ba(OH)_2$ B) H_3PO_4 C) NaCl D) CH_3COOCH_3

86) An Arrhenius base yields which ion as the only negative ion in an aqueous solution?

A) hydroxide ion B) hydronium ion C) hydride ion D) hydrogen ion

87) Which of the following substances is an Arrhenius base?

A) CH_3Cl B) CH_3OH C) LiCl D) LiOH

88) Which compound releases hydroxide ions in an aqueous solution?

A) CH_3COOH B) CH_3OH C) KOH D) HCl

89) One alternate acid-base theory states that an acid is an

A) H^+ donor B) H^+ acceptor C) OH^- acceptor D) OH^- donor

90) An acid can be defined as an

A) OH^- acceptor B) OH^- donor C) H^+ donor D) H^+ acceptor

91) One acid-base theory defines a base as an

A) H^+ acceptor B) H donor C) H^+ donor D) H acceptor

92) Given the equation representing a reaction at equilibrium:

$$NH_3(g) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

The H^+ acceptor for the forward reaction is

A) $OH^-(aq)$ B) $NH_3(g)$ C) $NH_4^+(aq)$ D) $H_2O(\ell)$

93) 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

94) Which change in pH represents a hundredfold increase in the concentration of hydronium ions in a solution?

A) pH 2 to pH 1 B) pH 1 to pH 2 C) pH 1 to pH 3 D) pH 3 to pH 1

95) Atoms of one element are converted to atoms of another element through

A) fermentation B) transmutation C) polymerization D) oxidation

96) Which reaction converts an atom of one element to an atom of another element?

A) polymerization B) transmutation C) saponification D) combustion

97) A change in the nucleus of an atom that converts the atom from one element to another element is called

A) neutralization B) polymerization C) combustion D) transmutation

98) The energy released by a nuclear reaction results primarily from the

A) formation of bonds between atoms
B) conversion of energy into mass
C) conversion of mass into energy
D) breaking of bonds between atoms

99) The amount of energy released from a fission reaction is much greater than the energy released from a chemical reaction because in a fission reaction

A) mass is converted into energy
B) ionic bonds are broken
C) covalent bonds are broken
D) energy is converted into mass

100) Which change takes place in a nuclear fusion reaction?

A) Covalent bonds are converted to ionic bonds.
B) Energy is converted to matter.
C) Matter is converted to energy.
D) Ionic bonds are converted to covalent bonds.

100 Questions That You Should Never Get Wrong
Answer Key

1. A
2. C
3. D
4. C
5. B
6. C
7. A
8. B
9. D
10. D
11. B
12. D
13. C
14. B
15. A
16. C
17. A
18. B
19. D
20. D
21. B
22. C
23. A
24. A
25. B
26. D
27. A
28. B
29. A
30. D
31. C
32. A
33. B
34. C
35. B
36. D
37. A
38. C
39. D
40. D
41. C
42. B
43. D
44. C
45. B
46. C
47. A
48. C
49. A
50. A
51. D
52. C
53. B
54. C
55. B
56. B
57. D
58. A
59. D
60. C
61. B
62. A
63. B
64. C
65. D
66. D
67. A
68. A
69. A
70. A
71. D
72. B
73. D
74. B
75. C
76. B
77. C
78. A
79. C
80. B
81. A
82. D
83. C
84. D
85. B
86. A
87. D
88. C
89. A
90. C
91. D
92. B
93. B
94. D
95. B
96. B
97. D
98. C
99. A
100. C

Name: **"State in Terms Of ..."**

____ 1. Base your answer to the following question on the information below and on your knowledge of chemistry.

Chemical concepts are applied in candy making. A recipe for making lollipops is shown below.

Hard-Candy Lollipops Recipe

Ingredients:
414 grams of sugar
177 grams of water
158 milliliters of light corn syrup

Step 1: In a saucepan, mix the sugar and water. Heat this mixture, while stirring, until all of the sugar dissolves.
Step 2: Add the corn syrup and heat the mixture until it boils.
Step 3: Continue boiling the mixture until the temperature reaches 143°C at standards pressure.
Step 4: Remove the pan from the heat and allow it to stand until the bubbling stops. Pour the mixture into lollipop molds that have been coated with cooking oil spray.

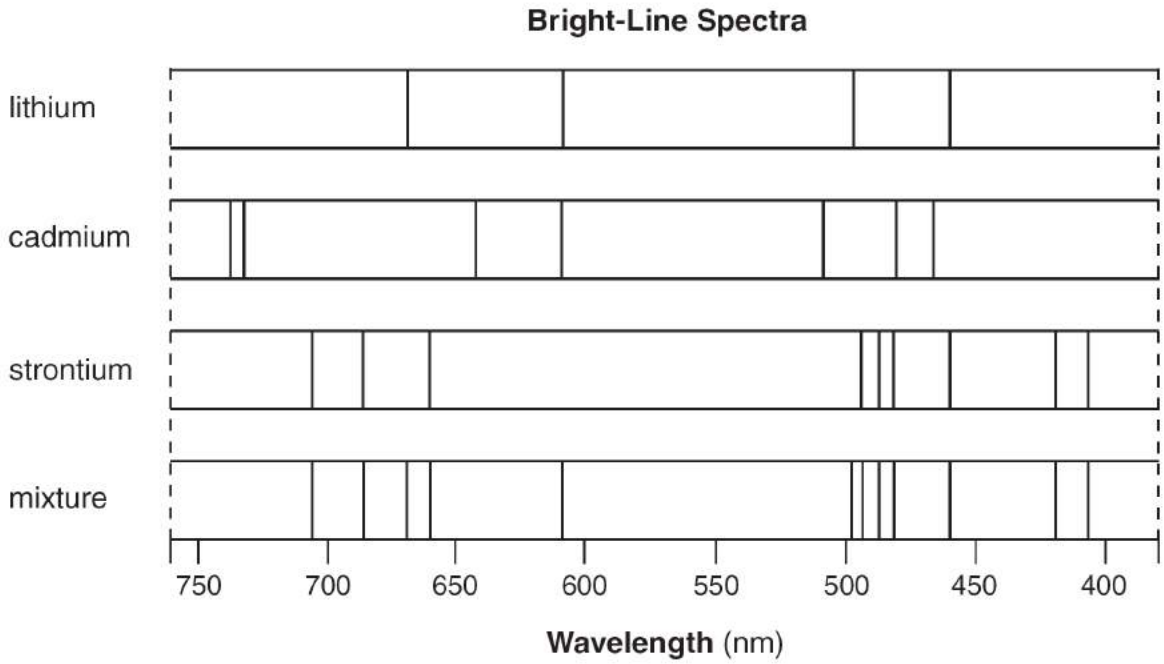
Explain, in terms of the concentration of sugar molecules, why the boiling point of the mixture in step 3 increases as water evaporates from the mixture.

____ 2. Base your answer to the following question on the information below

The element boron, a trace element in Earth's crust, is found in foods produced from plants. Boron has only two naturally occurring stable isotopes, boron-10 and boron-11.

State, in terms of subatomic particles, *one* difference between the nucleus of a carbon-11 atom and the nucleus of a boron-11 atom.

3. Base your answer to the following question on the information below.
The bright-line spectra for three elements and a mixture of elements are shown below.



Explain, in terms of *both electrons and energy, how the bright-line spectrum of an element is produced.*

4. Explain, in terms of protons and neutrons, why
U-235 and U-238 are different isotopes of uranium.

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

Before atomic numbers were known, Mendeleev developed a classification system for the 63 elements known in 1872, using oxide formulas and atomic masses. He used an R in the oxide formulas to represent any element in each group. The atomic mass was listed in parentheses after the symbol of each element. A modified version of Mendeleev's classification system is shown in the table below.

Modified Version of Mendeleev's Table

Group →		I	II	III	IV	V	VI	VII
Oxide formulas		R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7
Series	1	H(1)						
	2	Li(7)	Be(9.4)	B(11)	C(12)	N(14)	O(16)	F(19)
	3	Na(23)	Mg(24)	Al(27.3)	Si(28)	P(31)	S(32)	Cl(35.5)
	4	K(39)	Ca(40)		Ti(48)	V(51)	Cr(52)	Mn(55)
	5	Cu(63)	Zn(65)			As(75)	Se(78)	Br(80)
	6	Rb(85)	Sr(87)	Yt(88)	Zr(90)	Nb(94)	Mo(96)	
	7	Ag(108)	Cd(112)	In(113)	Sn(118)	Sb(122)	Te(125)	I(127)
	8	Cs(133)	Ba(137)	Di(138)	Ce(140)			

Explain , in terms of chemical reactivity, why the elements in Group 18 on the modern Periodic Table were *not* identified by Mendeleev at that time.

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

When magnesium is ignited in air, the magnesium reacts with oxygen and nitrogen. The reaction between magnesium and nitrogen is represented by the unbalanced equation below:

$$Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

Explain, in terms of electrons, why an atom of the metal in this reaction forms an ion that has a smaller radius than its atom.

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

The atomic radius and the ionic radius for some Group 1 and some Group 17 elements are given in the tables below.

Atomic and Ionic Radii of Some Elements

Group 1

Particle	Radius (pm)
Li atom	130.
Li^+ ion	78
Na atom	160.
Na^+ ion	98
K atom	200.
K^+ ion	133
Rb atom	215
Rb^+ ion	148

Group 17

Particle	Radius (pm)
F atom	60.
F^- ion	133
Cl atom	100.
Cl^- ion	181
Br atom	117
Br^- ion	?
I atom	136
I^- ion	220.

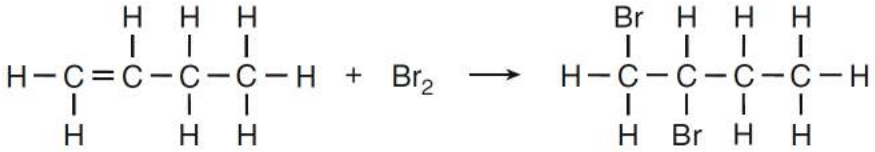
Explain, in terms of electron shells, why the radius of a K^+ ion is greater than the radius of an Na^+ ion.

8. Base your answer to the following question on the elements in Group 2 on the Periodic Table.

Explain, in terms of atomic structure, why the elements in Group 2 have similar chemical properties.

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

The equation below represents the reaction between 1-butene and bromine to form the compound 1,2-dibromobutane, $C_4H_8Br_2$.



Explain, in terms of bonding, why the hydrocarbon reactant is an unsaturated hydrocarbon.

Base your answers to questions **10** and **11** on the information below and on your knowledge of chemistry.

Rubbing alcohol is a product available at most pharmacies and supermarkets. One rubbing alcohol solution contains 2-propanol and water. The boiling point of 2-propanol is 82.3°C at standard pressure.

10. Explain, in tenns of charge distribution, why a molecule of the 2-propanol is a polar molecule.

11. Explain in terms of electronegativity differences, why a C–O bond is more polar than a C–H bond.

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

Ammonium chloride is dissolved in water to form a 0.10 M $NH_4Cl(aq)$ solution. This dissolving process is represented by the equation below.

$$NH_4Cl(s) + \text{heat} \xrightarrow{H_2O} NH_4{}^+(aq) + Cl^-(aq)$$

Explain, in terms of ions, why a 10.0-milliliter sample of 0.30 M $NH_4Cl(aq)$ is a better conductor of electricity than a 10.0-milliliter sample of the 0.10 M $NH_4Cl(aq)$.

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

Ozone, $O_3(g)$, is produced from oxygen, $O_2(g)$ by electrical discharge during thunderstorms. The unbalanced equation below represents the reaction that forms ozone.

$$O_2(g) \xrightarrow{\text{electricity}} O_3(g)$$

Explain, in terms of electron configuration, why an oxygen molecule is more stable than an oxygen atom.

14. Explain, in terms of valence electrons, why the bonding in magnesium oxide, MgO, is similar to the bonding in barium chloride, $BaCl_2$.

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

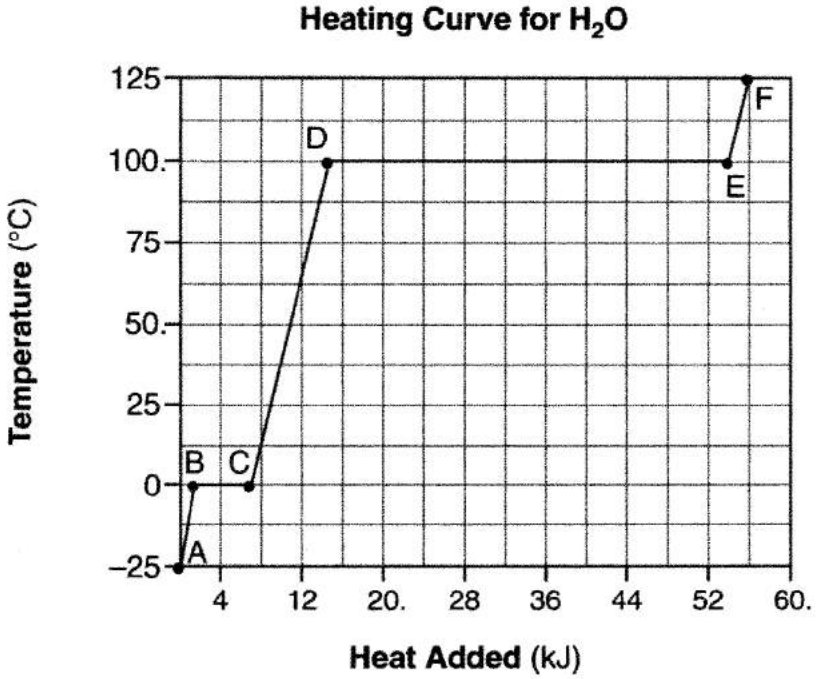
Carbon has three naturally occurring isotopes, C-12, C-13, and C-14. Diamond and graphite are familiar forms of solid carbon. Diamond is one of the hardest substances known, while graphite is a very soft substance. Diamond has a rigid network of bonded atoms. Graphite has atoms bonded in thin layers that are held together by weak forces.

Recent experiments have produced new forms of solid carbon called fullerenes. One fullerene, C_{60}, is a spherical, cagelike molecule of carbon.

State, in terms of the arrangement of atoms, the difference in hardness between diamond and graphite.

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

Starting as a solid at –25°C, a sample of H_2O is heated at a constant rate until the sample is at 125°C. This heating occurs at standard pressure. The graph below represents the relationship between temperature and heat added to the sample.



Explain, in terms of heat of fusion and heat of vaporization, why the heat added during interval *DE* is greater than the heat added during interval *BC* for this sample of water.

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

Nitrogen gas and oxygen gas make up about 99% of Earth's atmosphere. Other atmospheric gases include argon, carbon dioxide, methane, ozone, hydrogen, etc.
The amount of carbon dioxide in the atmosphere can vary. Data for the concentration of CO2(g) from 1960 to 2000 are shown in the table below.

Atmospheric Concentration of $CO_2(g)$

Year	Concentration (ppm)
1960	316.9
1980	338.7
2000	369.4

Explain, in terms of types of matter, why methane can be broken down by chemical means, but argon can *not* be broken down by chemical means. Your response must include *both methane and argon.*

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

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. ºC. This process is represented by the balanced equation below.

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

Explain, in terms of the distribution of particles, why the solution is a homogeneous mixture.

Base your answers to questions **19** and **20** on the information below.

Some Properties of Three Compounds at Standard Pressure

Compound	Boiling Point (°C)	Solubility in 100. Grams of H_2O at 20.°C (g)
ammonia	−33.2	56
methane	−161.5	0.002
hydrogen chloride	−84.9	72

19. Explain, in terms of intermolecular forces, why ammonia has a higher boiling point than the other compounds in the table.

20. Explain, in terms of molecular polarity, why hydrogen chloride is more soluble than methane in water at 20.ºC and standard pressure.

21. Base your answer to the following question 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.

Classify, in terms of saturation, the type of solution made by the student.

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

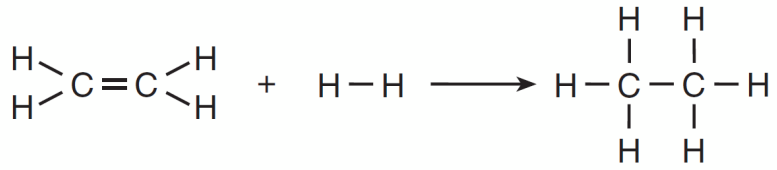
Molar Mass and Boiling Point of Four Substances

Substance	Molar Mass (g/mol)	Boiling Point at 1 atm (K)
methane	16	112
ethane	30.	185
propane	44	231
butane	58	273

State, in terms of intermolecular forces, why the boiling point of propane at 1 atmosphere is *lower* than the boiling point of butane at 1 atmosphere.

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

Ethene and hydrogen can react at a faster rate in the presence of the catalyst platinum. The equation below represents a reaction between ethene and hydrogen.



Explain, in terms of activation energy, why the catalyzed reaction occurs at a faster rate.

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

At standard pressure, hydrogen peroxide, H_2O_2, melts at –0.4°C, boils at 151°C, and is very soluble in water. A bottle of aqueous hydrogen peroxide, $H_2O_2(aq)$, purchased from a pharmacy has a pressure-releasing cap. Aqueous hydrogen peroxide decomposes at room temperature, as represented by the balanced equation below.

$$2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g) + 196.0\ kJ$$

State, in terms of *both* melting point and boiling point, why H_2O_2 is a liquid at room temperature.

Base your answers to questions **25** and **26** on the information below.

Several steps are involved in the industrial production of sulfuric acid. One step involves the oxidation of sulfur dioxide gas to form sulfur trioxide gas. A catalyst is used to increase the rate of production of sulfur trioxide gas. In a rigid cylinder with a movable piston, this reaction reaches equilibrium, as represented by the equation below.

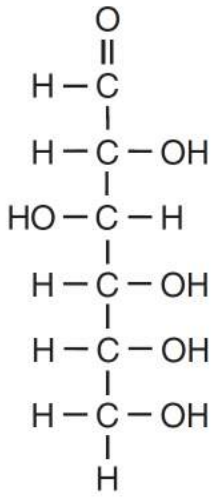
$$2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g) + 392\ kJ$$

25. State, in terms of the concentration of $SO_3(g)$, what occurs when more $O_2(g)$ is added to the reaction at equilibrium.

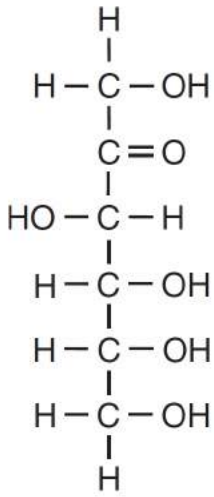
26. Explain, in terms of collision theory, why increasing the pressure of the gases in the cylinder increases the rate of the forward reaction.

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

Table sugar, sucrose, is a combination of two simple sugars, glucose and fructose. The formulas below represent these simple sugars.



Glucose

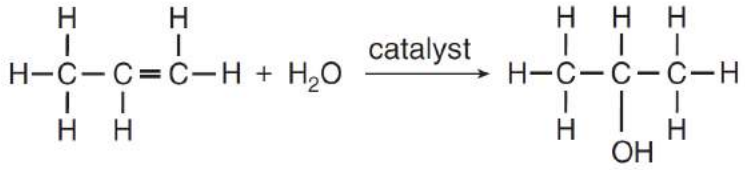


Fructose

Explain, in terms of atoms and molecular structure, why glucose and fructose are isomers of each other.

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

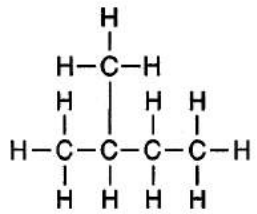
In one industrial organic reaction, C_3H_6 reacts with water in the presence of a catalyst. This reaction is represented by the balanced equation below.



Explain, in terms of bonding, why C_3H_6 is classified as an unsaturated hydrocarbon.

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

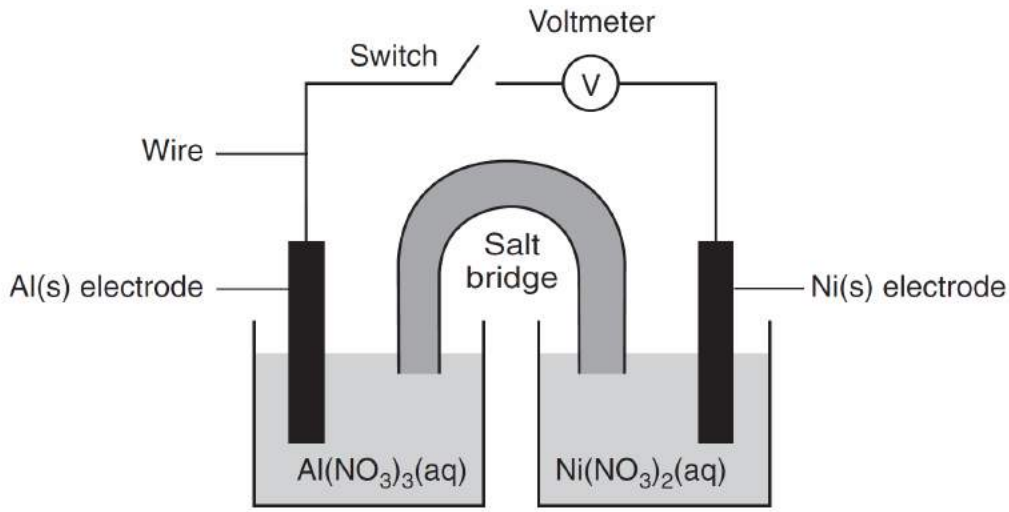
The formula below represents a hydrocarbon.



Explain, in terms of carbon-carbon bonds, why this hydrocarbon is saturated.

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

A student constructs an electrochemical cell during a laboratory investigation. When the switch is closed, electrons flow through the external circuit. The diagram and equation below represent this cell and the reaction that occurs.

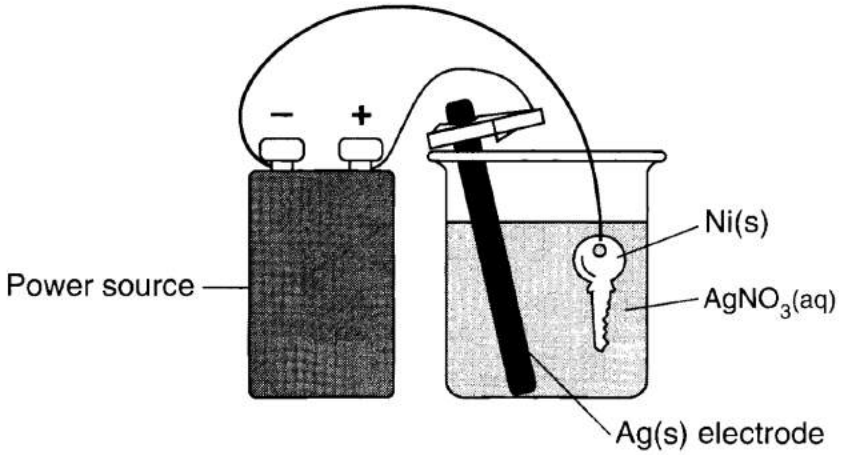


$$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$$

State, in terms of energy, why this cell is a voltaic cell.

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

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.



Explain, in terms of Ag atoms and $Ag^+(aq)$ ions, why the mass of the silver electrode *decreases* as the cell operates.

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

A laboratory worker filled a bottle with a hydrochloric acid solution. Another bottle was filled with methanol, while a third bottle was filled with a sodium hydroxide solution. However, the worker neglected to label each bottle. After a few days, the worker could not remember which liquid was in each bottle.

The worker needed to identify the liquid in each bottle. The bottles were labeled *A*, *B*, and *C*. Using materials found in the lab (indicators, conductivity apparatus, and pieces of Mg metal), the worker tested samples of liquid from each bottle. The test results are shown in the table below.

Table of Tests and Results

Test	Test Results		
	Bottle A	Bottle B	Bottle C
methyl orange indicator	yellow	yellow	yellow
bromthymol blue indicator	blue	green	yellow
electrical conductivity	conductor	nonconductor	conductor
reactivity with Mg metal	no reaction	no reaction	reaction

Explain, in terms of pH, why the methyl orange indicator test results were the same for each of the three liquids.

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

Iodine has many isotopes, but only iodine-127 is stable and is found in nature. One radioactive iodine isotope, I-108, decays by alpha particle emission. Iodine-131 is also radioactive and has many important medical uses.

Explain, in terms of protons and neutrons, why I-127 and I-131 are different isotopes of iodine.

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

The radioisotope uranium-238 occurs naturally in Earth's crust. The disintegration of this radioisotope is the first in a series of spontaneous decays.

The sixth decay in this series produces the radioisotope radon-222. The decay of radon-222 produces the radioisotope polonium-218 that has a half life of 3.04 minutes. Eventually, the stable isotope lead-206 is produced by the alpha decay of an unstable nuclide.

Explain, in terms of electron configuration, why atoms of the radioisotope produced by the sixth decay in the U-238 disintegration series do not readily react to form compounds.

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

Scientists are investigating the production of energy using hydrogen-2 nuclei (deuterons) and hydrogen-3 nuclei (tritons). The balanced equation below represents one nuclear reaction between two deuterons.

$$^{2}_{1}H + {}^{2}_{1}H \rightarrow {}^{3}_{2}He + {}^{1}_{0}n + 5.23 \times 10^{-13}\,J$$

State, in terms of subatomic particles, how a deuteron differs from a triton.

Answer Key

"State in Terms Of ..."

1. –The boiling point of the mixture increases as water evaporates because the concentration of dissolved molecules increases. –An increase in the concentration of sugar particles increases the boiling point.
2. –The carbon-11 nucleus has one more proton than the nucleus of boron-11. –A B-11 atom has a different number of neutrons than a C-11 atom.
3. – When electrons in an excited state return to a lower energy state, specific amounts of energy are emitted. These energies are associated with specific wavelengths of light that are characteristic of the bright-line spectrum of an element. – Energy is emitted when excited electrons fall back to lower shells.
4. – A U-235 atom has 92 protons and 143 neutrons, and a U-238 atom has 92 protons and 146 neutrons. – A U-235 atom and a U-238 atom have the same number of protons but a different number of neutrons.
5. –Since the Group 18 elements tend not to react with other elements, there were no oxide compounds for Mendeleev to study. –Group 18 elements are generally unreactive.
6. – An atom of magnesium loses its outer shell electrons to form the Mg^{2+} ion. – The electron configuration of a magnesium atom is 2-8-2, and the electron configuration of the magnesium ion is 2-8. – An atom of the metal loses electrons to form the ion.
7. —A K^+ ion has three electron shells and an Na^+ ion has only two. —A sodium ion has fewer electron shells than a potassium ion.
8. –In the ground state, an atom of each element has two valence electrons. –The number of electrons in the outermost shell of each atom is the same.
9. –Each reactant hydrocarbon molecule has a double carbon-carbon bond. –There is a multiple carbon-carbon bond in each molecule. –More hydrogen atoms can be bonded with this hydrocarbon.
10. – A 2-proponal molecule is polar because it has an asymmetrical distribution of charge. – The charge distribution is uneven. – The center of positive charge and the center of negative charge do *not* coincide
11. – There is a greater electronegativity difference in a CO bond than in a CH bond. – The CO bond is more polar because the electronegativity difference for a CO bond is 0.8, and the electronegativity difference for a CH bond is 0.4. – The CH bond has a smaller difference. – The CO is .8 and the CH is .4
12. –The 0.30 M NH_4 Cl(aq) sample has more mobile ions in solution. –The 0.10 M NH_4Cl solution has a lower concentration of ions.
13. – Both atoms in an O2 molecule have achieved a noble gas electron configuration. – An oxygen atom does not have a stable octet of valence electrons.
14. –The bonding in each compound involves a transfer of valence electrons from the metal to the nonmetal. –Both metals lose all of their valence electrons.
15. Diamond has atoms bonded strongly in a three-dimensional network. Graphite has atoms that are held weakly between layers.
16. –The heat of vaporization of water is 2260 J/g and the heat of fusion for water is only 334 J/g. –The heat of fusion of water is much less than its heat of vaporization.

Answer Key

"State in Terms Of ..."

17. –Methane is a compound consisting of two elements, so it can be broken down by chemical means, but argon is an element, which cannot be broken down. –Methane is a compound and argon is an element.

18. —The H^+ ions and the Cl^- ions are distributed uniformly throughout the solution. — There is an even distribution of $H^+(aq)$ and $Cl^-(aq)$.

19. –Ammonia has stronger intermolecular forces than either methane or hydrogen chloride. –Ammonia has hydrogen bonding

20. –Molecules of CH_4 are nonpolar, but molecules of HCl and H_2O are both polar. –Hydrogen chloride and water are both polar.

21. The solution made by the student is unsaturated.

22. The boiling point of propane at 1 atm is lower than the boiling point of butane at 1 atm because propane has weaker intermolecular forces than butane; Butane has stronger intermolecular forces.

23. – The catalyzed reaction pathway has a lower activation energy than the original reaction. – Less energy is needed.

24. —Room temperature is above the melting point and below the boiling point of H_2O_2 . —Room temperature is between –0.4°C and 151°C. — –0.4°C < room temperature < 151°C

25. The concentration of $SO_3(g)$ increases.

26. When the pressure in the cylinder is increased, the $SO_2(g)$ molecules and $O_2(g)$ molecules collide more frequently, producing more $SO_3(g)$.

27. – The number of each kind of atom is the same in both, but their structures are not he same. – Their molecular formulas are the same, but their structural arrangement of atoms is different. – same molecular formula but different structural formulas – The only different is the arrangement of the atoms.

28. Acceptable responses include, but are not limited to: The C_3H_6 is unsaturated because each molecule has a double covalent bond between two of its carbon atoms. There is a carbon-carbon double bond in each molecule

29. • All of the carbon-carbon bonds are single covalent bonds. • There are only single bonds between the carbon atoms.

30. —A spontaneous reaction converts chemical energy to electrical energy. —A battery is not required to provide energy for the cell to operate.

31. —Silver atoms lose electrons and become silver ions in the solution. — Some of the Ag atoms become Ag^+ ions. —Silver atoms are oxidized to silver ions.

32. *Examples*: -All three solutions have a pH greater than 4.4. -Methyl orange changes to yellow at a pH of 4.4, which is still in the acid range. -A solution with a pH greater than 4.4 could be acidic, basic, or neutral.

33. – I-127 atoms and I-131 atoms have the same number of protons, but different numbers of neutrons. – Both have 53 p, but I-127 has 74 n while I-131 has 78 n. – They have the same atomic number but different mass numbers. – same atomic number but different numbers of neutrons – The only difference is the number of neutrons.

34. –Radon-222 atoms have a complete outer shell of electrons and tend not to bond. –There are eight valence electrons in a radon atom. –Octet in valence shell.

Answer Key

"State in Terms Of ..."

35. Examples: A deuteron has one neutron and a triton has two neutrons.; A deuteron has one fewer neutron than a triton.