ATOMIC

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**

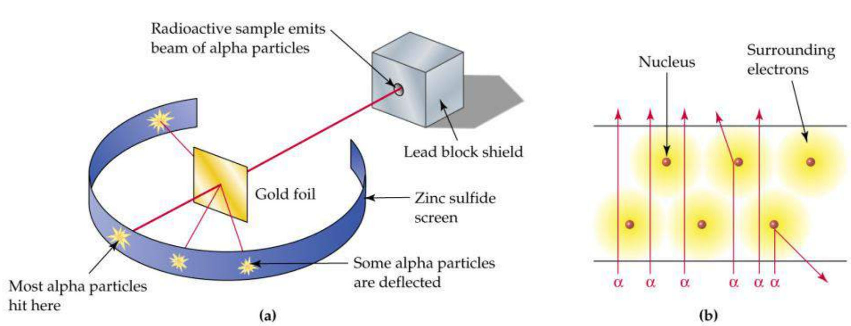
ATOMIC MODELS: helps scientists imagine on a macroscopic level what happens microscopically. Three contrasting models include:

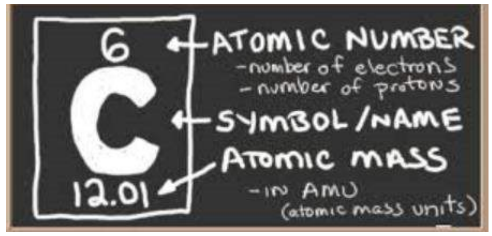
1. **Rutherford's Model:** Most of the mass of the atom is in the center (nucleus) is positive. Protons are in the nucleus, most of the atom is empty space. Electrons go around the nucleus
2. **Bohr's Model:** Protons are in the nucleus, which is positive. Electrons revolve around the nucleus in concentric 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 × 0.9312 = 36.28 amu

Potassium-41 40.962 amu × 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)

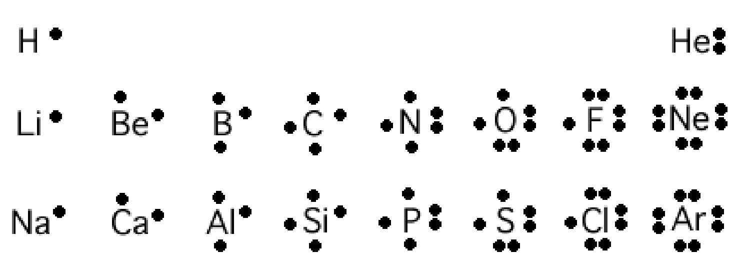
Metals

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

Non-Metals

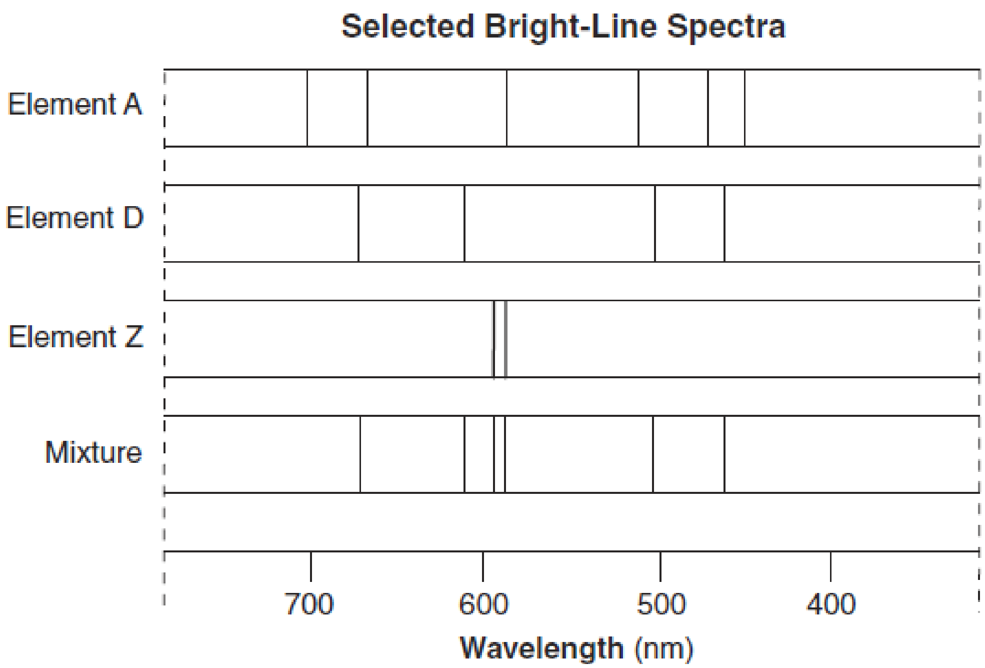
Atoms	Number of Protons = Number of Electrons	Ion	# Electons
F	9	F^{-}	10
S	16	S^{2-}	18

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



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$1p –$_1^1H$ –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. $^{11}_{5}B$

11. –Boron-11 is about four times more abundant than boron-10. –The B-10 is less abundant.

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 **<u>about 1,</u>** 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.

<u>RADIOACTIVITY:</u> 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.**

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
hydrogen-1 (protium)	hydrogen-2 (deuterium)	hydrogen-3 (tritium)

- **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 **<u>Natural Transmutation</u>**, 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

- **<u>Artificial Transmutation</u>:** elements can be made radioactive by <u>bombarding</u> 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 **<u>not</u>** spontaneously. Remember that the atomic # and the mass # have to equal the same thing on both sides.

Table O:

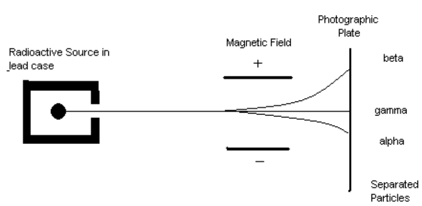
Table O
Symbols Used in Nuclear Chemistry

	Name	Notation	Symbol
Weakest penetrating power-HEAVIEST	alpha particle	$^{4}_{2}He$ or $^{4}_{2}\alpha$	α
	beta particle	$^{0}_{-1}e$ or $^{0}_{-1}\beta$	β^-
Strongest penetrating power-LIGHTEST	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^+

- 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.

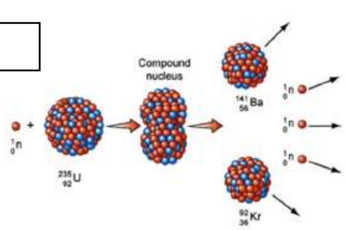


<u>Nuclear Energy:</u> in a nuclear reaction, mass is converted into energy. Two types are fission and fusion.

- **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).

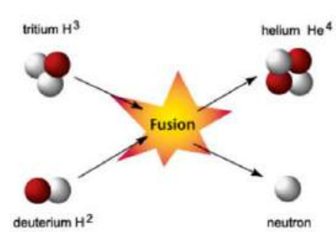
- **Fusion:** two nuclei unite to form a heavier nucleus ("u"- unit). High temperature and pressure are needed, which is why this occurs on the SUN. **<u>Fusion creates more energy than fission.</u>**

Fission



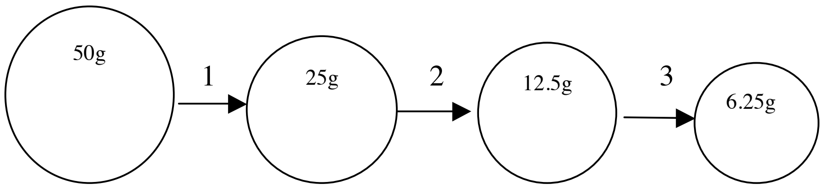
$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{93}_{36}Kr + ^{140}_{56}Ba + 3^{1}_{0}n$$

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.



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^{60}Co	5.271 y	β^-	cobalt-60

QUESTIONS:

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: **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

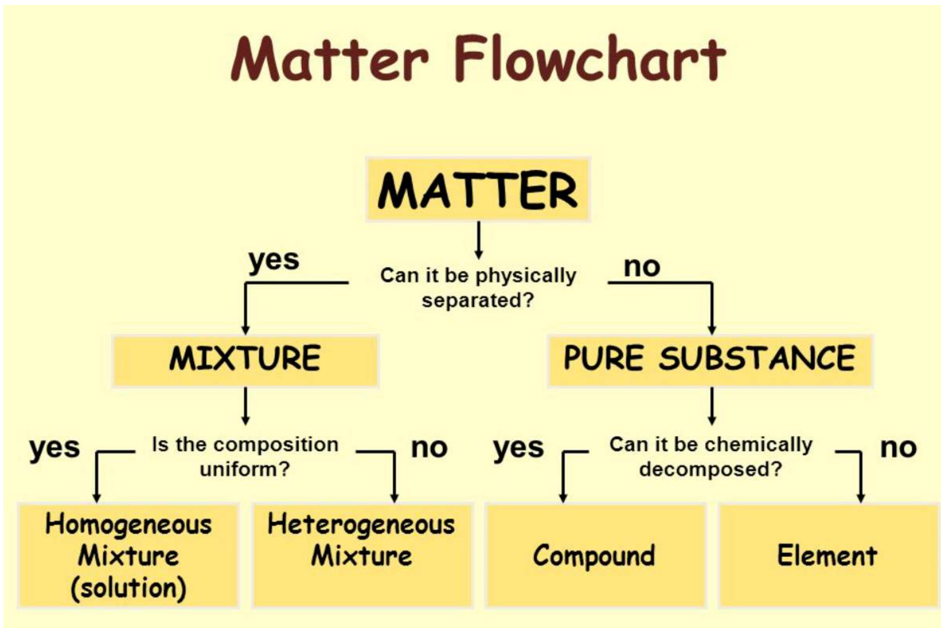
Nuclear Part 2

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.

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:

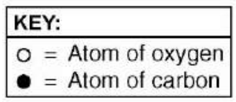
1. **Magnet:** ex. Sulfur and iron – separate iron out with the magnet
2. **Distillation**: Separate by boiling – salt water. Water will evaporate and salt will be left
3. **Filtration**: use filter, liquid goes through and solids stay behind on the filter paper.

III. States of Matter:

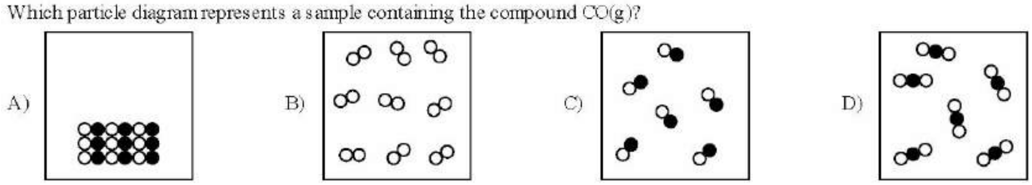
- Solids, liquids, gases (must know the general definition of each)
- MUST KNOW how to draw and interpret particle diagrams

Example:

Given the key:



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



- 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:

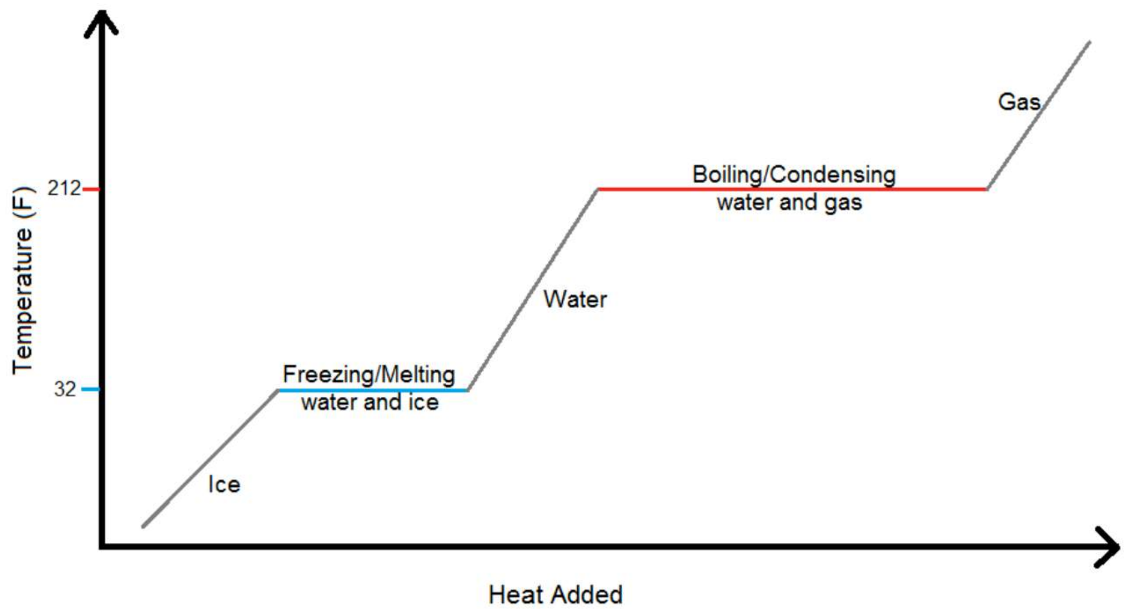
- a measure of the **average kinetic energy** of molecules. The higher the temperature, the more the kinetic energy; the lower the temperature, the lower the kinetic energy.

- **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.

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 $\mathrm{NH_4Cl}$, sand, and $\mathrm{H_2O}$ 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 $-78°C$ are placed in a flask that contains air at $21°C$. 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.

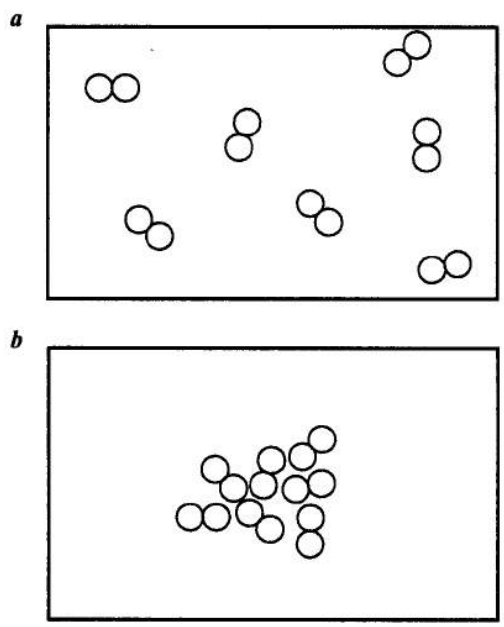
10. Compare the average kinetic energy of the water molecules in the bottle at 7 a.m. to the average kinetic energy of the water molecules in the bottle at 3 p.m.

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.

GASES/GAS LAWS

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

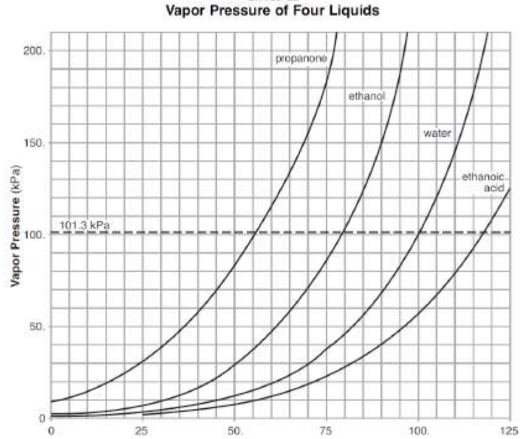
Table A
Standard Temperature and Pressure

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

Table H
Vapor Pressure of Four Liquids



<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!!!!!**

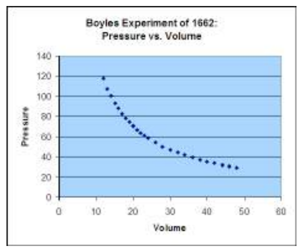
Combined Gas Law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	P = pressure V = volume T = temperature

Temperature MUST be in Kelvin

<u>BOYLE'S LAW:</u> 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.

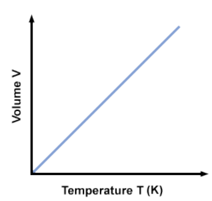
2 times as much pressure = ½ volume of gas
3 times as much pressure = 1/3 volume of gas

$$P_1V_1 = P_2V_2$$



<u>CHARLES'S LAW:</u> At constant pressure, volume is directly proportional to Kelvin (absolute) temperature. Temperature **<u>must be in degrees Kelvin</u>**, 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 <u>COMBINED GAS LAW</u> 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.

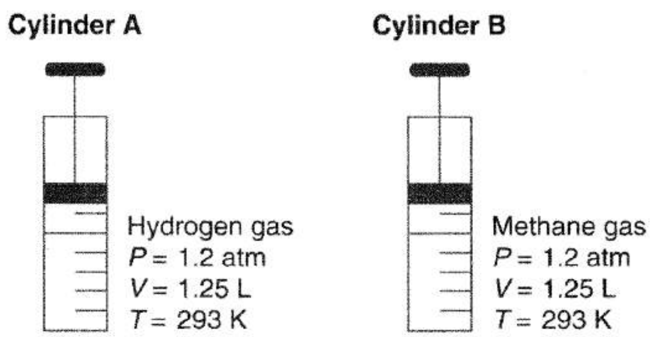
Ethane, C_2H_6, has a boiling point of $-89°C$ at standard pressure. Ethanol, C_2H_5OH, has a much higher boiling point than ethane at standard pressure. At STP, ethane is a gas and ethanol is a liquid.

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. $27°\text{C}$
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
   $V_2 = \frac{(273\text{ K})(1.30\text{ atm})(40.0\text{ L})}{(301\text{ K})(1.00\text{ atm})}$
   $\frac{(273)(1.30)(40.0)}{(301)(1.00)}$
10. $\underline{2}\ NaN_3(s) \rightarrow \underline{2}\ Na(s) + \underline{3}\ N_2(g)$

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.

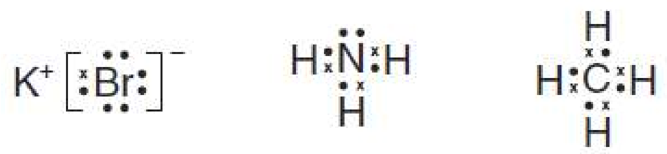


Diagram 1 Diagram 2 Diagram 3

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 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 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.

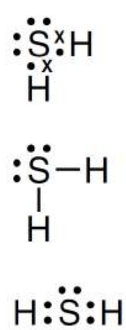
12. Base your answer to the following question on the information below and on your knowledge of chemistry.

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. 

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 $C_6H_{12}O_6$.**

<u>NAMING TYPE 1 COMPOUNDS</u>: 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

<u>NAMING TYPE 2 COMPOUNDS</u>: (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.

<u>NAMING TYPE 3 COMPOUNDS</u>: 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

<u>CLASSIFYING CHEMICAL REACTIONS</u>:

- 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)

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

- **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)

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

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.

<u>ALKANE:</u> (-ane)
- Hydrocarbons are made up of only **single bonds**, and are therefore considered **saturated.**

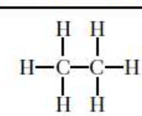
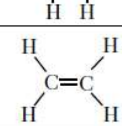
<u>ALKENE:</u> (-ene)
- Alkenes have one **double bond** and classified as **unsaturated.**

<u>ALKYNE</u> **:** (-yne)
- Alkynes have one **triple bond** and are considered **unsaturated.**

Table P
Organic Prefixes

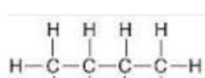
Prefix	Number of Carbon Atoms
meth-	1
eth-	2
prop-	3
but-	4
pent-	5
hex-	6
hept-	7
oct-	8
non-	9
dec-	10

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	H—C≡C—H

Note: n = number of carbon atoms

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



$CH_3CH_2CH_2CH_3$

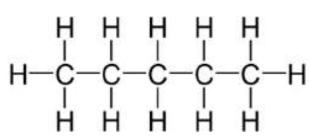
Alkyl Group: (-yl)

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

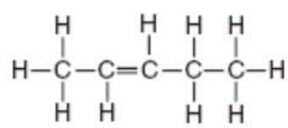
NUMBER OF CARBON ATOMS	STEM NAME	SIDE CHAIN (ALKYL GROUP) NAME	
1	meth	methyl	CH_3
2	eth	ethyl	C_2H_5
3	prop	propyl	C_3H_7
4	but	butyl	C_4H_9
5	pent	pentyl	C_5H_{11}
6	hex	hexyl	C_6H_{13}
7	hep	heptyl	C_7H_{15}
8	oct	octyl	C_8H_{17}
9	non	nonyl	C_9H_{19}
10	dec	decyl	$C_{10}H_{21}$

SUMMARY ON HOW TO DRAW CARBON COMPOUNDS:

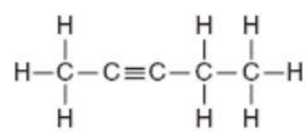
1. Look at the prefix (table P), which will tell you the number of carbons.
2. Put the bonds between the carbons (-ane = single bonds; -ene = double bond; and –yne means a triple bond)
3. The number before –ene and –yne tells you where the double or triple bond is after that carbon atom, Ex: 2-pentene (db after 2^{nd} carbon), or 1-butyne (tb after 1^{st} 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 2^{nd} carbon).
5. Put all of the appropriate hydrogen atoms around the carbon atoms if they belong there.



Pentane



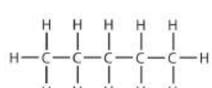
2-Pentene



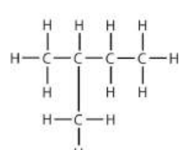
3-Pentyne

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

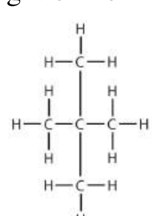
- The rule to naming these structures is to count the number of carbon atoms in the longest unbroken chain. You want to use the lowest possible number, so this means you might have to count from the left to right or from right to left.



Pentane



2-methyl butane



2,2 dimethyl propane

<u>OTHER ORGANIC COMPOUNDS & FUNCTIONAL GROUPS</u> (Table R):

<u>Alcohols:</u> (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)

<u>Ether:</u>

- 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.

<u>Aldehyde:</u> (CHO)

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

<u>Ketone:</u>

- 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.

<u>Organic Acids:</u> (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.

<u>Halides:</u> (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.

<u>Amine:</u>

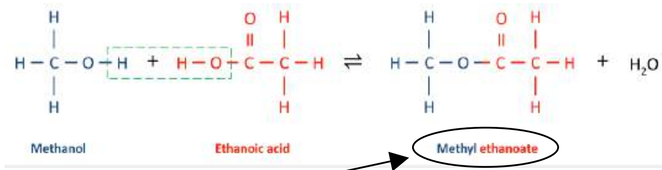
- Named by dropping the final –e in the hydrocarbon and adding **–amine.**
- Can be placed on any carbon, so a number is required.

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	$-\overset{\overset{O}{\Vert}}{C}-H$	$R-\overset{\overset{O}{\Vert}}{C}-H$	$CH_3CH_2\overset{\overset{O}{\Vert}}{C}-H$ propanal
ketone	$-\overset{\overset{O}{\Vert}}{C}-$	$R-\overset{\overset{O}{\Vert}}{C}-R'$	$CH_3\overset{\overset{O}{\Vert}}{C}CH_2CH_2CH_3$ 2-pentanone
organic acid	$-\overset{\overset{O}{\Vert}}{C}-OH$	$R-\overset{\overset{O}{\Vert}}{C}-OH$	$CH_3CH_2\overset{\overset{O}{\Vert}}{C}-OH$ propanoic acid
ester	$-\overset{\overset{O}{\Vert}}{C}-O-$	$R-\overset{\overset{O}{\Vert}}{C}-O-R'$	$CH_3CH_2\overset{\overset{O}{\Vert}}{C}OCH_3$ methyl propanoate
amine	$-\overset{\vert}{N}-$	$R-\overset{\overset{R'}{\vert}}{N}-R''$	$CH_3CH_2CH_2NH_2$ 1-propanamine
amide	$-\overset{\overset{O}{\Vert}}{C}-\overset{\vert}{N}H$	$R-\overset{\overset{O}{\Vert}}{C}-\overset{\overset{R'}{\vert}}{N}H$	$CH_3CH_2\overset{\overset{O}{\Vert}}{C}-NH_2$ propanamide

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

<u>Effects of Functional Groups on Boiling Point</u>

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)

<u>ORGANIC REACTIONS:</u>

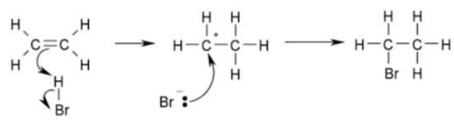
Substitution: Occurs in **ALKANES** only.

- 2 reactants and 2 products.

$$CH_4 + Cl_2 \xrightarrow{\text{heat or light}} CH_3Cl + HCl$$

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

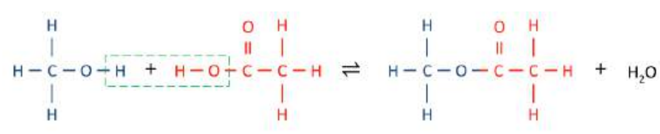
- 2 reactants and 1 product.



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

$$\underset{\text{Glucose}}{C_6H_{12}O_{6\,(aq)}} \xrightarrow{\text{yeast}} \underset{\text{Ethanol}}{2CH_3CH_2OH_{(aq)}} + \underset{\text{Carbon dioxide}}{2CO_{2\,(g)}}$$

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.

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.

---

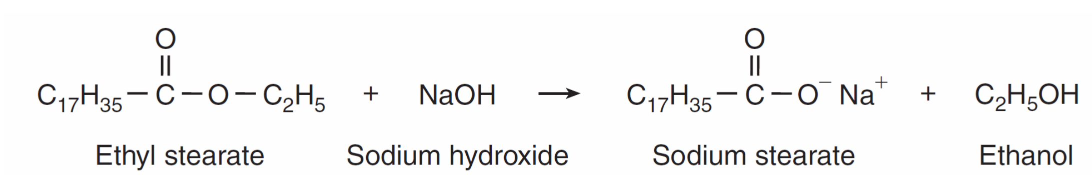
3. Base your answer to the following question on the information below and on your knowledge of chemistry.

   Ethane, C_2H_6, has a boiling point of $-89°C$ 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.

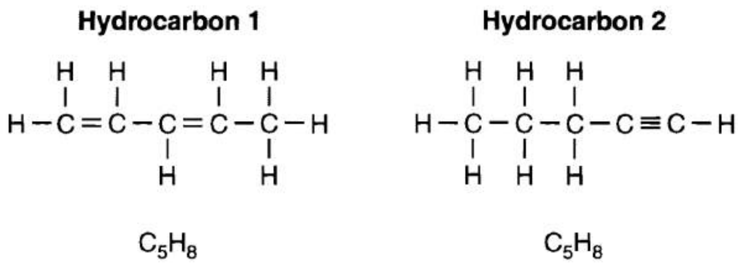


4. To which class of organic compounds does ethyl stearate belong?

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

Organic Review

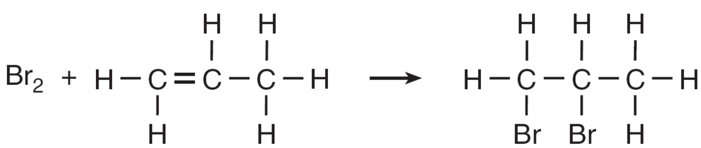
6. Two hydrocarbons that are isomers of each other are represented by the structural formulas and molecular formulas below.



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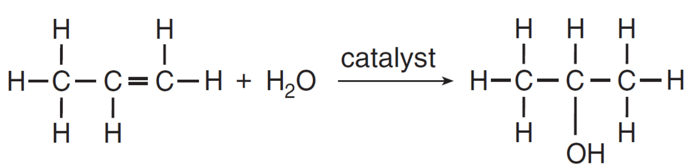
9. Base your answer to the following question on the information below.

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.

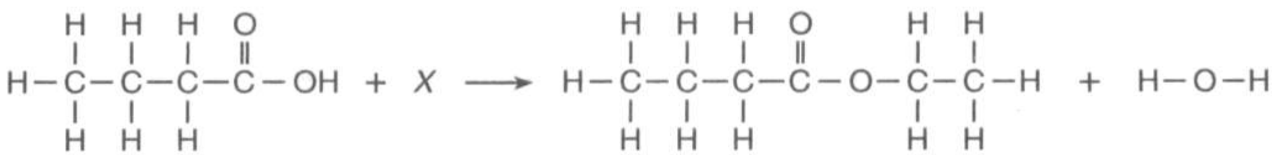
An alternative fuel, E-85, can be used in some automobiles. This fuel is a mixture of 85% ethanol and 15% gasoline.

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

Organic Review

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.

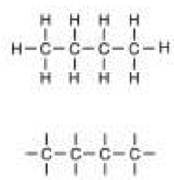
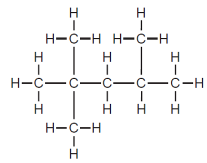
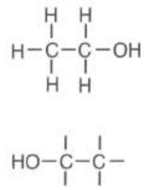
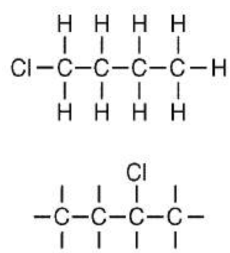
Given the balanced equation for an organic reaction between butane and chlorine that takes place at 300.ºC and 101.3 kilopascals:

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

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

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 (-)

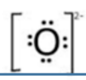
- If electronegativity difference is >1.7, then it is an IONIC BOND.

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.

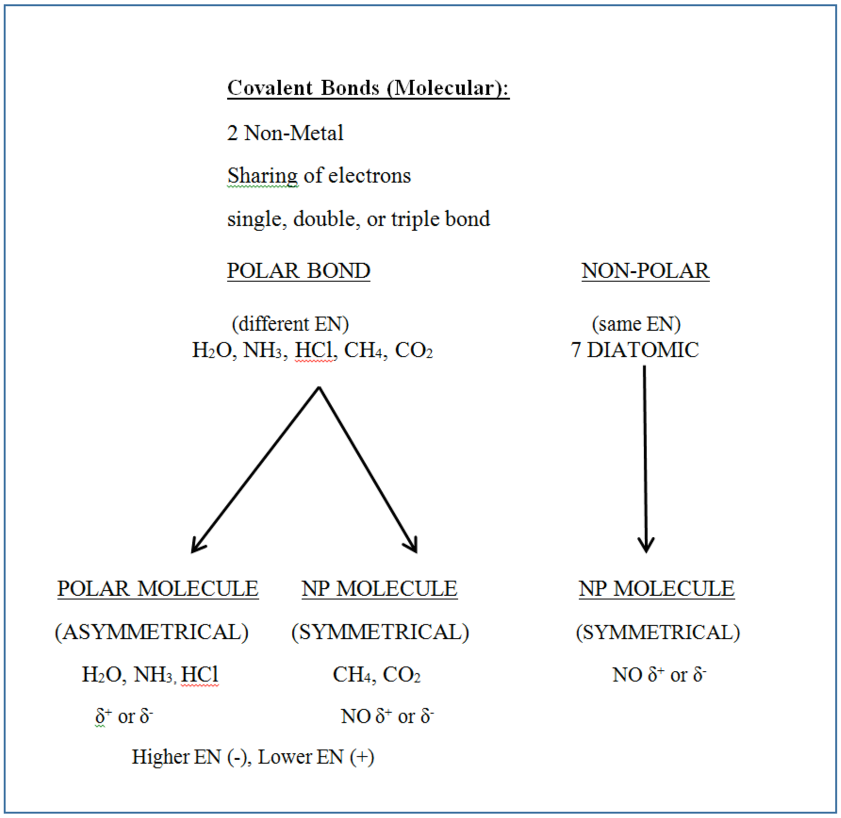
Na_3N

$Zn(OH)_2$

$(NH_4)_2O$

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

- Also called **molecular bonds.**

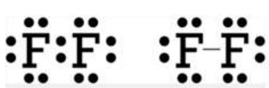


PROPERTIES OF COVALENT BONDS:

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

H_2

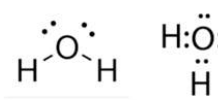
F_2, Cl_2, Br_2, I_2



O_2

N_2

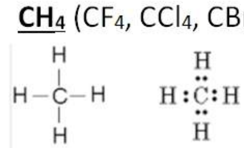
H_2O (H_2S)



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

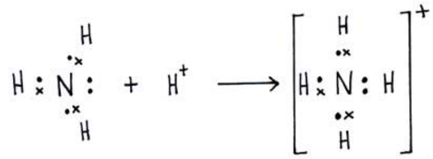
HCl (HF, HBr, HI)

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:

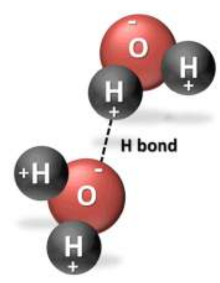
1. Hard
2. Poor conductors of heat and electricity
3. High melting points

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

<u>INTERMOLECULAR FORCES</u>: forces of attraction between molecules

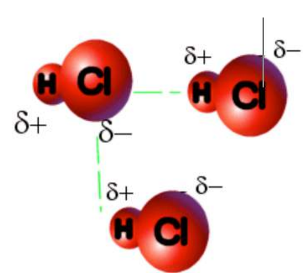
<u>HYDROGEN BONDING:</u>

- 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.



<u>DIPOLE ATTRACTION:</u>

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

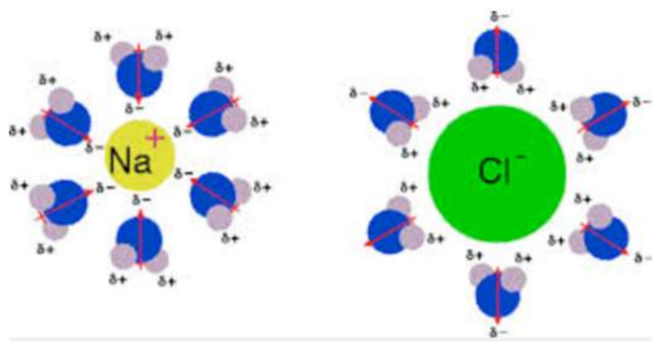


DISPERSION FORCES:

- 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

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

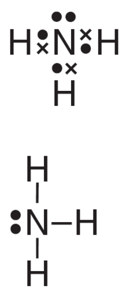
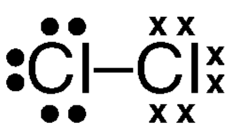
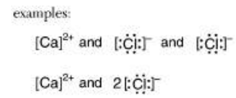
Compound	Melting Point (°C)	Boiling Point (°C)	Solubility in Water at 20.0°C
CF_4	–183.6	–127.8	insoluble
NH_3	–77.7	–33.3	soluble

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

12Explain, 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_4 Cl.*
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. 

MOLES/STOICHIOMETRY

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

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

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

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

<u>Mole – Mole:</u>

- 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:

<u>Percent Composition</u> (Table T)

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

<u>Hydrate:</u> 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: $CaSO_4{\bullet}2H_2O$. 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.**

<u>% Composition of a Hydrate from a lab:</u>

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

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.

Chemical Calculations Review

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.

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10. A hydrated compound contains water molecules within its crystal structure. The percent composition by mass of water in the hydrated compound $CaSO_4•2H_2O$ has an accepted value of 20.9%. A student did an experiment and determined that the percent composition by mass of water in $CaSO_4•2H_2O$ 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. $\frac{28.0855\ \mathrm{g}}{28.0855\ \mathrm{g}+2(15.9994\ \mathrm{g})} \times 100$

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

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

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

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

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.

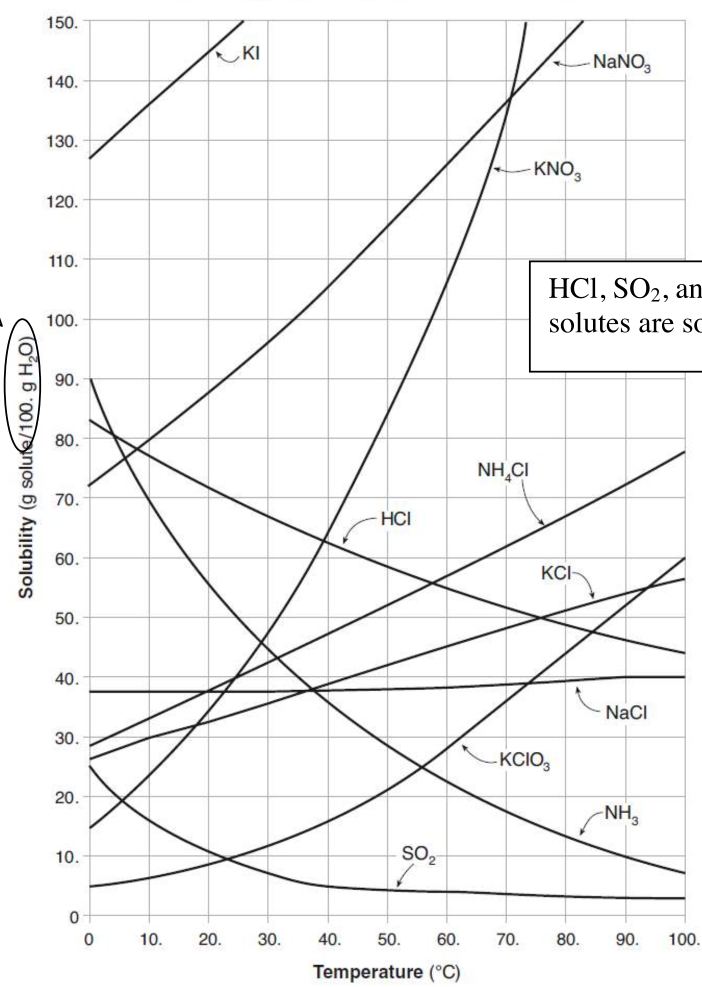
13. a) 186 grams b) .095 moles c) $C_2H_5O_4$

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.

Table G
Solubility Curves at Standard Pressure

100 grams of H_2O



HCl, SO_2, and NH_3 are gases. The rest of the solutes are solids.

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

<u>Soluble</u>	<u>Insoluble</u>
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.

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. $23°C$ to $26°C$
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

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)

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

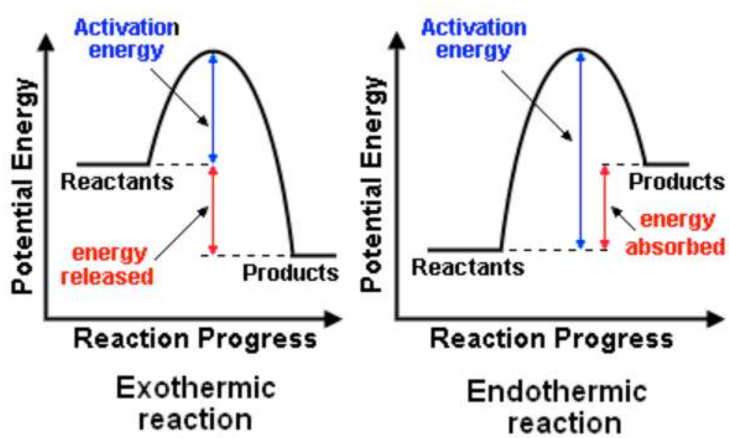
- ΔH = Heat of Reaction (Enthalpy)
- If ΔH is (+) then it is an endothermic reaction
- If ΔH is (-) then it is an exothermic reaction

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.



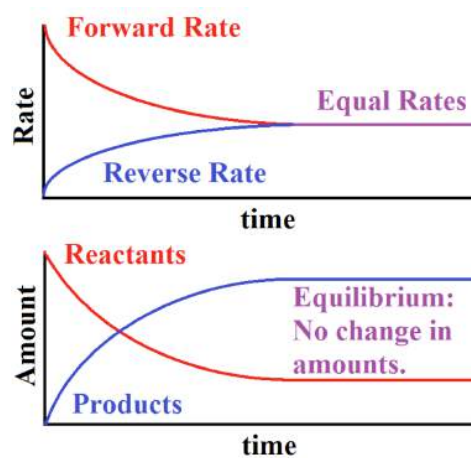
3. Equilibrium:

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).

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 $-78°C$ are placed in a flask that contains air at $21°C$. 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.

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

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.

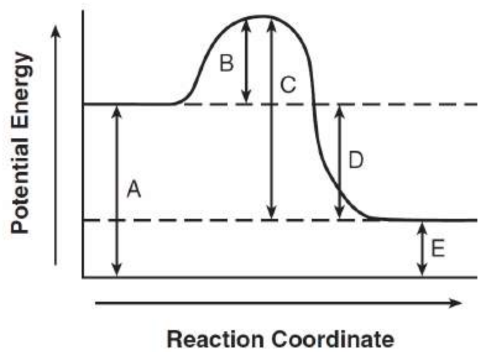
9. Several steps are involved in the industrial production of sulfuric acid. One step involves the oxidation of sulfur dioxide gas to form sulfur trioxide gas. A catalyst is used to increase the rate of production of sulfur trioxide gas. In a rigid cylinder with a movable piston, this reaction reaches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

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

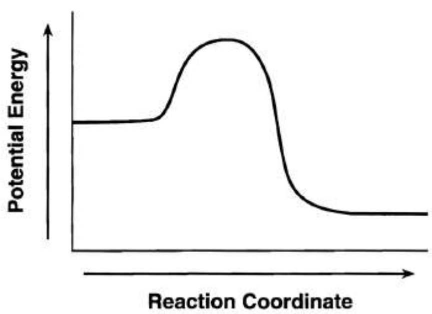
    Explain, in terms of Le Chatelier's principle, why the equilibrium shifts to the right to relieve the stress when the pressure on the system is increased at constant temperature.

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 equlbrium. – 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

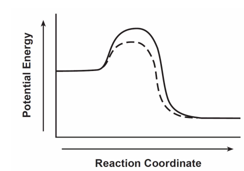
10. Acceptable responses include, but are not limited to: • A lower concentration of oxygen gas decreases the number of effective collisions between O_2 molecules and CH_4 molecules.

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.



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**

Metals	Nonmetals
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
H_2	
Cu	
Ag	
Au	

Most Active → Least Active

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

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.

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!!!!

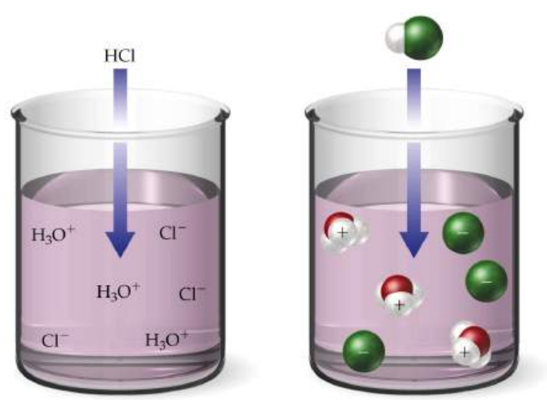
SALT:

- 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).



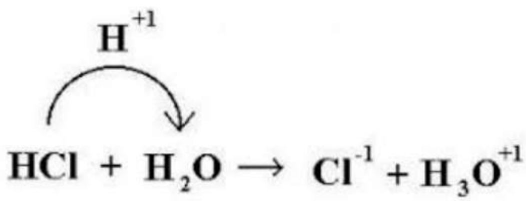
$HCl_{(g)} + H_2O_{(l)} \text{-->} H_3O^+_{(aq)} + Cl^-_{(aq)}$

hydrogen chloride molecule + water ---> hydronium ion in solution + chloride ion in solution

- 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.

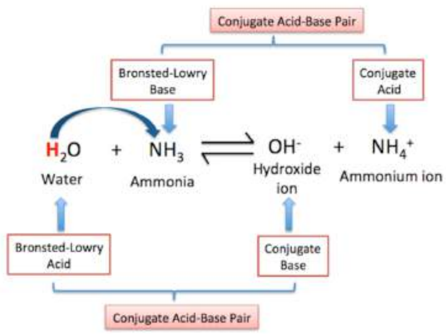
BRONSTED-LOWERY THEORY (ALTERNATE THEORY): BAAD

- 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



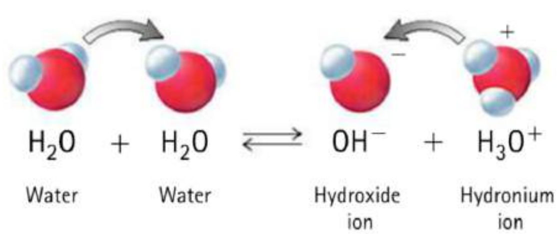
CONJUGATE ACIDS/BASES:

- 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:

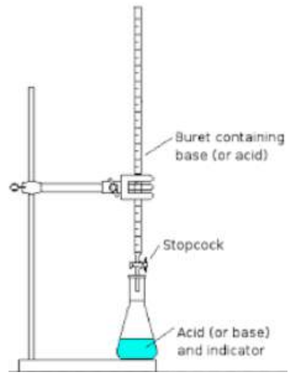
- An amphoteric substance can sometimes act like an acid and sometimes act like a base. Examples are $\mathbf{H_2O}$ **&** $\mathbf{HSO_4^-}$**.**



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^+ V_A = volume of acid	M_B = molarity of OH^- V_B = volume of base



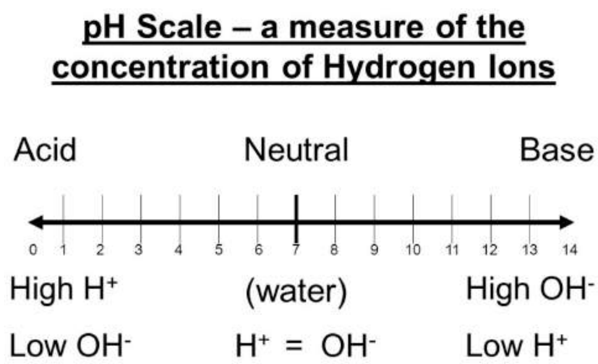
INDICATORS (TABLE M):

- 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



- If the H^+ concentration is 1.0 x 10^{-3}, then the pH is 3 (acidic)
- If the H^+ concentration is 1.0 x 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

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
$NaCl(aq)$	
$NaOH(aq)$	

Acids & Bases Review

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. $HCl(aq) + NaOH(aq) \rightarrow \underline{H_2O(\ell)} + \underline{NaCl(aq)}$
   $HCl(aq) + NaOH(aq) \rightarrow \underline{NaCl} + \underline{HOH}$
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.

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. **<u>Example:</u>** sulfate ion SO_4^{-2} . O has an oxidation number of -2, and therefore $(-2) \times (4) = -8$. 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**

	Metals	Nonmetals	
Most Active ↓	Li	F_2	Most Active ↓
	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.

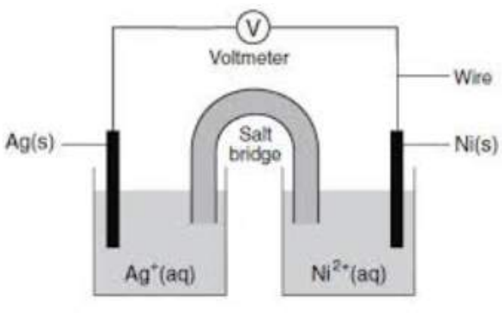
- SINGLE REPLACEMENT REACTIONS ARE **<u>ALWAYS</u>** REDOX
- DOUBLE REPLACMENT REACTIONS ARE **<u>NEVER</u>** REDOX
- Metals are easily oxidized (they lose electrons), Non-metals are easily reduced (gain electrons)
- A spontaneous redox reaction will occur when during a single replacement reaction the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

<u>Voltaic Cell</u>

- 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.

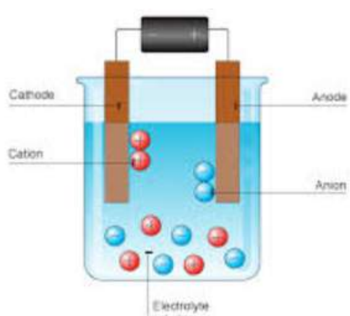
<u>Electrolytic Cell</u>

- 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.**



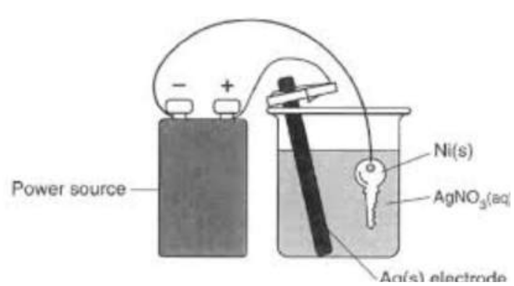
$2Ag^{+}(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

VOLTAIC



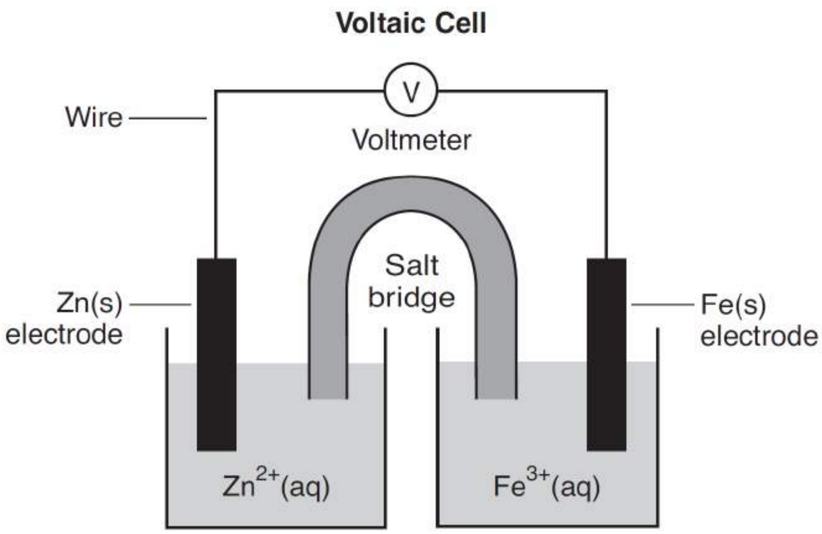
ELECTOLYTIC

ELECTROPLATING



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.



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

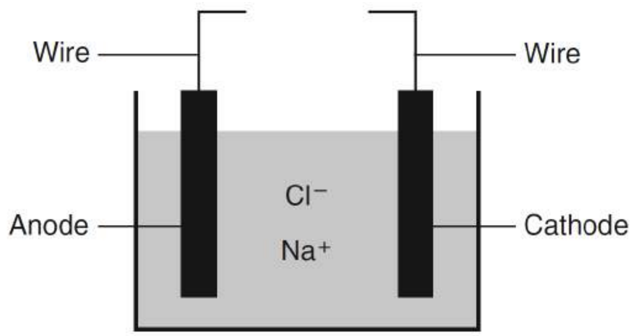
1. Explain, in terms of Zn atoms and Zn ions, why the mass of the Zn electrode *decreases as the cell operates.*

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.



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

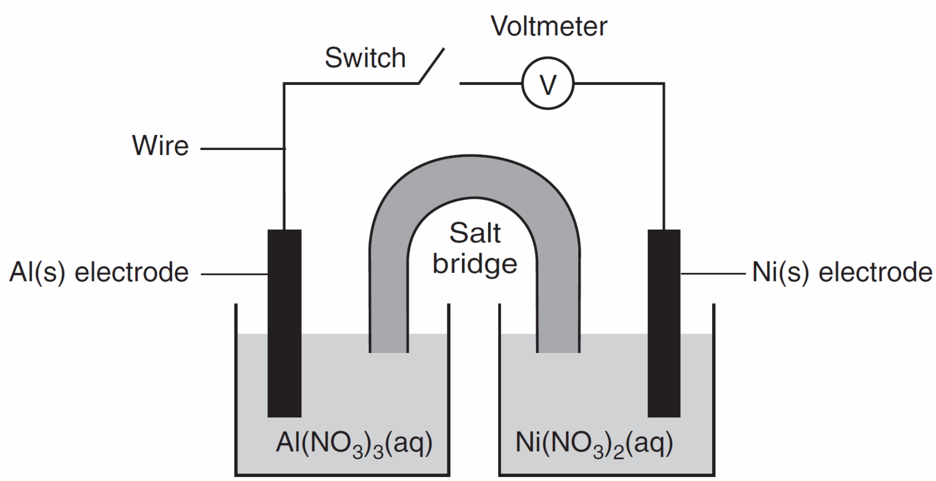
3. Write a balanced half-reaction equation for the reduction of the iron ions in iron(III) oxide to iron atoms.

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.

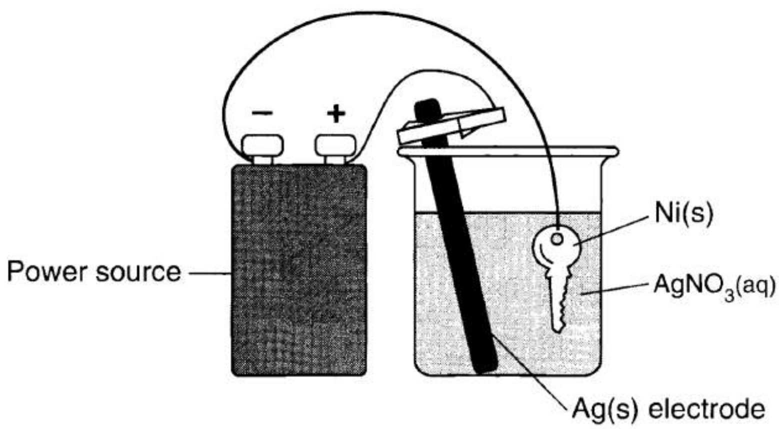
7. Determine the number of moles of Al(s) needed to completely react with 9.0 moles of Ni^{2+}(aq) ions.

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.



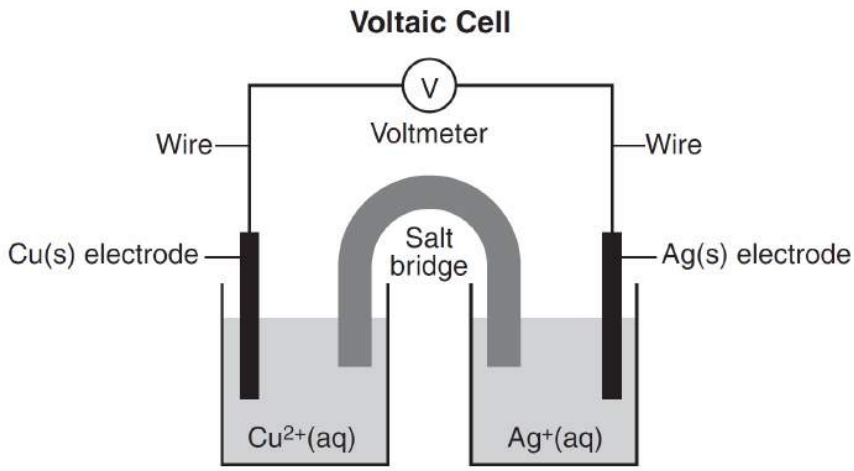
10. Explain, in terms of Ag atoms and Ag^{+}(aq) ions, why the mass of the silver electrode *decreases* as the cell operates.

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