Regents Chemistry

Topic Review Packet

Name: _____

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Topic 1: Matter, Its Properties and Changes Outline

1. Matter is classified as a pure substance or a mixture of substances. A pure substance (element or compound) has a constant composition and constant properties throughout a given sample, and from sample to sample.
- ✓ You can use particle models/diagrams to differentiate among elements, compounds, and mixtures.

2. The proportions of components in a mixture can be varied. Each component in a mixture retains its original properties. Differences in properties such as density, particle size, molecular polarity, boiling point and freezing point, and solubility permit physical separation of the components of the mixture.
- ✓ Methods of separating mixtures include evaporation, filtration, distillation, and chromatography.
- ✓ Mixtures can be homogeneous or heterogeneous. Solutions are always homogeneous. Heterogeneous mixtures are things like soil, fruit salad, where the composition is NOT uniform throughout the mixture.

3. The structure and arrangement of particles and their interactions determine the physical state of a substance at a given temperature and pressure.
- ✓ Know the states (phases) of the elements at STP; Br and Hg are the only 2 liquids, the noble gases as well as N, O, F, H, and Cl are gases, the rest are solids
- ✓ Know the 7 elements that are diatomic in their natural states; "7-Up" or "HOFBrINCl".
- ✓ Draw particle models of solids, liquids, and gases.

4. A physical change results in the rearrangement of existing particles in a substance; no new types of particles result from this type of change.   A chemical change results in the formation of different particles with changed properties.
- ✓ Distinguish between chemical and physical changes based on whether new substances form or not.

5. Properties can be physical or chemical. Physical properties describe those characteristics that can be observed with the senses or measured. Chemical properties describe how the substance interacts with other substances.

- ✓ Distinguish between chemical and physical properties.
- ✓ One of the more useful properties is density. The density equation is on Table T; D=m/V.
- ✓ Some common properties of the elements are found on Table S, such as melting and boiling points.

Matter – Cut from Jan 2007 – Jan 2008 Exams

1. A sample composed only of atoms having the same atomic number is classified as
    (1) a compound   (3) an element
    (2) a solution   (4) an isomer

2. A dilute, aqueous potassium nitrate solution is best classified as a
(1) homogeneous compound
(2) homogeneous mixture
(3) heterogeneous compound
(4) heterogeneous mixture

3. At which Celsius temperature does lead change from a solid to a liquid?
    (1) 874°C   (3) 328°C
    (2) 601°C   (4) 0°C

4. Which statement describes a chemical property of hydrogen gas?
(1) Hydrogen gas burns in air.
(2) Hydrogen gas is colorless.
(3) Hydrogen gas has a density of 0.000 09 g/cm^3 at STP.
(4) Hydrogen gas has a boiling point of 20. K at standard pressure.

5. Which element has the greatest density at STP?
    (1) calcium   (3) chlorine
    (2) carbon   (4) copper

6. Which statement describes a chemical property of the element magnesium?
(1) Magnesium is malleable.
(2) Magnesium conducts electricity.
(3) Magnesium reacts with an acid.
(4) Magnesium has a high boiling point.

7. Matter that is composed of two or more different elements chemically combined in a fixed proportion is classified as
    (1) a compound   (3) a mixture
    (2) an isotope   (4) a solution

8. Which element is a solid at STP and a good conductor of electricity?
    (1) iodine   (3) nickel
    (2) mercury   (4) sulfur

9. The table below shows mass and volume data for four samples of substances at 298 K and 1 atmosphere.

Masses and Volumes of Four Samples

Sample	Mass (g)	Volume (mL)
A	30.	60.
B	40.	50.
C	45	90.
D	90.	120.

Which two samples could consist of the same substance?
    (1) *A* and *B*   (3) *B* and *C*
    (2) *A* and *C*   (4) *C* and *D*

10. Bronze contains 90 to 95 percent copper and 5 to 10 percent tin. Because these percentages can vary, bronze is classified as
    (1) a compound   (3) a mixture
    (2) an element (4) a substance

11. At STP, which list of elements contains a solid, a liquid, and a gas?
    (1) Hf, Hg, He (3) Ba, Br$_2$, B
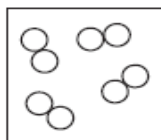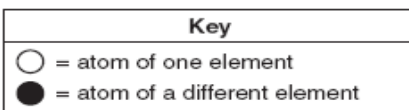   (2) Cr, Cl$_2$, C   (4) Se, Sn, Sr

12. A 10.0-gram sample of which element has the *smallest* volume at STP?
    (1) aluminum   (3) titanium
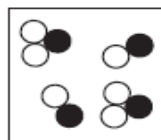   (2) magnesium   (4) zinc

13. At room temperature, a mixture of sand and water can be separated by
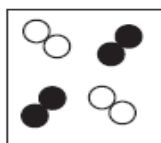   (1) ionization   (3) filtration
    (2) combustion   (4) sublimation

14. Which particle diagram represents a sample of one compound, only?
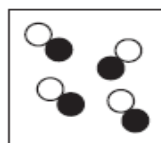
15. A 1.00-mole sample of neon gas occupies a volume of 24.4 liters at 298 K and 101.3 kilopascals. Calculate the density of this sample. Your response must include *both* a correct numerical setup and the calculated result.

Base your answers to questions 16 through 18 on the information below.

In an investigation, a dripless wax candle is massed and then lighted. As the candle burns, a small amount of liquid wax forms near the flame. After 10 minutes, the candle's flame is extinguished and the candle is allowed to cool. The cooled candle is massed.

16. Identify *one* physical change that takes place in this investigation.

17. State *one* observation that indicates a chemical change has occurred in this investigation.

18. Draw a particle diagram showing the change from solid wax to liquid wax. Use "●" for particles of wax. Draw separate diagrams for the liquid and the solid states.

Base your answers to questions 19 through 21 on the particle diagrams below, which show atoms and/ or molecules in three different samples of matter at STP.

Key
⊗ atom of x
● atom of y
○ atom of z

Sample 1   Sample 2   Sample 3
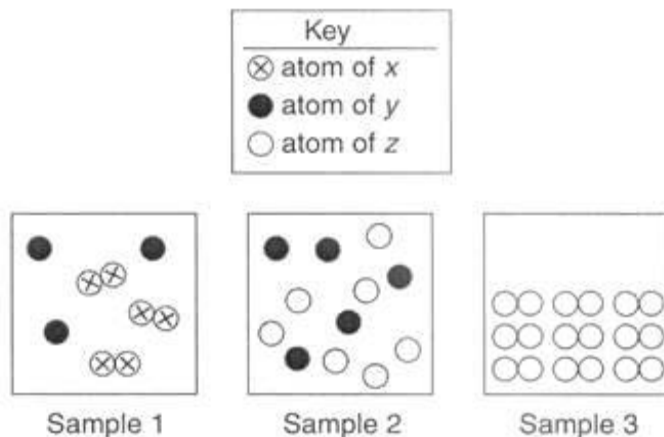
19. Which sample represents a pure substance?

20. When two atoms of y react with one atom of z, a compound forms. Using the number of atoms shown in sample 2, what is the maximum number of molecules of this compound that can be formed?

21. Explain why xx does not represent a compound.

Topic 2: Atomic Concepts Outline

1. The modern model of the atom has evolved over a long period of time through the work of many scientists.
- ✓ *Dalton's Model:*
    - Elements are made of atoms
    - Atoms of an element are the same.
    - Compounds are formed from combinations of atoms.
- ✓ *Rutherford Experiment*
    - Bombarded gold foil with alpha particles. Showed atoms were mostly empty space with small, dense positively charged nucleus.
- ✓ *Bohr Model*
    - Small, dense, positively charged nucleus surrounded by electrons in circular orbits.
- ✓ *Wave-Mechanical Model (Modern Atomic Theory)*
    - Small, dense, nucleus positively charged nucleus surrounded by electrons moving in "electron cloud".
    - **"Orbitals"** are areas where an electron with a certain amount of energy is *most likely* to be found.

2. Each atom is made of a positively charged nucleus with one or more orbiting, negatively charged electrons.

3. Protons and neutrons are found in the nucleus.
- ✓ The number of protons in an atom's nucleus gives the nucleus a positive charge. Li has a "nuclear charge" of +3, since it has 3 protons.

4. Protons have a positive charge, neutrons no charge, and electrons a negative charge.

5. The number of protons in an atom equals the number of electrons.
- ✓ The positive charges of the protons are cancelled by the negative charges of the electrons, so overall an *atom* has a neutral charge.

6. The mass of a proton is 1 amu. The mass of a neutron is 1 amu. The mass of an electron is almost 0 amu.
- ✓ The mass of an atom is contained in its nucleus.
- ✓ The <u>atomic mass</u> of an atom is equal to the total number of protons and neutrons.

7. Each electron in an atom has its own distinct amount of energy.
- ✓ When all electrons are at their lowest possible energy, it is called the "<u>ground state</u>."
- ✓ Electrons fill in energy levels and orbitals starting with the one that requires the least energy and progressively move to those levels and orbitals that require increasing amounts of energy.

8. When the electron gains a specific amount of energy, it moves to a higher orbital and is in the "<u>excited state</u>".
- ✓ You can recognize an excited state electron configuration. If the configuration does not match that on the Periodic Table for that number of electrons, then it is an excited state.

9. When an electron returns from a higher energy state to a lower energy state, it emits a specific amount of energy usually in the form of <u>light</u>. This can be used to identify an element (bright line spectrum).
- ✓ The instrument used to see the bright line spectrum is called a spectroscope.

10. The outermost electrons are called <u>valence electrons</u>. These affect the chemical properties of the element.
- ✓ Atoms with a filled valence level are stable (noble gases).
- ✓ Most elements can have up to 8 electrons in their valence level. The exceptions are H and He, which can have only 2 valence electrons.
- ✓ Atoms form bonds in order to fill their valence levels.
- ✓ You can use <u>Lewis structures</u> to show the configuration of the valence electrons.

11. Atoms of the same element all contain the same number of protons.
- ✓ Changing the number of protons changes the atom into a different element.
- ✓ The atomic number is the number of protons in an atom of an element.

12. <u>Isotopes</u> are atoms with equal numbers of protons but different numbers of neutrons.
- ✓ Isotopes of an element have the same atomic number (protons only), but different atomic masses (protons + neutrons).

13. The average atomic mass of an element is the weighted average of its naturally occurring isotopes.
- ✓ You need to know how to do the calculation of "weighted atomic mass" given isotope masses and percent abundances.

14. When an atom gains an electron, it becomes a negative ion and its radius increases.

15. When an atom loses an electron, it becomes a positive ion and its radius decreases.

16. <u>Electronegativity</u> indicates how strongly an atom of an element attracts electrons in a chemical bond. These values are based on an arbitrary scale.
- ✓ Fluorine has the highest electronegativity of all elements (4.00).

Atomic Structure – Practice Questions

1. Experiments performed to reveal the structure of atoms led scientists to conclude that an atom's
   (1) positive charge is evenly distributed throughout its volume
   (2) negative charge is mainly concentrated in its nucleus
   (3) mass is evenly distributed throughout its volume
   (4) volume is mainly unoccupied

2. The modern model of the atom shows that electrons are
   (1) orbiting the nucleus in fixed paths
   (2) found in regions called orbitals
   (3) combined with neutrons in the nucleus
   (4) located in a solid sphere covering the nucleus

3. An experiment in which alpha particles were used to bombard thin sheets of gold foil led to the conclusion that an atom is composed mostly of
   (1) empty space and has a small, negatively charged nucleus
   (2) empty space and has a small, positively charged nucleus
   (3) a large, dense, positively charged nucleus
   (4) a large, dense, negatively charged nucleus

4. What is the atomic number of an element that has six protons and eight neutrons?
   (1) 6   (2) 2   (3) 8   (4) 14

5. An atom of fluorine has a mass of 19 atomic mass units. The total number of protons and neutrons in its nucleus is
   (1) 9   (2) 10   (3) 19   (4) 28

6. What is the total number of protons contained in the nucleus of a carbon-14 atom?
   (1) 6   (2) 8   (3) 12   (4) 14

7. What is the nuclear charge of an iron atom?
   (1) +26   (2) +30   (3) +56   (4) +82

8. Which of these elements has an atom with the most stable outer electron configuration?
   (1) Ne   (2) Cl   (3) Ca   (4) Na

9. How many electrons are in the outermost principal energy level of an atom of carbon in the ground state?
   (1) 6   (2) 2   (3) 3   (4) 4

10. Which electron configuration is correct for a sodium ion?
    (1) 2-7   (2) 2-8   (3) 2-8-1   (4) 2-8-2

11. What is the electron configuration of a sulfur atom in the ground state?
    (1) 2-4   (2) 2-6   (3) 2-8-4   (4) 2-8-6

12. The nucleus of which atom contains 48 neutrons?

(1) $^{32}_{16}S$ (2) $^{48}_{22}Ti$ (3) $^{85}_{37}Rb$ (4) $^{112}_{48}Cd$

13. The number of neutrons in the nucleus of an atom can be determined by
    (1) adding the atomic number to the mass number
    (2) subtracting the atomic number from the mass number
    (3) adding the mass number to the atomic mass
    (4) subtracting the mass number from the atomic number

14. When an atom loses an electron, the atom becomes an ion that is
    (1) positively charged and gains a small amount of mass
    (2) positively charged and loses a small amount of mass
    (3) negatively charged and gains a small amount of mass
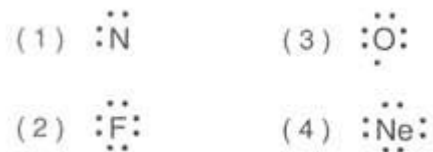   (4) negatively charged and loses a small amount of mass

15. In which pair of elements do the nuclei of the atoms contain the same number of neutrons?

(1) $^{7}_{3}Li$ and $^{9}_{4}Be$ (3) $^{23}_{11}Na$ and $^{24}_{12}Mg$

(2) $^{14}_{7}N$ and $^{16}_{8}O$ (4) $^{32}_{16}S$ and $^{35}_{17}Cl$

16. The characteristic spectral lines of elements are caused when electrons in an excited atom move from
    (1) lower to higher energy levels, releasing energy
    (2) lower to higher energy levels, absorbing energy
    (3) higher to lower energy levels, releasing energy
    (4) higher to lower energy levels, absorbing energy

17. Which Lewis electron-dot structure is drawn correctly for the atom it represents?

(1) $:\!\overset{\cdot\cdot}{N}$ (3) $:\!\overset{\cdot\cdot}{\underset{\cdot}{O}}\!:$

(2) $:\!\overset{\cdot\cdot}{\underset{\cdot\cdot}{F}}\!:$ (4) $:\!\overset{\cdot\cdot}{\underset{\cdot\cdot}{Ne}}\!:$

18. When a lithium atom forms a Li^{+} ion, the lithium atom

(1) gains a proton   (3) loses an electron
(2) loses a proton   (4) gains an electron

19. What is the total number of electrons in the valence shell of an atom of aluminum in the ground state?

(1) 8   (2) 2   (3) 3   (4) 10

20. An electron in an atom moves from the ground state to an excited state when the energy of the electron

(1) increases   (2) decreases   (3) remains the same

21. During a flame test, ions of a specific metal are heated in the flame of a gas burner. A characteristic color of light is emitted by these ions in the flame when the electrons
    (1) emit energy as they move to higher energy levels
    (2) emit energy as they return to lower energy levels
    (3) gain energy as they move to higher energy levels
    (4) gain energy as they return to lower energy levels

22. What is the total number of electrons in a Cu^+ ion?
    (1) 36   (2) 29   (3) 30   (4) 28

Base your answers to questions 23 and 24 on the information and the bright-line spectra represented below.

Many advertising signs depend on the production of light emissions from gas-filled glass tubes that are subjected to a high-voltage source. When light emissions are passed through a spectroscope, bright-line spectra are produced.

Gas A

Gas B

Gas C

Gas D
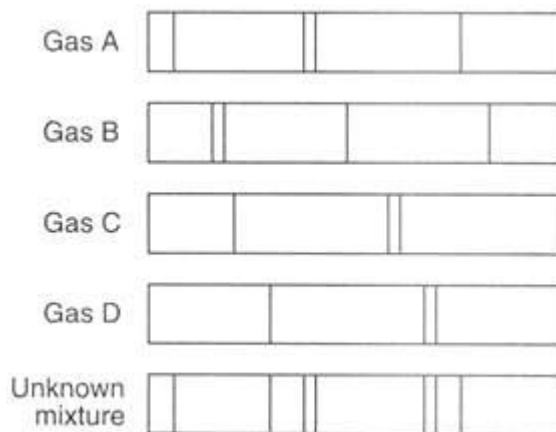
Unknown mixture

23. Identify the two gases in the unknown mixture.

24. Explain the production of an emission spectrum in terms of the energy states of an electron.

Atomic Concepts Review – Cut from Jan 2007 – Jan 2008 Exams

1. Which subatomic particles are located in the nucleus of a neon atom?
 (1) electrons and positrons
 (2) electrons and neutrons
 q(3) protons and neutrons
 (4) protons and electrons

2. The total mass of the protons in an atom of gold-198 is approximately
 (1) 79 atomic mass units
 (2) 119 atomic mass units
 (3) 198 atomic mass units
 (4) 277 atomic mass units

3. In a calcium atom in the ground state, the electrons that possess the *least* amount of energy are located in the
 (1) first electron shell
 q(2) second electron shell
 q(3) third electron shell
 (4) fourth electron shell

4. Which group of atomic models is listed in historical order from the earliest to the most recent?
(1) hard-sphere model, wave-mechanical model, electron-shell model
(2) hard-sphere model, electron-shell model, wave-mechanical model
(3) electron-shell model, wave-mechanical model, hard-sphere model
(4) electron-shell model, hard-sphere model, wave-mechanical model

5 Which isotopic notation represents an atom of carbon-14?
 (1) $^6_8 C$ (3) $^6_{14} C$
 (2) $^8_6 C$ (4) $^{14}_6 C$

6. Which isotopic notation identifies a metalloid that is matched with the corresponding number of protons in each of its atoms?
 (1) ^{24}Mg and 12 protons
 (2) ^{28}Si and 14 protons
 (3) ^{75}As and 75 protons
 (4) ^{80}Br and 80 protons

7. According to the wave-mechanical model of the atom, electrons in an atom
 (1) travel in defined circles
 (2) are most likely found in an excited state
 (3) have a positive charge
 (4) are located in orbitals outside the nucleus

8. What is the total charge of the nucleus of a carbon atom?
 (1) –6   (3) +6
 (2) 0   (4) +12

9. A sample composed only of atoms having the same atomic number is classified as
 (1) a compound   (3) an element
 (2) a solution   (4) an isomer

10. Which two particles each have a mass approximately equal to one atomic mass unit?
 (1) electron and neutron
 (2) electron and positron
 (3) proton and electron
 (4) proton and neutron

11. Which electron configuration could represent a strontium atom in an excited state?
 (1) 2–8–18–7–1   (3) 2–8–18–8–1
 (2) 2–8–18–7–3   (4) 2–8–18–8–2

12. What is the total number of neutrons in an atom of $^{57}_{26}Fe$?
 (1) 26   (3) 57
 (2) 31   (4) 83

13. What is the total number of electrons in a

Mg^{+2} ion?

(1) 10   (3) 14

(2) 12   (4) 24

14. What was concluded about the structure of the atom as the result of the gold foil experiment?
(1) A positively charged nucleus is surrounded by positively charged particles.
(2) A positively charged nucleus is surrounded by mostly empty space.
(3) A negatively charged nucleus is surrounded by positively charged particles.
(4) A negatively charged nucleus is surrounded by mostly empty space.

15. An atom is electrically neutral because the
(1) number of protons equals the number of electrons
(2) number of protons equals the number of neutrons
(3) ratio of the number of neutrons to the number of electrons is 1:1
(4) ratio of the number of neutrons to the number of protons is 2:1

16. How do the energy and the most probable location of an electron in the third shell of an atom compare to the energy and the most probable location of an electron in the first shell of the same atom?
(1) In the third shell, an electron has more energy and is closer to the nucleus.
(2) In the third shell, an electron has more energy and is farther from the nucleus.
(3) In the third shell, an electron has less energy and is closer to the nucleus.
(4) In the third shell, an electron has less energy and is farther from the nucleus.

17. What is the net charge on an ion that has 9 protons, 11 neutrons, and 10 electrons?

(1) 1+   (3) 1–

(2) 2+   (4) 2–

18. Which value of an element is calculated using both the mass and the relative abundance of each of the naturally occurring isotopes of this

element?

(1) atomic number   (3) half-life

(2) atomic mass   (4) molar volume

19. Which two notations represent different isotopes of the same element?

(1) $^{6}_{4}Be$ and $^{9}_{4}Be$ (3) $^{14}_{7}N$ and $^{14}_{6}C$

(2) $^{7}_{3}Li$ and $^{7}_{3}Li$ (4) $^{32}_{15}P$ and $^{32}_{16}S$

Base your answers to questions 20 through 22 on the information below.

The accepted values for the atomic mass and percent natural abundance of each naturally occurring isotope of silicon are given in the data table below.
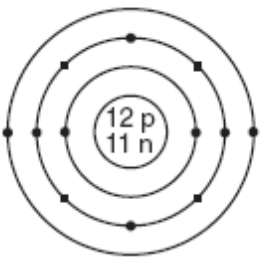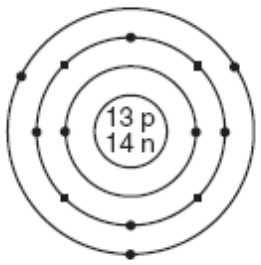
Naturally Occurring Isotopes of Silicon

Isotope	Atomic Mass (atomic mass units)	Percent Natural Abundance (%)
Si-28	27.98	92.22
Si-29	28.98	4.69
Si-30	29.97	3.09

20. Determine the total number of neutrons in an atom of Si-29. [1]

21. Show a correct numerical setup for calculating the atomic mass of Si. [1]

22. A scientist calculated the percent natural abundance of Si-30 in a sample to be 3.29%. Determine the percent error for this value. [1]

23. Write *one* electron configuration for an atom of silicon in an excited state.

Base your answers to questions 24 through 26 on the information below.

Atomic Diagrams of Magnesium and Aluminum

Key	Element	Lewis Electron-Dot Diagram	Electron-Shell Diagram
• = electron	magnesium	Mg:	12 p / 11 n
	aluminum	Al:	13 p / 14 n

24. Identify *one* piece of information shown in the electron-shell diagrams that is *not* shown in the Lewis electron-dot diagrams. [1]

25. Determine the mass number of the magnesium atom represented by the electron-shell diagram. [1]

26. Explain why Lewis electron-dot diagrams are generally more suitable than electron-shell diagrams for illustrating chemical bonding. [1]

Topic 3: Periodic Table Outline

1. The placement of an element on the Periodic Table gives an indication of the chemical and physical properties of that element.
- ✓ Elements to the left of the stair step line are metals, and therefore are easily oxidized (lose electrons) in bonding situations, are good electrical conductors, are shiny, malleable, ductile, and have low ionization energies and electronegativities.
- ✓ Elements to the right of the stair step line, but not in Group 18 are nonmetals, and therefore react to gain electrons (get reduced), are not conductors, are dull appearing, brittle, have high ionization energies and electronegativities.
- ✓ Some of the elements along the stair step line have properties of both metals and non-metals and are known as "metalloids" or "semi-metals".
- ✓ Elements in Group 18 are the noble gases and they are chemically inert (unreactive) and have extremely high ionization energies.

2. Elements are arranged in order of increasing atomic number (NOT MASS!)

3. The number of protons in an atom (atomic number) identifies the element.
- ✓ The number of protons in an atom only changes through nuclear reactions.

4. The atomic mass is the sum of protons and neutrons in the nucleus.
- ✓ The mass number given on the periodic table is a weighted average of the different isotopes of that element.
- ✓ Electrons do not significantly add to the atomic mass.

5. Isotopes of an element are identified by the sum of protons and neutrons.
- ✓ Isotopes of the same element have the same number of protons and a different number of neutrons.
- ✓ Examples of isotopic notation are: $^{14}_{6}C$, ^{14}C, carbon-14, C-14

6. Elements can be classified by their properties and their location on the Periodic Table as metals, non-metals, metalloids, and noble gases.

7. Elements may be differentiated by their physical properties.
- ✓ Ex: Density, conductivity, malleability, hardness, ductility, solubility

8. Elements may be differentiated by their chemical properties.
- ✓ Chemical properties describe how an element behaves in a chemical reaction.

9. Elements are arranged into periods and groups.

10. Elements of the same period have the same number of occupied energy levels.

11. Elements of the same group have the same valence configuration and similar chemical properties.
✓ Group 1 elements other than H are *alkali metals*.
✓ Group 2 elements are *alkali earth metals*.
✓ Group 17 elements are *halogens.*
✓ Alkali metals, alkali earth metals, and halogens all are highly reactive and do not exist as free elements in nature (they are all found in compounds).
✓ Group 18 elements are *noble or inert gases*. These elements have filled valence levels and are do not normally react with other substances.

12. The succession of elements within a group demonstrates characteristic trends in properties. As you progress *down* a group:
✓ atomic radius increases.
✓ electronegativity decreases.
✓ first ionization energy decreases.
✓ metallic character increases.

13. The succession of elements within a period demonstrates characteristic trends in properties. As you progress *across* a group from *left to right*:
✓ atomic radius decreases.
✓ electronegativity increases.
✓ first ionization energy increases.
✓ metallic character decreases.

14. Some elements may exist in two or more forms in the same phase. These forms differ in their molecular or crystal structure, hence their different properties. These different forms are called "allotropes,"
✓ Ex: Solid carbon exists in three different forms: graphite, diamond (a network solid) and coal.
✓ Ex: the element oxygen can exist in two different forms: O_2 gas and ozone (O_3 gas)

Periodic Table – Practice Questions

1. Elements in the Periodic Table are arranged according to their
   - (1) atomic number
   - (2) atomic mass
   - (3) relative activity
   - (4) relative size

2. Elements in a given period of the Periodic Table contain the same number of
   - (1) protons in the nucleus
   - (2) neutrons in the nucleus
   - (3) electrons in the outermost level
   - (4) occupied principal energy levels

3. Atoms of metals tend to
   - (1) lose electrons and form negative ions
   - (2) lose electrons and form positive ions
   - (3) gain electrons and form negative ions
   - (4) gain electrons and form positive ions

4. Which properties are most common in nonmetals?
   - (1) low ionization energy and low electronegativity
   - (2) low ionization energy and high electronegativity
   - (3) high ionization energy and low electronegativity
   - (4) high ionization energy and high electronegativity

5. Which two elements have chemical properties that are most similar?
   - (1) Cl and Ar
   - (2) Li and Na
   - (3) K and Ca
   - (4) C and N

6. Which of the following Period 4 elements has the most metallic characteristics?
   - (1) Ca
   - (2) Ge
   - (3) As
   - (4) Br

7. If M represents an alkali metal, what is the formula for the compound formed by M and oxygen?
   - (1) MO_2
   - (2) M_2O
   - (3) M_2O_3
   - (4) M_3O_2

8. As the elements in Group 15 are considered in order of increasing atomic number, which sequence in properties occurs?
   - (1) nonmetal, metalloid, metal
   - (2) metalloid, metal, nonmetal
   - (3) metal, metalloid, nonmetal
   - (4) metal, nonmetal, metalloid

9. Which group contains a metalloid?
   - (1) 1
   - (2) 11
   - (3) 15
   - (4) 18

10. As elements of Group 15 of the Periodic Table are considered in order from top to bottom, the metallic character of the atoms of each successive element generally
    - (1) increases
    - (2) decreases
    - (3) remains the same

11. Which statement best describes Group 2 elements as they are considered in order from top to bottom of the Periodic Table?
    (1) The number of principal energy levels increases, and the number of valence electrons increases.
    (2) The number of principal energy levels increases, and the number of valence electrons remains the same.
    (3) The number of principal energy levels remains the same, and the number of valence electrons increases.
    (4) The number of principal energy levels remains the same, and the number of valence electrons decreases.

12. Which Group 15 element exists as a diatomic molecule at STP?
    (1) phosphorous   (3) bismuth
    (2) nitrogen   (4) arsenic

13. Which Group 16 element when combined with hydrogen forms a compound that would exhibit the strongest hydrogen bonding?
    (1) selenium   (3) oxygen
    (2) tellurium   (4) sulfur

14. Which ion has the largest radius?
    (1) Na^+ (2) Mg^{2+} (3) K^+ (4) Ca^{2+}

15. What occurs as the atomic number of the elements in Period 2 increases?
    (1) The nuclear charge of each successive atom decreases, and the covalent radius decreases.
    (2) The nuclear charge of each successive atom decreases, and the covalent radius increases.
    (3) The nuclear charge of each successive atom increases, and the covalent radius decreases.
    (4) The nuclear charge of each successive atom increases, and the covalent radius increases.

Periodic Table – Cut from Jan 2007 – Jan 2008 Exams

1. Which element is a solid at STP and a good conductor of electricity?
    (1) iodine   (3) nickel
    (2) mercury   (4) sulfur

2. Which element has both metallic and nonmetallic properties?
    (1) Rb   (3) Si
    (2) Rn   (4) Sr

3. The carbon atoms in graphite and the carbon atoms in diamond have different
    (1) atomic numbers
    (2) atomic masses
    (3) electronegativities
    (4) structural arrangements

4. Atoms of which element have the greatest tendency to gain electrons?
    (1) bromine   (3) fluorine
    (2) chlorine   (4) iodine

5. Which statement describes a chemical property of the element magnesium?
    (1) Magnesium is malleable.
    (2) Magnesium conducts electricity.
    (3) Magnesium reacts with an acid.
    (4) Magnesium has a high boiling point.

6. Which statement explains why sulfur is classified as a Group 16 element?
    (1) A sulfur atom has 6 valence electrons.
    (2) A sulfur atom has 16 neutrons.
    (3) Sulfur is a yellow solid at STP.
    (4) Sulfur reacts with most metals.

7. How do the atomic radius and metallic properties of sodium compare to the atomic radius and metallic properties of phosphorus?
    (1) Sodium has a larger atomic radius and is more metallic.
    (2) Sodium has a larger atomic radius and is lessmetallic.
    (3) Sodium has a smaller atomic radius and is more metallic.
    (4) Sodium has a smaller atomic radius and is less metallic.

8. Which list of elements consists of metalloids, only?
    (1) B, Al, Ga   (3) O, S, Se
    (2) C, N, P   (4) Si, Ge, As

9. Which general trend is found in Period 2 on the Periodic Table as the elements are considered in order of increasing atomic number?
    (1) decreasing atomic mass
    (2) decreasing electronegativity
    (3) increasing atomic radius
    (4) increasing first ionization energy

10. Which two characteristics are associated with metals?
    (1) low first ionization energy and low electronegativity
    (2) low first ionization energy and high electronegativity
    (3) high first ionization energy and low electronegativity
    (4) high first ionization energy and high electronegativity

11. Which element is most chemically similar to chlorine?
    (1) Ar   (3) Fr
    (2) F   (4) S

12. Which grouping of circles, when considered in order from the top to the bottom, best represents the relative size of the atoms of Li, Na, K, and Rb, respectively?

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

13. At STP, which element is brittle and *not* a conductor of electricity?
    (1) S   (3) Na
    (2) K   (4) Ar

14. An atom of argon rarely bonds to an atom of another element because an argon atom has
    (1) 8 valence electrons
    (2) 2 electrons in the first shell
    (3) 3 electron shells
    (4) 22 neutrons

15. The elements on the Periodic Table are arranged in order of increasing
    (1) boiling point   (3) atomic number
    (2) electronegativity   (4) atomic mass

16. Which element is classified as a nonmetal?
    (1) Be   (3) Si
    (2) Al   (4) Cl

17. Solid samples of the element phosphorus can be white, black, or red in color. The variations in color are due to different
    (1) atomic masses
    (2) molecular structures
    (3) ionization energies
    (4) nuclear charges

18. Lithium and potassium have similar chemical properties because the atoms of both elements have the same
    (1) mass number
    (2) atomic number
    (3) number of electron shells
    (4) number of valence electrons

19. At STP, which list of elements contains a solid, a liquid, and a gas?
    (1) Hf, Hg, He   (3) Ba, Br_2, B
    (2) Cr, Cl_2, C   (4) Se, Sn, Sr

Base your answers to questions 20 through 22 on the information below.

Elements with atomic numbers 112 and 114 have been produced and their IUPAC names are pending approval. However, an element that would be put between these two elements on the Periodic Table has not yet been produced. If produced, this element will be identified by the symbol Uut until an IUPAC name is approved.

20. Draw a Lewis electron-dot diagram for an atom of Uut. [1]

21. Determine the charge of an Uut nucleus. Your response must include *both* the numerical value and the sign of the charge. [1]

22. Identify *one* element that would be chemically similar to Uut. [1]

Base your answers to questions 23 through 26 on the information below, which describes the proposed discovery of element 118.

In 1999, a nuclear chemist and his team announced they had discovered a new element by crashing krypton atoms into lead. The new element, number 118, was assigned the name ununoctium and the symbol Uuo. One possible isotope of ununoctium could have been Uuo-291.

However, the discovery of Uuo was not confirmed because other scientists could not reproduce the experimental results published by the nuclear chemist and his team. In 2006, another team of scientists claimed that they produced Uuo. This claim has yet to be confirmed.

Adapted from Discover January 2002

23. Based on atomic number, in which group on the Periodic Table would element 118 be placed? [1]

24. What would be the total number of neutrons present in a theoretical atom of Uuo-291?[1]

25. What would be the total number of electrons present in a theoretical atom of Uuo-291? [1]

26. Explain why being able to reproduce scientific results is an important component of scientific research. [1]

Base your answers to questions 27 through 30 on the information below.

The table below lists physical and chemical properties of six elements at standard pressure that correspond to known elements on the Periodic Table. The elements are identified by the code letters, *D, E, G, J, L*, and *Q*.

Properties of Six Elements at Standard Pressure

Element D	Element E	Element G
Density 0.00018 g/cm³	Density 1.82 g/cm³	Density 0.53 g/cm³
Melting point –272°C	Melting point 44°C	Melting point 181°C
Boiling point –269°C	Boiling point 280°C	Boiling point 1347°C
Oxide formula (none)	Oxide formula E_2O_5	Oxide formula G_2O
Element J	**Element L**	**Element Q**
Density 0.0013 g/cm³	Density 0.86 g/cm³	Density 0.97 g/cm³
Melting point –210°C	Melting point 64°C	Melting point 98°C
Boiling point –196°C	Boiling point 774°C	Boiling point 883°C
Oxide formula J_2O_5	Oxide formula L_2O	Oxide formula Q_2O

27. What is the total number of elements in the "Properties of Six Elements at Standard Pressure" table that are solids at STP? [1]

28. An atom of element *G* is in the ground state. What is the total number of valence electrons in this atom? [1]

29. Letter *Z* corresponds to an element on the Periodic Table other than the six listed elements. Elements *G, Q, L,* and *Z* are in the same group on the Periodic Table, as shown in the diagram below.

G
Q
L
Z

Based on the trend in the melting points for elements *G, Q,* and *L* listed in the "Properties of Six Elements at Standard Pressure" table, estimate the melting point of element *Z*, in degrees Celsius. [1]

30. Identify, by code letter, the element that is a noble gas in the "Properties of Six Elements at Standard Pressure" table. [1]

Topic 4: Formulas & Names, Equations, Moles, Molar Mass, & Types of Reactions Outline

1. A compound is a substance composed of two or more different elements that are chemically combined in a fixed proportion. A chemical compound can only be broken down by chemical means.

2. Chemical compounds can be represented by a specific formula and assigned a name based on the IUPAC system.

3. Types of chemical formulas include empirical, molecular, and structural.
- ✓ Empirical formulas show elements in their simplest whole number ratios. This may or may not be the same as the molecular formula.
- ✓ Molecular formulas show the actual number of atoms per element in a single molecule.
- ✓ Structural formulas show the number of each type of atom as well as their physical arrangement.

4. All chemical reactions show a conservation of mass, energy and charge.

5. A balanced chemical equation represents conservation of atoms.

6. The coefficients in a balanced chemical equation can be used to determine <u>mole ratios</u> in the reaction, and can further be used to predict relationships about amounts between products and reactants.

7. The molar mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gram formula mass) equals the mass of one mole of that substance.

8. The percent composition by mass of each element in a compound can be calculated mathematically.

9. Types of chemical reactions include synthesis, decomposition, single replacement, and double replacement.

Equations & Stoichiometry – Practice Questions

1. Which substance has the greatest molecular mass?
   (1) H_2O_2 (2) NO   (3) CF_4 (4) I_2

2. What is the gram formula mass of $Ca(OH)_2$?
   (1) 29 g   (2) 34 g   (3) 57 g   (4) 74 g

3. What is the total number of moles of atoms present in 1 gram formula mass of $Pb(C_2H_3O_2)_2$?
   (1) 9   (2) 14   (3) 3   (4) 15

4. The percent by mass of carbon in $HC_2H_3O_2$ is equal to
   (1) $\dfrac{12}{60} \times 100$ (2) $\dfrac{24}{60} \times 100$ (3) $\dfrac{60}{24} \times 100$ (4) $\dfrac{60}{12} \times 100$

5. What is the empirical formula of C_3H_6?
   (1) CH   (2) CH_2 (3) CH_3 (4) CH_6

6. The name of the compound $KClO_2$ is potassium
   (1) hypochlorite   (3) chlorate
   (2) chlorite   (4) perchlorate

7. Which formula is correct for ammonium sulfate?
   (1) NH_4SO_4 (2) $(NH_4)_2SO_4$ (3) $NH_4(SO_4)_2$ (4) $(NH)_3(SO_4)_2$

8. The molecular formula of a compound is represented by X_3Y_6. What is the empirical formula of this compound?
   (1) X_3Y (2) X_2Y (3) XY_2 (4) XY

9. The number of moles of molecules in a 12.0-gram samples of Cl_2 is
   (1) $\dfrac{12.0}{35.5}$ mole   (2) $\dfrac{12.0}{71.0}$ mole   (3) 12.0 moles   (4) 12.0 x 35.5 moles

10. What is the total number of moles of sulfur atoms in 1 mole of $Fe_2(SO_4)_3$?
    (1) 1   (2) 15   (3) 3   (4) 17

11. Given the unbalanced equation:

    $$\_\_\_\ CaSO_4 + \_\_\_\ AlCl_3 \rightarrow \_\_\_\ Al_2(SO_4)_3 + \_\_\_\ CaCl_2$$

    What is the coefficient of $Al_2(SO_4)_3$ when the equation is completely balanced using the smallest whole-number coefficients?
    (1) 1   (2) 2   (3) 3   (4) 4

12. Given the unbalanced equation:

$$\_\_\_ Al\ (s)\ +\ \_\_\_ O_2\ (g) \rightarrow \_\_\_ Al_2O_3\ (s)$$
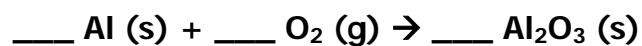
When this equation is correctly balanced using smallest whole numbers, what is the coefficient of O_2 (g)?
(1) 6   (2) 2   (3) 3   (4) 4

13. Given the reaction:

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

What is the total number of moles of NO produced when 1.0 mole of O_2 is completely consumed?
(1) 1.0 mole   (2) 1.2 moles   (3) 0.80 mole   (4) 4.0 moles

14. Given the equation:

$$H_2\ (g)\ +\ Cl_2\ (g) \rightarrow 2\ HCl\ (g)$$
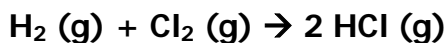
What is the total number of moles of HCl (g) produced when 3 moles of H_2 (g) is completely consumed?
(1) 5 moles   (2) 2 moles   (3) 3 moles   (4) 6 moles

Formulas, Equations & Stoichiometry Review –
Cut from Jan 2007 – Jan 2008 Exams

1. Which equation shows conservation of atoms?
   (1) $H_2 + O_2 \rightarrow H_2O$
   (2) $H_2 + O_2 \rightarrow 2\ H_2O$
   (3) $2\ H_2 + O_2 \rightarrow 2\ H_2O$
   (4) $2\ H_2 + 2\ O_2 \rightarrow 2\ H_2O$

2. Which substance can be broken down by a chemical change?
   (1) antimony   (3) hexane
   (2) carbon   (4) sulfur

3. What is the gram formula mass of $Ca_3(PO_4)_2$?
   (1) 248 g/mol   (3) 279 g/mol
   (2) 263 g/mol   (4) 310 g/mol

4. In which compound is the ratio of metal ions to nonmetal ions 1 to 2?
   (1) calcium bromide
   (2) calcium oxide
   (3) calcium phosphide
   (4) calcium sulfide
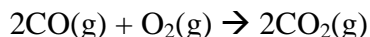
5. Given the balanced equation representing a reaction:
   $$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$
   What is the mole ratio of CO(g) to $CO_2(g)$ in this reaction?
   (1) 1:1   (3) 2:1
   (2) 1:2   (4) 3:2
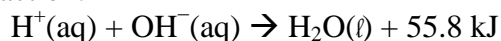
6. Given the balanced equation representing a reaction:
   $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) + 55.8$ kJ
   In this reaction there is conservation of
   (1) mass, only
   (2) mass and charge, only
   (3) mass and energy, only
   (4) mass, charge, and energy

7. Which polyatomic ion contains the greatest number of oxygen atoms?
   (1) acetone   (3) hydroxide
   (2) carbonate   (4) peroxide

8. Which formula represents an ionic compound?
   (1) H_2 (3) CH_3OH
   (2) CH_4 (4) NH_4Cl

9. What is the total number of different elements present in NH_4NO_3?
   (1) 7   (3) 3
   (2) 9   (4) 4

10. Which formula represents lead (II) chromate?
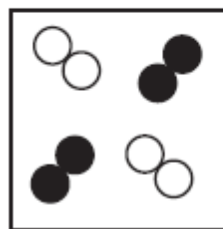   (1) $PbCrO_4$ (3) Pb_2CrO_4
   (2) $Pb(CrO_4)_2$ (4) $Pb_2(CrO_4)_3$
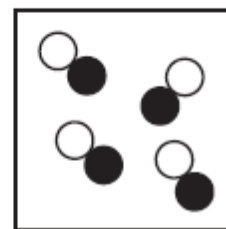
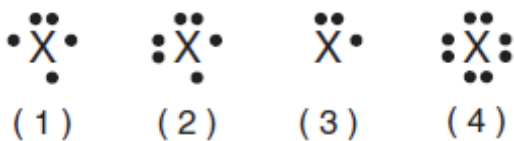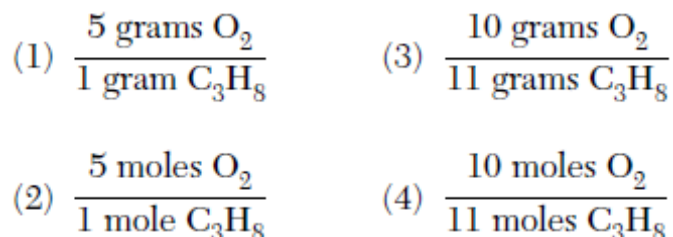11. Which particle diagram represents a sample of one compound, only?
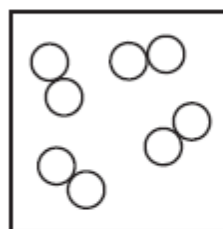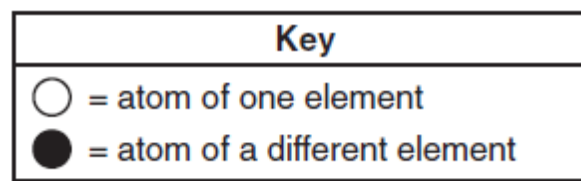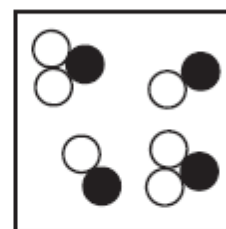
12. An atom in the ground state contains a total of 5 electrons, 5 protons, and 5 neutrons. Which Lewis electron-dot diagram represents this atom?

$\cdot \ddot{X} \cdot$ $: \ddot{X} \cdot$ $\dot{X} \cdot$ $: \ddot{X} :$

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

13. Given the balanced equation representing the reaction between propane and oxygen:
$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$
According to this equation, which ratio of oxygen to propane is correct?

(1) $\dfrac{5 \text{ grams } O_2}{1 \text{ gram } C_3H_8}$

(2) $\dfrac{5 \text{ moles } O_2}{1 \text{ mole } C_3H_8}$

(3) $\dfrac{10 \text{ grams } O_2}{11 \text{ grams } C_3H_8}$

(4) $\dfrac{10 \text{ moles } O_2}{11 \text{ moles } C_3H_8}$

14. Which substance can be decomposed by chemical means?
    (1) tungsten   (3) krypton
    (2) antimony   (4) methane

15. Given the balanced equation representing a reaction:
$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$
What is the *minimum* number of moles of O_2 that are needed to completely react with 16 moles of NH_3?
    (1) 16 mol   (3) 64 mol
    (2) 20. mol   (4) 80. mol

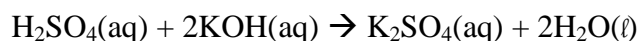16. Element X reacts with iron to form two different compounds with the formulas FeX and Fe_2X_3.
To which group on the Periodic Table does element X belong?
    (1) Group 8   (3) Group 13
    (2) Group 2   (4) Group 16

17. The molar mass of $Ba(OH)_2$ is
    (1) 154.3 g   (3) 171.3 g
    (2) 155.3 g   (4) 308.6 g

18. Given the balanced equation representing a reaction:
$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(\ell)$$
Which type of reaction is represented by this equation?
    (1) decomposition
    (2) neutralization
    (3) single replacement
    (4) synthesis

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

Calculate the percent error of the student's experimental result. Your response must include *both* a correct numerical setup and the calculated result. [2]

20. Write the empirical formula for the compound C_8H_{18}. [1]   _____

Base your answers to questions 21 through 23 on the information below.

Some dry chemicals can be used to put out forest fires. One of these chemicals is $NaHCO_3$. When $NaHCO_3(s)$ is heated, one of the products is $CO_2(g)$, as shown in the balanced equation below.

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

21. Show a correct numerical setup for calculating the percent composition by mass of carbon in the product Na_2CO_3. [1]

22. Identify whether the reaction is endothermic or exothermic. [1] _____

23. Determine the total number of moles of $CO_2(g)$ produced when 7.0 moles of $NaHCO_3(s)$ is completely reacted. [1]

_____ moles

24. Balance this chemical equation: [1]

_____ S(s) + _____ $KClO_3(s) \rightarrow$ _____$SO_2(g) +$ _____ KCl(s) + energy
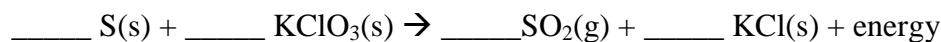
Base your answers to questions 25 through 27 on the information below.

Rust on an automobile door contains $Fe_2O_3(s)$. The balanced equation representing one of the reactions between iron in the door of the automobile and oxygen in the atmosphere is given below.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$
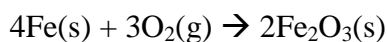
25. Identify the type of chemical reaction represented by this equation. [1]_____

26. Determine the gram-formula mass of the product of this reaction. [1]

27. Write the IUPAC name for Fe_2O_3. [1] _____

Base your answers to questions 28 through 30 on the information below.

Ozone gas, O_3, can be used to kill adult insects in storage bins for grain without damaging the grain. The ozone is produced from oxygen gas, O_2, in portable ozone generators located near the storage bins. The concentrations of ozone used are so low that they do not cause any environmental damage. This use of ozone is safer and more environmentally friendly than a method that used bromomethane, CH_3Br. However, bromomethane was more effective than ozone because CH_3Br killed immature insects as well as adult insects.

Adapted From: *The Sunday Gazette* (Schenectady, NY) 3/9/03

28. Determine the total number of moles of CH3Br in 19 grams of CH3Br (gram-formula mass = 95 grams/mol). [1]

29. Given the balanced equation for producing bromomethane:

$$Br_2 + CH_4 \rightarrow CH_3Br + HBr$$

Identify the type of organic reaction shown. [1] _____

30. Based on the information in the passage, state *one* advantage of using ozone instead of bromomethane for insect control in grain storage bins. [1]

Base your answers to questions 31 through 33 on the information below.

A hydrate is a compound that has water molecules within its crystal structure. The formula for the hydrate $CuSO_4 \cdot 5H_2O(s)$ shows that there are five moles of water for every one mole of $CuSO_4(s)$. When $CuSO_4 \cdot 5H_2O(s)$ is heated, the water within the crystals is released, as represented by the balanced equation below.

$$CuSO_4 \cdot 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$$

A student first masses an empty crucible (a heat-resistant container). The student then masses the crucible containing a sample of $CuSO_4 \cdot 5H_2O(s)$. The student repeatedly heats and masses the crucible and its contents until the mass is constant. The student's recorded experimental data and calculations are shown below.

Data and calculation before heating:

mass of $CuSO_4 \cdot 5H_2O(s)$ and crucible	21.37 g
$-$ mass of crucible	19.24 g
mass of $CuSO_4 \cdot 5H_2O(s)$	2.13 g

Data and calculation after heating to a constant mass:

mass of $CuSO_4(s)$ and crucible	20.61 g
$-$ mass of crucible	19.24 g
mass of $CuSO_4(s)$	1.37 g

Calculation to determine the mass of water:

mass of $CuSO_4 \cdot 5H_2O(s)$	2.13 g
$-$ mass of $CuSO_4(s)$	1.37 g
mass of $H_2O(g)$	0.76 g

31. Identify the total number of significant figures recorded in the calculated mass of $CuSO_4 \cdot 5H_2O(s)$. [1]

32. In the space below, use the student's data to show a correct numerical setup for calculating the percent composition by mass of water in the hydrate. [1]

33. Explain why the sample in the crucible must be heated until the constant mass is reached. [1]

Topic 5: Bonding Outline

1. Chemical compounds are formed when atoms are bonded together.
✓ Breaking a chemical bond is an <u>endothermic</u> process.
✓ Forming a chemical bond is an <u>exothermic</u> process.
✓ Compounds have <u>less</u> potential energy than the individual atoms they are formed from.

2. Two major categories of compounds are ionic and molecular (covalent) compounds.
✓ Ionic compounds tend to be a metal bonding with a nonmetal; or a metal with a polyatomic ion
✓ Molecular (covalent) compounds tend to be two or more nonmetals combined.

3. Compounds can be differentiated by their chemical and physical properties.
✓ Ionic substances have high melting and boiling points, form crystals, dissolve in water (<u>dissociate</u>), and conduct electricity in solution and as liquids.
✓ Covalent or molecular substances have lower melting and boiling points, do not conduct electricity.

4. Atoms gain a stable electron configuration by bonding with other atoms.
✓ Atoms are stable when they have a full valence level.
✓ Most atoms need <u>8 electrons</u> to fill their valence level.
✓ H and He only need <u>2 electrons</u> to fill their valence level.
✓ The noble gases (group 18) have filled valence levels. They do not normally bond with other atoms.

5. Chemical bonds are formed when valence electrons are:
✓ Transferred from one atom to another – <u>ionic.</u>
✓ Shared between atoms – <u>covalent</u>.
✓ Mobile in a free moving "sea" of electrons – <u>metallic</u>.

6. In multiple (double or triple) covalent bonds more than 1 pair of electrons are shared between two atoms.
✓ oxygen and it's family (group 16) form double bonds with each other (O_2)
✓ nitrogen and it's family (group 15) form triple bonds with each other (NH_3)
✓ carbon can form double and triple bonds with itself & group 16 and 15 elements (ex: CO_2)

7. Polarity of a molecule can be determined by its shape and the distribution of the charge.
- ✓ Polar molecules have an asymmetrical (uneven) distribution of electrons in them.
- ✓ As a result, polar molecules have (+) and (-) charged ends.
- ✓ Water is the most common substance composed of polar molecules; O end is (-), H ends are (+).
- ✓ Nonpolar molecules have symmetrical (even) distribution of electrons in them.
- ✓ Polar substances are dissolved only by another polar substance. Non-polar substances are dissolved only by other non-polar substances.

8. The electronegativity difference between two bonded atoms can determine the type of bond and its polarity.
0.0 = non-polar covalent
0.0 -1.7 = polar covalent
1.7+ = ionic

9. Bonding guidelines:
- ✓ Metals react with nonmetals to form ionic compounds.
- ✓ Nonmetals bond with nonmetals to form covalent compounds (molecules).
- ✓ Ionic compounds with polyatomic ions have both ionic and covalent bonds.

10. Intermolecular forces allow different particles to be attracted to each other to form solids and liquids.
- ✓ Hydrogen bonds are an example of a strong IMF between polar molecules.
- ✓ Hydrogen bonds exist between atoms of hydrogen on one molecule and atoms of either oxygen, fluorine, or nitrogen on a neighboring molecule.
- ✓ Substances with hydrogen bonds tend to have much higher melting and boiling points than those without hydrogen bonds. Water is one such substance
- ✓ Ordinary polar molecules simply attract each other as their oppositely charged ends line up.
- ✓ Nonpolar molecules use weak VanderWaal's forces of attraction and as a result tend to have lower melting points, and higher vapor pressures.

11. Metallic bonding occurs between atoms of metal. The valence electrons are loosely held by all atoms in a mobile "sea" of valence electrons.
- ✓ This type of bonding accounts for some of the unique properties of metals, such as their ability to conduct electricity, luster, and malleability.

12. Physical properties of a substance can be explained in terms of chemical bonds and intermolecular forces. These include conductivity, malleability, solubility, ductility, hardness, melting point and boiling point, vapor pressure.

Bonding – Practice Questions

1. The forces between atoms that create chemical bonds are the result of interactions between
   (1) nuclei
   (2) electrons
   (3) protons and electrons
   (4) protons and nuclei

2. According to Reference Table S, which sequence correctly places the elements in order of increasing ionization energy?
   (1) H → Li → Na → K
   (2) I → Br → Cl → F
   (3) O → S → Se → Te
   (4) H → Be → Al → Ga

3. Electronegativity is a measure of an atom's ability to
   (1) attract the electrons in the bond between the atom and another atom
   (2) repel the electrons in the bond between the atom and another atom
   (3) attract the protons of another atom
   (4) repel the protons of another atom

4. If the electronegativity difference between the elements in compound NaX is 2.0, what is element X?
   (1) bromine   (2) chlorine   (3) fluorine   (4) oxygen

5. An element with an electronegativity of 0.9 bonds with an element with an electronegativity of 3.1. Which phrase best describes the bond between these elements?
   (1) mostly ionic in character and formed between two nonmetals
   (2) mostly ionic in character and formed between a metal and a nonmetal
   (3) mostly covalent in character and formed between two nonmetals
   (4) mostly covalent in character and formed between a metal and a nonmetal

6. Which type of bond exists between an atom of carbon and an atom of fluorine?
   (1) ionic   (2) metallic   (3) polar covalent   (4) nonpolar covalent

7. Which pair of atoms is held together by a covalent bond?
   (1) HCl   (2) LiCl   (3) NaCl   (4) KCl

8. Which substance contains nonpolar covalent bonds?
   (1) H_2 (2) H_2O (3) $Ca(OH)_2$ (4) CaO

9. Given the reaction: **Cl (g) + Cl (g) → Cl$_2$ (g) + energy**
   Which statement best describes the reaction?
   (1) A bond is formed and energy is absorbed.
   (2) A bond is formed and energy is released.
   (3) A bond is broken and energy is absorbed.
   (4) A bond is broken and energy is released.

10. The primary forces of attraction between water molecules in H_2O (l) are
    (1) ionic bonds
    (2) hydrogen bonds
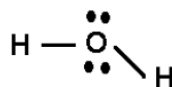   (3) molecule-ion attractions
    (4) van der Waals forces

11. Which structure represents a polar molecule?

(1) H – H

(3)
```
    H
    |
H—C—H
    |
    H
```
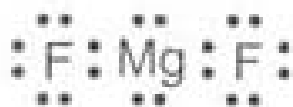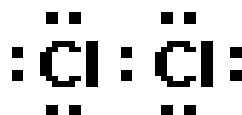
(2) H—C≡C—H

(4)
```
       ••
H—O
       ••  ＼H
```
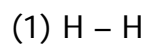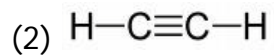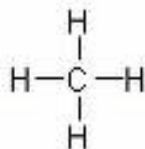
12. Which electron dot diagram represents a molecule that has a polar covalent bond?

1) $:\!\ddot{F}\!:\!\ddot{Mg}\!:\!\ddot{F}\!:$
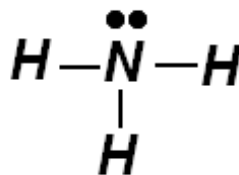
2) $:\!\ddot{Cl}\!:\!\ddot{Cl}\!:$

3) $:\!N\!\equiv\!N\!:$

4)
```
      ••
H—N—H
      |
      H
```

Bonding Review – Cut from Jan 2007 – Jan 2008 Exams

1. Given the balanced equation:

$$I + I \rightarrow I_2$$
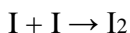
Which statement describes the process represented by this equation?
(1) A bond is formed as energy is absorbed.
(2) A bond is formed and energy is released.
(3) A bond is broken as energy is absorbed.
(4) A bond is broken and energy is released.

2. An oxygen molecule contains a double bond because the two atoms of oxygen share a total of
      (1) 1 electron   (3) 3 electrons
      (2) 2 electrons   (4) 4 electrons

3. A double carbon-carbon bond is found in a molecule of
      (1) pentane   (3) pentyne
      (2) pentene   (4) pentanol

4. At STP, fluorine is a gas and bromine is a liquid because, compared to fluorine, bromine has
      (1) stronger covalent bonds
      (2) stronger intermolecular forces
      (3) weaker covalent bonds
      (4) weaker intermolecular forces

5. Which term indicates how strongly an atom attracts the electrons in a chemical bond?
      (1) alkalinity
      (2) atomic mass
      (3) electronegativity
      (4) activation energy

6. Magnesium nitrate contains chemical bonds that are
      (1) covalent, only
      (2) ionic, only
      (3) both covalent and ionic
      (4) neither covalent nor ionic

7. A solid substance is an excellent conductor of electricity. The chemical bonds in this substance are most likely
(1) ionic, because the valence electrons are shared between atoms
(2) ionic, because the valence electrons are mobile
(3) metallic, because the valence electrons are stationary
(4) metallic, because the valence electrons are mobile

8. When sodium and fluorine combine to produce the compound NaF, the ions formed have the same electron configuration as atoms of
      (1) argon, only
      (2) neon, only
      (3) both argon and neon
      (4) neither argon nor neon

9. Atoms of which element have the greatest tendency to gain electrons?
      (1) bromine   (3) fluorine
      (2) chlorine   (4) iodine

10. Which polyatomic ion contains the greatest number of oxygen atoms?
      (1) acetate   (3) hydroxide
      (2) carbonate   (4) peroxide

11. Which formula represents an ionic compound?
      (1) H_2 (3) CH_3OH
      (2) CH_4 (4) NH_4Cl

12. Which liquid has the highest vapor pressure at 75°C?
      (1) ethanoic acid   (3) propanone
      (2) ethanol   (4) water

13. Given the balanced equation representing a reaction:
$$Cl_2(g) \rightarrow Cl(g) + Cl(g)$$
What occurs during this change?
(1) Energy is absorbed and a bond is broken.
(2) Energy is absorbed and a bond is formed.
(3) Energy is released and a bond is broken.
(4) Energy is released and a bond is formed.

14. At standard pressure, a certain compound has a low boiling point and is insoluble in water. At STP, this compound most likely exists as
      (1) ionic crystals
      (2) metallic crystals
      (3) nonpolar molecules
      (4) polar molecules

15. Which group on the Periodic Table of the Elements contains elements that react with oxygen to form compounds with the general formula X_2O?

      (1) Group 1   (3) Group 14
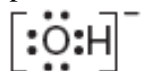   (2) Group 2   (4) Group 18

16. Which two substances are covalent compounds?
      (1) $C_6H_{12}O_6$(s) and KI(s)
      (2) $C_6H_{12}O_6$(s) and HCl(g)
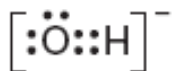   (3) KI(s) and NaCl(s)
      (4) NaCl(s) and HCl(g)

17. Which compound has hydrogen bonding between its molecules?
      (1) CH_4 (3) KH
      (2) CaH_2 (4) NH_3

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

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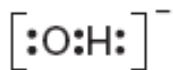
      (2)   (4)

19. Explain, in terms of electronegativity, why a P–Cl bond in a molecule of PCl5 is more polar than a P–S bond in a molecule of P2S5. [1]

Base your answers to questions 20 and 21 on the information below.

The graph below shows the relationship between boiling point and molar mass at standard pressure for pentane, hexane, heptane, and nonane.

Boiling Point Versus Molar Mass of Some Alkanes
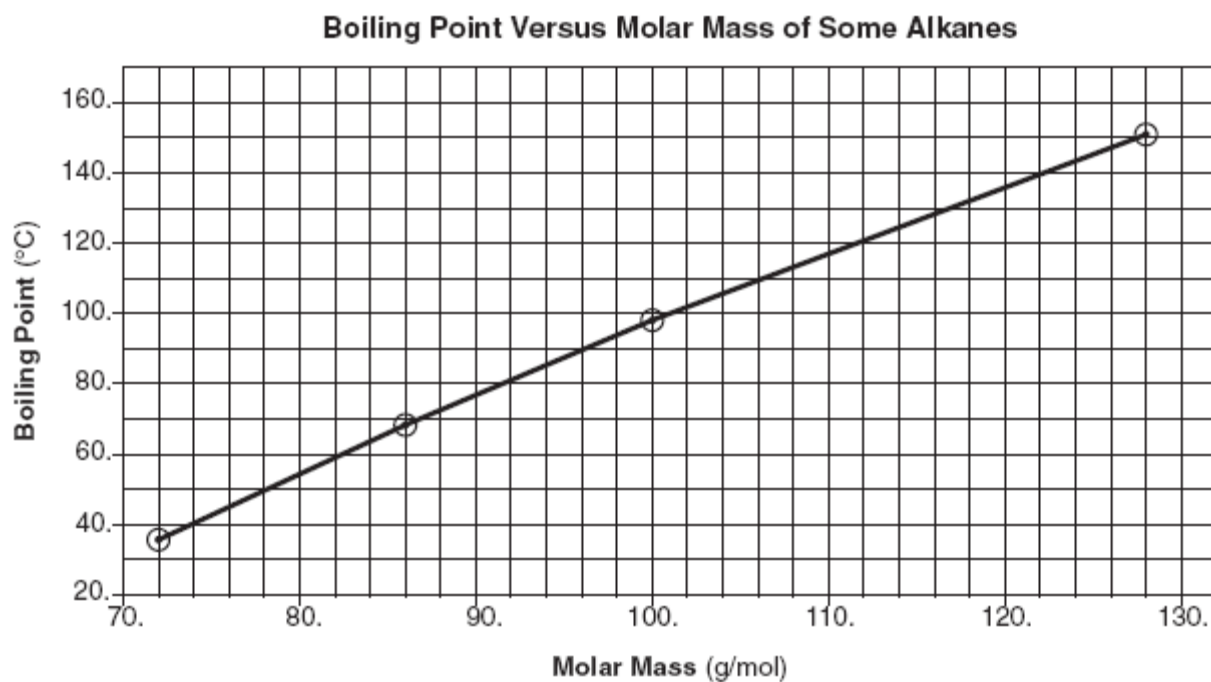
20. Octane has a molar mass of 114 grams per mole. According to this graph, what is the boiling point of octane at standard pressure? [1] _____

21. State the relationship between molar mass and the strength of intermolecular forces for the selected alkanes. [1]

Base your answers to questions 22 through 24 on the information below.

The particle diagrams below represent the reaction between two nonmetals, A_2 and Q_2.

Key
● = Atom of element A
○ = Atom of element Q
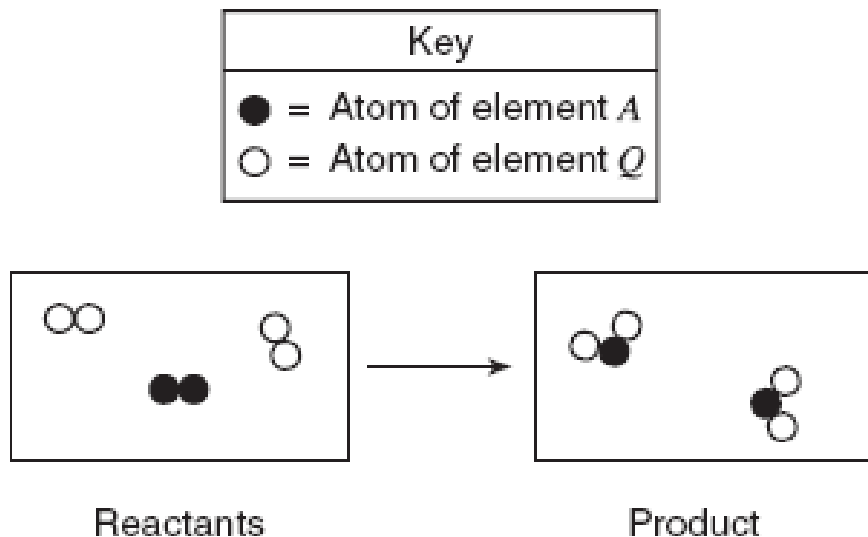
Reactants   Product

22. Using the symbols A and Q, write the chemical formula of the product. [1] _____

23. Identify the type of chemical bond between an atom of element A and an atom of element Q. [1]

24. Compare the total mass of the reactants to the total mass of the product. [1]

25. Explain, in terms of molecular structure or distribution of charge, why a molecule of methane is nonpolar. [1]

26. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Using Reference Table H, determine the boiling point of water when the atmospheric pressure is 90. kPa. [1]

Base your answers to questions 27 through 30 on the information below.

Have you ever seen an insect called a water strider "skating" across the surface of a calm pond? Have you ever "floated" a sewing needle on the water in a glass? If you have, then you've observed one of water's many amazing properties. Water's surface tension keeps the water strider and the sewing needle from sinking into the water. Simply stated, the surface tension is due to the forces that hold the water molecules together. Without these intermolecular forces, the water strider and the sewing needle would sink below the surface of the water. The surface tension of water at various temperatures is given in the data table below.

Surface Tension at Different Water Temperatures

Water Temperature (°C)	Surface Tension (mN/m)
10.	74.2
25	72.0
50.	67.9
75	63.6
100.	58.9

27. On a piece of graph paper, plot the data from the data table. Circle and connect the five points. [1]

28. According to your graph, what is the surface tension of water at 60.°C? [1] _____ mN/m

29. State the relationship between the surface tension and the temperature of water. [1]

30. The surface tension of liquid tetrachloromethane, CCl_4, at 25°C is 26.3 millinewtons/ meter (mN/m). Compare the intermolecular forces between molecules of CCl_4 to the intermolecular forces between molecules of water, H_2O, at 25°C. [1]

Topic 6 Overview

Topic 6A: Heat & Temperature

1. Energy can exist in different forms – chemical, electrical, electromagnetic, thermal, mechanical, nuclear.

✓ Stored energy is referred to as <u>potential energy</u>.
✓ Energy of motion is <u>kinetic energy.</u>

2. The Law of Conservation of Energy states that energy can not be lost or destroyed, only changed from one form to another.

3. Heat is a transfer of energy (often but not always thermal energy) from a body of higher temperature to a body of lower temperature.

4. Temperature is a <u>measure</u> of the average kinetic energy of the particles in a sample. <u>Temperature is NOT a form of energy</u> and should not be confused with heat.

5. The concepts of kinetic and potential energy can be used to explain physical processes such as fusion (melting), solidification (freezing), vaporization (boiling, evaporation), condensation, sublimation, and deposition.

6. Processes that are <u>exothermic</u> give off heat energy. This typically causes the surrounding environment to become warmer.

7. Processes that are <u>endothermic</u> absorb energy. This typically causes the surrounding environment to become colder.

Topic 6A: Heat & Temperature Outline

1. Temperature is a "measure of the average kinetic energy of the particles in a sample of matter."
- ✓ Kinetic energy is energy due to motion. So as temperature increases, the particles move faster, on average.
- ✓ Temperature does NOT depend on the mass of the sample.

2. Temperature scales used by chemists are the Celsius and Kelvin scales.
- ✓ The freezing point of water is a reference point often used in science, and is referred to as "standard temperature." Its value is 0^oC or 273 K, and is noted on Table A.
- ✓ The boiling point of water is 100^oC or 373 K.
- ✓ Converting from C^o to K: $K = C^o + 273$ (on Table T)

3. Heat is a form of energy and IS NOT the same as temperature.
- ✓ Heat is dependent on mass. There is more heat in an iceberg that is at 0^0C than a cup full of boiling water.
- ✓ Heat can be transferred from one substance to another when their particles are in contact (when the objects touch). Heat will move from the object with more particle KE (higher temp) to the one with less.
- ✓ The amount of heat needed to cause a temperature change is dependent on the mass of the sample, its "specific heat" and the amount of temperature change: $q = m c \Delta T$ (Table T)   When heat is absorbed to cause a temperature change, it is resulting in a change in KE of particles.
- ✓ The amount of heat needed to cause a phase change can be calculated using the $q = mHf$ (melting), or $q = mHv$ (boiling)   (Table T).   When heat is added to cause a phase change, it is causing a change in intermolecular forces between particles.
- ✓ The values for water are on Table B.

4. Heat of fusion (H_f) is the energy needed to convert one gram of a substance from solid to liquid.

5. Heat of vaporization (H_v) is the energy needed to convert one gram of a substance from liquid to gas.

6. Specific heat (C) is the energy required to raise one gram of a substance 1 degree (Celcius or Kelvin).
- ✓ The specific heat of liquid water is 1 cal/g*J or 4.2 J/g*K.

7. The three phases of matter are solid, liquid and gas. Each has its own properties.
- ✓ Solids have a constant volume and shape. Particles are held in a rigid, crystalline structure.
- ✓ Liquids have a constant volume but a changing shape. Particles are mobile but still held together by strong attraction.
- ✓ Gases have no set volume or shape. They will completely fill any closed contained. Particles have largely broken free of the forces holding them together.
- ✓ The phase a substance is in is dependent on the temperature. Melting points and boiling points are on Table S (in Kelvin degrees).

8. Phase changes are a type of physical change. If they are changes that involve heat being absorbed, they are endothermic changes.
- ✓ Endothermic phase changes are melting, boiling, evaporating and subliming (s→l).
- ✓ Opposite type of phase changes (freezing, condensing, depositing) are exothermic.

9. A heating curve (or cooling curve) traces the changes in temperature of a substance as it changes from solid to liquid to gas (or gas to liquid to solid).
- ✓ When the substance undergoes a phase change, there is no change in temperature. The line "flattens" until the phase change is complete.
- ✓ When a phase change is occurring, the potential energy of the substance changes while kinetic energy remains the same.
- ✓ As temperature increases, kinetic energy increases.

10. The amount of heat involved in some chemical changes is shown on Table I, called "heat of reaction" or ΔH.
- ✓ If the value is negative, the reaction is exothermic.
- ✓ This can be expressed as a potential energy diagram.
- ✓ If the energy is written into the equation, and is on the reactants side, the reaction is endothermic.
- ✓ ΔH is the difference between the energy stored in the products (PE) and the potential energy of the reactants.

11. Breaking bonds is ALWAYS endothermic, and forming bonds is ALWAYS exothermic.
- ✓ I + I → I_2 Bond is forming, I atoms are become stable by bonding, so they release energy (Exo)
- ✓ H_2 → H + H Bond is breaking, requires energy in order to put atoms in unbonded state (endo)

Heat and Temperature – Cut from Jan 2007 – Jan 2008 Exams
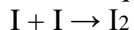
1. Given the balanced equation:
$$I + I \rightarrow I_2$$
Which statement describes the process represented by this equation?
(1) A bond is formed as energy is absorbed.
(2) A bond is formed and energy is released.
(3) A bond is broken as energy is absorbed.
(4) A bond is broken and energy is released.

2. Which term is defined as a measure of the average kinetic energy of the particles in a sample?
      (1) temperature   (3) thermal energy
      (2) pressure   (4) chemical energy

3. Which term refers to the difference between the potential energy of the products and the potential energy of the reactants for any chemical change?
(1) heat of deposition
(2) heat of fusion
(3) heat of reaction
(4) heat of vaporization

4. Which kelvin temperature is equal to 56°C?
      (1) –329 K   (3) 217 K
      (2) –217 K   (4) 329 K

5. Which reaction releases the greatest amount of energy per 2 moles of product?
(1) $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$
(2) $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
(3) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$
(4) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Use the reaction shown below to answer questions 6 and 7.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(\ell) + 2219.2 \text{ kJ}$$

6. Draw a potential energy diagram for this reaction. [1]

7. Determine the total amount of energy released when 2.50 moles of propane is completely reacted with oxygen. [1]

8. Given the balanced equation representing a reaction: $N_2(g) + O_2(g) + 182.6 \text{ kJ} \rightarrow 2NO(g)$
Draw a potential energy diagram for this reaction. [1]

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

A 5.00-gram sample of liquid ammonia is originally at 210. K. The diagram of the partial heating curve below represents the vaporization of the sample of ammonia at standard pressure due to the addition of heat. The heat is *not* added at a constant rate.

Partial Heating Curve for Ammonia
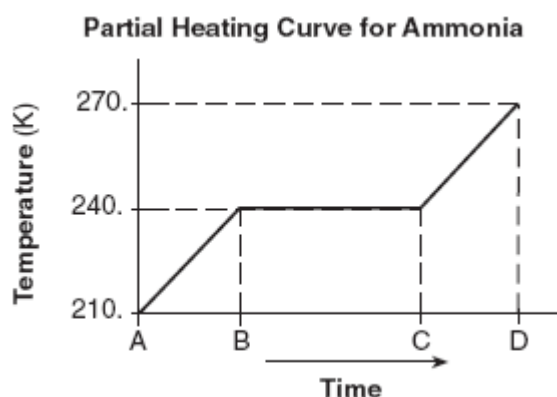
Some physical constants for ammonia are shown in the data table below.

Some Physical Constants for Ammonia

specific heat capacity of $NH_3(\ell)$	4.71 J/g•K
heat of fusion	332 J/g
heat of vaporization	1370 J/g

9. Calculate the total heat absorbed by the 5.00-gram sample of ammonia during time interval *AB*. Your response must include *both* a correct numerical setup and the calculated result. [2]

10. Describe what is happening to *both* the potential energy and the average kinetic energy of the molecules in the ammonia sample during time interval *BC*. Your response must include *both* potential energy and average kinetic energy. [1]

11. Determine the total amount of heat required to vaporize this 5.00-gram sample of ammonia at its boiling point. [1]

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Base your answers to questions 12 through 14 on the information below.

A 100.0-gram sample of NaCl(s) has an initial temperature of 0°C. A chemist measures the temperature of the sample as it is heated. Heat is *not* added at a constant rate. The heating curve for the sample is shown below.

Heating Curve for Sodium Chloride
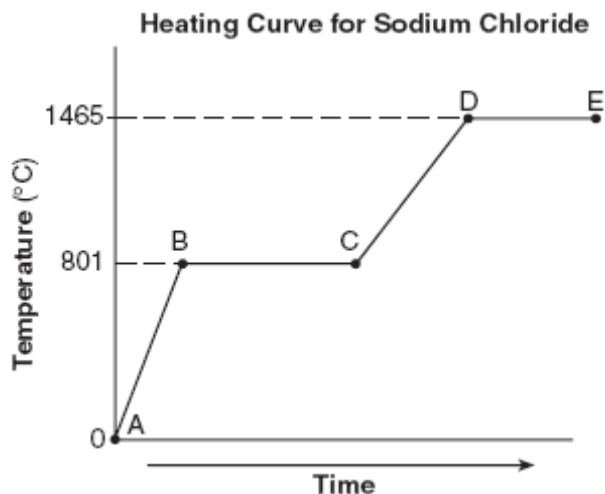
12. Determine the temperature range over which the entire NaCl sample is a liquid. [1]

13. Identify *one* line segment on the curve where the average kinetic energy of the particles of the NaCl sample is changing. [1]

14. Identify *one* line segment on the curve where the NaCl sample is in a single phase and capable of conducting electricity. [1]

Base your answers to questions 15 and 16 on the information below.

 A student performed an experiment to determine the total amount of energy stored in a peanut. The accepted value for the energy content of a peanut is 30.2 kilojoules per gram. The student measured 100.0 grams of water into a metal can and placed the can on a ring stand, as shown in the diagram below. The peanut was attached to a wire suspended under the can. The initial temperature of the water was recorded as 22.0°C. The peanut was ignited and allowed to burn. When the peanut finished burning, the final water temperature was recorded as 57.0°C. The student's experimental value for the energy content of this peanut was 25.9 kilojoules per gram.
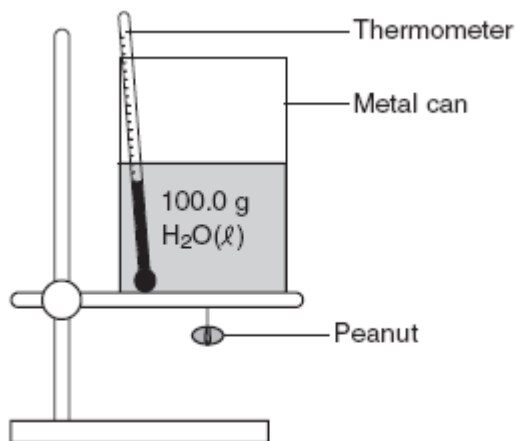
15. Calculate the total amount of heat absorbed by the water. Your response must include *both* a correct numerical setup and the calculated result. [2]

16. Determine the student's percent error for the energy content of this peanut. [1]

Base your answers to questions 17 through 20 on the information below.

The temperature of a sample of a substance is increased from 20.°C to 160.°C as the sample absorbs heat at a constant rate of 15 kilojoules per minute at standard pressure. The graph below represents the relationship between temperature and time as the sample is heated.

Temperature Versus Time
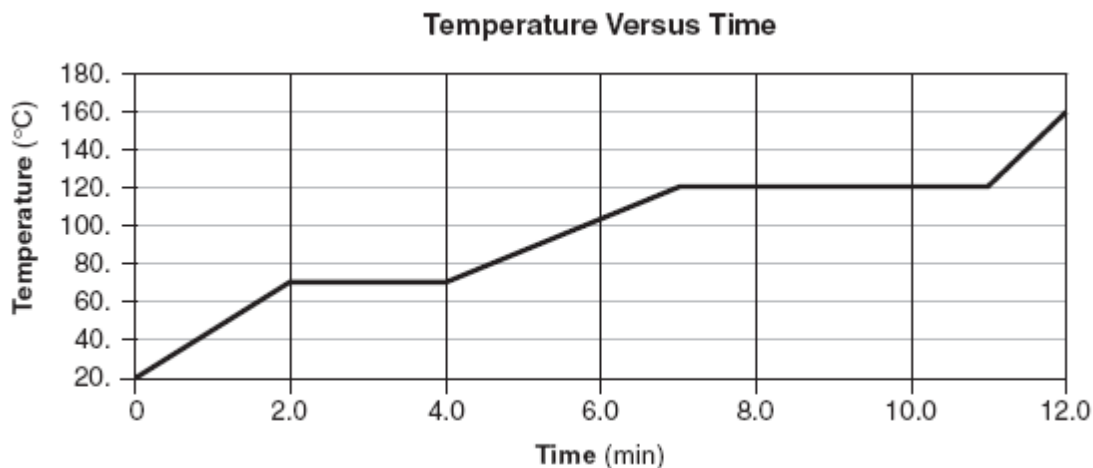
17. What is the boiling point of this sample? [1]

18. Draw *at least nine* particles in the box, showing the correct particle arrangement of this sample during the first minute of heating. [1]

19. What is the total time this sample is in the liquid phase, only? [1]

20. Determine the total amount of heat required to completely melt this sample at its melting point. [1]

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Topic 6B: Reaction Rate & Equilibrium Outline

1. Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation.
  - ✓ This is sometimes called an "effective collision."

2. The rate of a chemical reaction depends on temperature, concentration, nature of the reactants, surface area and the presence of a catalyst.

3. Energy absorbed or released by a chemical reaction can be represented by a potential energy diagram.

4. The amount of energy released or absorbed during a chemical reaction is the heat of reaction.
  - ✓ Heat of reaction equals the PE of the products – PE of reactants.
  - ✓ Positive heat of reaction implies an endothermic reaction.
  - ✓ Negative heat of reaction implies an exothermic reaction.

5. A catalyst provides an alternative pathway for a chemical reaction. The catalyst lowers reaction the activation energy required to start up the reaction.
  - ✓ Adding a catalyst increases the rate of the forward and reverse reactions equally, so there is no shift in equilibrium.
  - ✓ Know how the use of a catalyst affects the PE diagram.

6. Entropy is a measure of the randomness or disorder in a system. A system with greater disorder has greater entropy.

7. Systems in nature tend to undergo changes towards lower energy (tend to be exothermic) and higher entropy.

8. At equilibrium the rate of the forward reaction equals the rate of the reverse reaction.
  - ✓ This state can only be achieved IF the system (container) is closed and the conditions of Temp and Pressure are held steady.

9. The measurable quantities of reactants and products remain constant at equilibrium.
  - ✓ This does NOT mean the amounts of products and reactants is the same as each other, but rather that the amounts are no longer changing.

**10. Types of equilibrium include chemical, phase and solution.**
  - ✓ Solutions that are saturated represent an equilibrium between the processes of dissolving and precipitating.
  - ✓ An example of a phase equilibrium would be the simultaneous melting and freezing of water if the system is held at $0^{o}C$.

11. LeChatelier's principle can be used to predict the effect of stress on a system in equilibrium.
  - ✓ Stresses include a change in pressure, volume, concentration, and temperature.
  - ✓ You should be able to predict if a "shift left" or a "shift right" occurs due to a particular stress.

Rate of Reaction & Equilibrium –
Cut from Jan 2007 – Jan 2008 Exams

1. Given the equation representing a phase change at equilibrium:

$$C_2H_5OH(\ell) \rightleftharpoons C_2H_5OH(g)$$
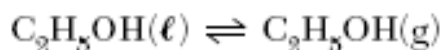
Which statement is true?
(1) The forward process proceeds faster than the reverse process.
(2) The reverse process proceeds faster than the forward process.
(3) The forward and reverse processes proceed at the same rate.
(4) The forward and reverse processes both stop.

2. A 5.0-gram sample of zinc and a 50.-milliliter sample of hydrochloric acid are used in a chemical reaction. Which combination of these samples has the fastest reaction rate?
(1) a zinc strip and 1.0 M HCl(aq)
(2) a zinc strip and 3.0 M HCl(aq)
(3) zinc powder and 1.0 M HCl(aq)
(4) zinc powder and 3.0 M HCl(aq)

3. For a given reaction, adding a catalyst increases the rate of the reaction by
(1) providing an alternate reaction pathway that has a higher activation energy
(2) providing an alternate reaction pathway that has a lower activation energy
(3) using the same reaction pathway and increasing the activation energy
(4) using the same reaction pathway and decreasing the activation energy

4. Given the equation representing a reaction at equilibrium:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + energy$$
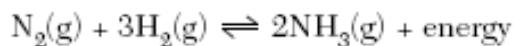
Which change causes the equilibrium to shift to the right?
(1) decreasing the concentration of $H_2(g)$
(2) decreasing the pressure
(3) increasing the concentration of $N_2(g)$
(4) increasing the temperature

5. Given the equation representing a system at equilibrium:

$$H_2O(s) \rightleftharpoons H_2O(\ell)$$

At which temperature does this equilibrium exist at 101.3 kilopascals?
(1) 0 K   (3) 32 K
(2) 0°C   (4) 273°C

6. Which statement must be true when solution equilibrium occurs?
(1) The solution is at STP.
(2) The solution is supersaturated.
(3) The concentration of the solution remains constant.
(4) The masses of the dissolved solute and the undissolved solute are equal.

7. Which statement must be true for any chemical reaction at equilibrium?
(1) The concentration of the products is greater than the concentration of the reactants.
(2) The concentration of the products is less than the concentration of the reactants.
(3) The concentration of the products and the concentration of the reactants are equal.
(4) The concentration of the products and the concentration of the reactants are constant.

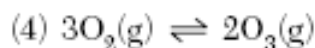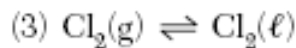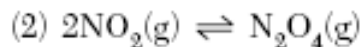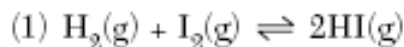8. Given the balanced equation representing a reaction at 101.3 kPa and 298 K:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) + 91.8 \text{ kJ}$$
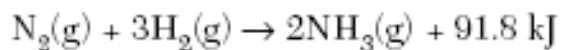
Which statement is true about this reaction?
(1) It is exothermic and $\otimes H$ equals –91.8 kJ.
(2) It is exothermic and $\otimes H$ equals +91.8 kJ.
(3) It is endothermic and $\otimes H$ equals –91.8 kJ.
(4) It is endothermic and $\otimes H$ equals +91.8 kJ.

9. Which balanced equation represents a phase

equilibrium?

(1) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

(2) $2NO_2(g) \rightleftharpoons N_2O_4(g)$

(3) $Cl_2(g) \rightleftharpoons Cl_2(\ell)$

(4) $3O_2(g) \rightleftharpoons 2O_3(g)$

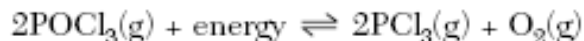11. In terms of energy and entropy, systems in nature tend to undergo changes toward
(1) higher energy and higher entropy
(2) higher energy and lower entropy
(3) lower energy and higher entropy
(4) lower energy and lower entropy

10. Given the system at equilibrium:

$2POCl_3(g) + energy \rightleftharpoons 2PCl_3(g) + O_2(g)$

Which changes occur when $O_2(g)$ is added to this system?
(1) The equilibrium shifts to the right and the concentration of $PCl_3(g)$ increases.
(2) The equilibrium shifts to the right and the concentration of $PCl_3(g)$ decreases.
(3) The equilibrium shifts to the left and the concentration of $PCl_3(g)$ increases.
(4) The equilibrium shifts to the left and the concentration of $PCl_3(g)$ decreases.

12. Explain, in terms of collision theory, why the rate of a chemical reaction increases with an increase in temperature. [1]

Base your answers to questions 13 through 15 on the information below.

A beaker contains 100.0 milliliters of a dilute aqueous solution of ethanoic acid at equilibrium. The equation below represents this system.

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

13. Compare the rate of the forward reaction to the rate of the reverse reaction for this system. [1]

14. Describe what happens to the concentration of $H^+(aq)$ when 10 drops of concentrated $HC_2H_3O_2(aq)$ are added to this system. [1]

15. Draw a structural formula for ethanoic acid. [1]

Base your answer to question 16 on the information below.

"Hand Blasters" is a toy that consists of a set of two ceramic balls, each coated with a mixture of sulfur and potassium chlorate, $KClO_3$. When the two balls are struck together, a loud popping noise is produced as sulfur and potassium chlorate react with each other.

16. Identify *one* source of the activation energy for this reaction. [1]

Base your answers to questions 17 through 18 on the reaction represented by the balanced equation below.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) + 571.6 \text{ kJ}$$
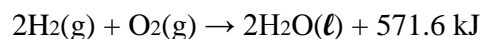
17. Identify the information in this equation that indicates the reaction is exothermic. [1]

18. Explain why the entropy of the system *decreases* as the reaction proceeds. [1]

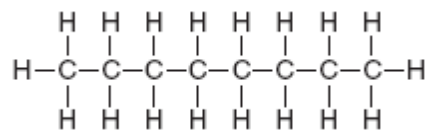Base your answers to question 19 on the information below.

The equilibrium equation below is related to the manufacture of a bleaching solution. In this equation, $Cl^-(aq)$ means that chloride ions are surrounded by water molecules.

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

19. Explain, in terms of collision theory, why increasing the concentration of $Cl_2(g)$ increases the concentration of $OCl^-(aq)$ in this equilibrium system. [1]

Base your answers to questions 20 and 21 on the information below.

A gasoline engine burns gasoline in the presence of excess oxygen to form carbon dioxide and water. The main components of gasoline are isomers of octane. A structural formula of octane is shown below.

$$\begin{array}{c} \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \\ | \ \ | \ \ | \ \ | \ \ | \ \ | \ \ | \ \ | \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ | \ \ | \ \ | \ \ | \ \ | \ \ | \ \ | \ \ | \\ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \ \text{H} \end{array}$$

One isomer of octane is 2,2,4-trimethylpentane.

20. In the space *in your answer booklet*, draw a structural formula for 2,2,4-trimethylpentane. [1]

21. Explain, in terms of the arrangement of particles, why the entropy of gasoline vapor is greater than the entropy of liquid gasoline. [1]

Topic 7: Water & Solutions

1. **Water has some unusual properties.**
   - ✓ The bonds between H and O inside a water molecule are "polar covalent."
   - ✓ Due to its structure, it is a polar molecule. This means it has an uneven distribution of electrons in it. The O end is (-) and the H ends are (+).
   - ✓ *Draw its Lewis dot structure here!*　　　　_____
   - ✓ Water is actually a VERY polar substance. As a result, it uses the very strong type of intermolecular (between molecule) forces of attraction called "HYDROGEN BONDS."
   - ✓ As a result of H-bonding, water has an unusually high melting point and boiling point compared with similar molecules like H_2S.
   - ✓ Water has a very high specific heat, so it heats up and cools down much more slowly than most materials. This value is found on Table B.
   - ✓ Water solutions that contain ions are capable of conducting electricity. The substances that form the ions in solution are called "electrolytes."

2. **Water is able to make solutions with many substances.**
   - ✓ Solutions are ALWAYS HOMOGENEOUS MIXTURES.
   - ✓ Water will dissolve many ionic compounds.
   - ✓ Water will dissolve molecular substances if they are also polar.
   - ✓ This reminds us of the "like dissolves like" principle.
   - ✓ Acids dissolve in water to form H^+ ions. (This includes organic acids:　R-COOH)
   - ✓ Bases dissolve in water to form OH^- ions. (This does NOT include alcohols: R-OH)

3. **Ionic compounds may be either soluble or insoluble in water.**
   - ✓ Use Table F to decide!

4. **Solubility describes how much of a particular solute will dissolve in a set amount of water at a certain temperature.**
   - ✓ Use Table G. The amount of water used is 100 g.
   - ✓ Saturated solutions hold all the solute possible at the temperature chosen for the water.
   - ✓ An increase in temperature of the water usually makes it capable of dissolving more solute. The opposite is true for gas solutes like O_2 gas, or NH_3 or SO_2 or HCl.
   - ✓ An increase in pressure over the solution increases the solubility of gas solutes. It does not affect solubility of solutes that are liquids or solids.

5. **Solutions have a lower freezing point and a higher boiling point that pure water. This effect becomes larger with more concentrated solutions.**

6. **"Concentration" describes how much solute is dissolved in a certain amount of water.**
   - ✓ You should know how to calculate:
     - o Molarity
     - o % mass
     - o Parts per million
       - • Use Table T and "plug and chug."

Water & Solutions –
Cut from Jan 2007 – Jan 2008 Exams

1. A 3.0 M HCl(aq) solution contains a total of
   (1) 3.0 grams of HCl per liter of water
   (2) 3.0 grams of HCl per mole of solution
   (3) 3.0 moles of HCl per liter of solution
   (4) 3.0 moles of HCl per mole of water

2. A dilute, aqueous potassium nitrate solution is best classified as a
   (1) homogeneous compound
   (2) homogeneous mixture
   (3) heterogeneous compound
   (4) heterogeneous mixture

3. According to one acid-base theory, a water molecule acts as an acid when the water molecule
   (1) accepts an H^+ (3) donates an H^+
   (2) accepts an OH^- (4) donates an OH^-

4. An Arrhenius base yields which ion as the only negative ion in an aqueous solution?
   (1) hydride ion   (3) hydronium ion
   (2) hydrogen ion   (4) hydroxide ion

5. Which barium salt is *insoluble* in water?
   (1) $BaCO_3$ (3) $Ba(ClO_4)_2$
   (2) $BaCl_2$ (4) $Ba(NO_3)_2$

6. Which unit can be used to express solution concentration?
   (1) J/mol   (3) mol/L
   (2) L/mol   (4) mol/s

7. Under which conditions of temperature and pressure is a gas most soluble in water?
   (1) high temperature and low pressure
   (2) high temperature and high pressure
   (3) low temperature and low pressure
   (4) low temperature and high pressure

8. Given the equation representing a system at equilibrium:

   $$H_2O(s) \rightleftharpoons H_2O(\ell)$$

   At which temperature does this equilibrium exist at 101.3 kilopascals?
   (1) 0 K   (3) 32 K
   (2) 0°C   (4) 273°C

9. As water is added to a 0.10 M NaCl aqueous solution, the conductivity of the resulting solution
   (1) decreases because the concentration of ions decreases
   (2) decreases, but the concentration of ions remains the same
   (3) increases because the concentration of ions decreases
   (4) increases, but the concentration of ions remains the same

10. Which substance is an Arrhenius acid?
    (1) $Ba(OH)_2$ (3) H_3PO_4
    (2) CH_3COOCH_3 (4) NaCl

11. Which compound releases hydroxide ions in an aqueous solution?
    (1) CH_3COOH (3) HCl
    (2) CH_3OH (4) KOH

12. Which liquid has the highest vapor pressure at 75°C?
    (1) ethanoic acid   (3) propanone
    (2) ethanol   (4) water

13. Which sample of matter is a single substance?
    (1) air   (3) hydrochloric acid
    (2) ammonia gas   (4) salt water

14. At standard pressure, a certain compound has a low boiling point and is insoluble in water. At STP, this compound most likely exists as
    (1) ionic crystals
    (2) metallic crystals
    (3) nonpolar molecules
    (4) polar molecules

15. An unsaturated solution is formed when 80. grams of a salt is dissolved in 100. grams of water at 40.°C. This salt could be
    (1) KCl   (3) NaCl
    (2) KNO_3 (4) $NaNO_3$

16. Which substance, when dissolved in water, forms a solution that conducts an electric current?
    (1) C_2H_5OH (3) $C_{12}H_{22}O_{11}$
    (2) $C_6H_{12}O_6$ (4) CH_3COOH

17. Compared to a 2.0 M aqueous solution of NaCl at 1 atmosphere, a 3.0 M aqueous solution of NaCl at 1 atmosphere has a
(1) lower boiling point and a higher freezing point
(2) lower boiling point and a lower freezing point
(3) higher boiling point and a higher freezing point
(4) higher boiling point and a lower freezing point

18. A student prepares four aqueous solutions, each with a different solute. The mass of each dissolved solute is shown in the table below.

Mass of Dissolved Solute for Four Aqueous Solutions

Solution Number	Solute	Mass of Dissolved Solute (per 100. g of H_2O at 20.°C)
1	KI	120. g
2	$NaNO_3$	88 g
3	KCl	25 g
4	$KClO_3$	5 g

Which solution is saturated?
(1) 1   (3) 3
(2) 2   (4) 4

Base your answers to question 19 on the information below.

The equilibrium equation below is related to the manufacture of a bleaching solution. In this equation, $Cl^-(aq)$ means that chloride ions are surrounded by water molecules.

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

19. Draw *two* water molecules in the box, showing the correct orientation of each water molecule toward the chloride ion. [1]

Base your answers to questions 20 through 22 on the information below.

Scientists who study aquatic ecosystems are often interested in the concentration of dissolved oxygen in water. Oxygen, O_2, has a very low solubility in water, and therefore its solubility is usually expressed in units of milligrams per 1000. grams of water at 1.0 atmosphere. The graph below shows a solubility curve of oxygen in water.

Solubility of Oxygen in Water Versus Temperature
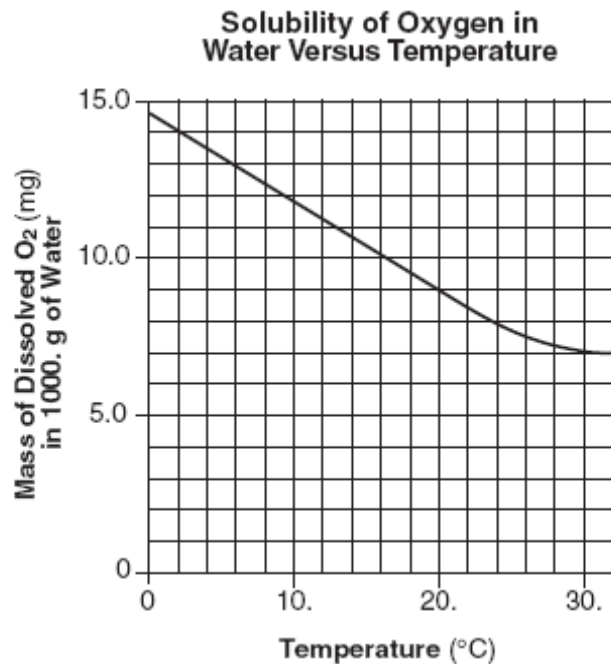
20. A student determines that 8.2 milligrams of oxygen is dissolved in a 1000.-gram sample of water at 15°C and 1.0 atmosphere. In terms of saturation, what type of solution is this sample? [1] _____

21. Explain, in terms of molecular polarity, why oxygen gas has low solubility in water. Your response must include *both* oxygen and water. [1]

22. An aqueous solution has 0.0070 gram of oxygen dissolved in 1000. grams of water. Calculate the dissolved oxygen concentration of this solution in parts per million. Your response must include *both* a correct numerical setup and the calculated result. [2]

Base your answers to questions 23 and 24 on the information below.

A solution is made by completely dissolving 90. grams of $KNO_3(s)$ in 100. grams of water in a beaker. The temperature of this solution is 65°C.

23. Describe the effect on the solubility of $KNO_3(s)$ in this solution when the pressure on the solution increases. [1]

24. Determine the total mass of $KNO_3(s)$ that settles to the bottom of the beaker when the original solution is cooled to 15°C. [1]

Base your answers to questions 25 through 27 on the information below.

The compound 1,2-ethanediol can be mixed with water. This mixture is added to automobile radiators as an engine coolant. The cooling system of a small van contains 6690 grams of 1,2-ethanediol. Some properties of water and 1,2-ethanediol are given in the table below.

Properties of Water and 1,2-ethanediol

Property	Water (H_2O)	1,2-ethanediol (CH_2OHCH_2OH)
gram-formula mass (g/mol)	18.0	62.0
boiling point at standard pressure (°C)	100.0	197.2

25. Identify the class of organic compounds to which 1,2-ethanediol belongs. [1] _____

26. State, in terms of molecular polarity, why 1,2-ethanediol is soluble in water. [1]

27. Calculate the total number of moles of 1, 2-ethanediol in the small van's cooling system. Your response must include *both* a correct numerical setup and the calculated result. [2]

28. An aqueous solution contains 300. parts per million of KOH. Determine the number of grams of KOH present in 1000. grams of this solution. [1]

29. A liquid boils when the vapor pressure of the liquid equals the atmospheric pressure on the surface of the liquid. Using Reference Table H, determine the boiling point of water when the atmospheric pressure is 90. kPa. [1]

Topic 8: Acids & Bases Outline

1. An electrolyte is a substance which when dissolved in water forms a solution capable of conducting an electric current. The ability of a solution to conduct an electric current depends upon the concentration of *ions* present.
  ✓ Ionic compounds are conductors of electricity when melted OR dissolved in water. Under these circumstances, the charged particles (ions in this case) are free to move (mobile).
  ✓ There are 3 categories of electrolytes: acids, bases and salts.
  ✓ Arrhenius theory states that an **acid** is an substance that dissolves in water to **produce H$^+$** (H$_3$O$^+$) ions (called "hydronium" ions on Table E).
  ✓ Arrhenius theory states that a **base** is an substance that dissolves in water to **produce OH$^-$** ions (called "hydroxide" ions on Table E).
  ✓ A salt is any ionic compound producing a positive ion other than H$^+$ and a negative ion other than OH$^-$.
  ✓ Common acid and base names and formulas are given on Tables K and L.
  ✓ *You should be able to sort compounds as acids, bases or salts, given their chemical formulas*

2. Properties of many acids and bases can be explained by the Arrhenius theory. Arrhenius acids and bases are electrolytes.
  ✓ Acid properties include sour taste, less than 7 pH, ability to neutralize bases, and ability to affect indicator colors. These properties are due to the H$^+$ ion.
  ✓ Base properties include bitter taste, greater than 7 pH, ability to neutralize acids, and ability to affect indicator colors. These properties are due to the OH$^-$ ion.
  ✓ *When given properties, you should be able to identify substances as Arrhenius acids or Arrhenius bases*

3. The acidity or alkalinity of a solution can be measured by its pH value.
  ✓ For every change in pH of one unit, the acidity changes by a factor of 10. A pH 4 solution is 10 times more acidic that a pH 5 solution. A pH 4 solution is 100 times more acidic that a pH 6 solution.
  ✓ *You should be able to identify solutions as acid, base, or neutral based upon the pH. Neutral is a pH of 7.*

4. The relative level of acidity or alkalinity of a solution can be shown by using indicators.
  ✓ Various indicators are shown on Table M. Make sure you know how to interpret the info on it!
  ✓ Example: IF bromothymol blue is yellow in color, we know the pH is 6 OR LESS. If its color is blue, we know the pH of the solution is 7.6 OR GREATER.

5. In the process of neutralization, an Arrhenius acid reacts with an Arrhenius base to form a salt and water.
  ✓ Example: Ba(OH)$_2$ + 2 HBr → BaBr$_2$ +   2 H$_2$O
  ✓ These reactions are double replacements.
  ✓ These reactions are NOT redox reactions.
  ✓ *You should be able to write simple neutralization reactions when given the reactants*

6. Titration is a laboratory process in which the volume of a solution of known concentration is used to determine the concentration of another solution.
  ✓ *You should be able to calculate the concentration or volume of a solution, using titration data*
  ✓ In order to do this, use the titration equation on Table T: (MxV) $_{acid}$ = (MxV) $_{base}$

7. The Bronsted acid - base theory views acids as "H$^+$ donors", and bases as "H$^+$ acceptors."

Acids & Bases – Cut from Jan 2007 – Jan 2008 Exams

1. An Arrhenius base yields which ion as the only negative ion in an aqueous solution?
    (1) hydride ion (3) hydronium ion
    (2) hydrogen ion   (4) hydroxide ion

2. According to one acid-base theory, a water molecule acts as an acid when the water molecule
    (1) accepts an H_+ (3) donates an H_+
    (2) accepts an OH_- (4) donates an OH_-

3. Which two formulas represent Arrhenius acids?
    (1) CH_3COOH and CH_3CH_2OH
    (2) $HC_2H_3O_2$ and H_3PO_4
    (3) $KHCO_3$ and $KHSO_4$
    (4) $NaSCN$ and $Na_2S_2O_3$

4. Which substance is an Arrhenius acid?
    (1) $Ba(OH)_2$ (3) H_3PO_4
    (2) CH_3COOCH_3 (4) $NaCl$

5. Which compound releases hydroxide ions in an aqueous solution?
    (1) CH_3COOH (3) HCl
    (2) CH_3OH (4) KOH

6. What are the products of a reaction between $KOH(aq)$ and $HCl(aq)$?
    (1) H_2 and $KClO$ (3) KH and $HClO$
    (2) H_2O and KCl (4) KOH and HCl

7. Which volume of 0.10 M $NaOH(aq)$ exactly neutralizes 15.0 milliliters of 0.20 M $HNO_3(aq)$?
    (1) 1.5 mL   (3) 3.0 mL
    (2) 7.5 mL   (4) 30. mL

8. Which indicator, when added to a solution, changes color from yellow to blue as the pH of the solution is changed from 5.5 to 8.0?
    (1) bromcresol green   (3) litmus
    (2) bromthymol blue   (4) methyl orange

9. What is the pH of a solution that has a hydronium ion concentration 100 times greater than a solution with a pH of 4?
    (1) 5   (3) 3
    (2) 2   (4) 6

10. The pH of an aqueous solution changes from 4 to 3 when the hydrogen ion concentration in the solution is
(1) decreased by a factor of 3/4
(2) decreased by a factor of 10
(3) increased by a factor of 4/3
(4) increased by a factor of 10

11. Which formula represents a hydronium ion?
    (1) H_3O_+ (3) OH_-
    (2) NH_{4+} (4) HCO_{3-}

12. Which compound is an Arrhenius acid?
    (1) H_2SO_4 (3) $NaOH$
    (2) KCl (4) NH_3

13. The table below shows the color of the indicators methyl orange and litmus in two samples of the same solution.

Results of Acid-Base Indicator Tests

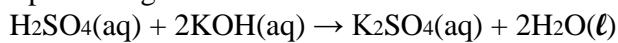| Indicator	Color Result from the Indicator Test
methyl orange	yellow
litmus	red

Which pH value is consistent with the indicator results?
    (1) 1   (3) 3
    (2) 5   (4) 10

14. Which ion is the only negative ion produced by an Arrhenius base in water?
    (1) NO_{3-} (3) OH_-
    (2) Cl_- (4) H_-

15. Given the balanced equation representing a reaction:
$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(\ell)$$
Which type of reaction is represented by this equation?
    (1) decomposition   (2) neutralization   (3) single replacement   (4) synthesis

16. In which 0.01 M solution is phenolphthalein pink?
    (1) $CH_3OH(aq)$ (2) $Ca(OH)_2(aq)$ (3) $CH_3COOH(aq)$ (4) $HNO_3(aq)$

17. As the pH of a solution is changed from 3 to 6, the concentration of hydronium ions
    (1) increases by a factor of 3
    (2) increases by a factor of 1000
    (3) decreases by a factor of 3
    (4) decreases by a factor of 1000

18. What color is bromcresol green after it is added to a sample of NaOH(aq)? [1]

19. Identify *two* indicators from Reference Table *M* that are yellow in solutions with a pH of 5.5. [1]

Base your answers to questions 20 and 21 on the information below.

    Sulfur dioxide, SO_2, is one gas produced when fossil fuels are burned. When this gas reacts with water in the atmosphere, an acid is produced forming acid rain. The pH of the water in a lake changes when acid rain collects in the lake.
    Two samples of the same rainwater are tested using two indicators. Methyl orange is yellow in one sample of this rainwater. Litmus is red in the other sample of this rainwater.

20. Identify a possible pH value for the rainwater that was tested. [1] _____

21. Write the formula for *one* substance that can neutralize the lake water affected by acid rain. [1]

Base your answers to questions 22 through 24 on the information below.

A laboratory worker filled a bottle with a hydrochloric acid solution. Another bottle was filled with methanol, while a third bottle was filled with a sodium hydroxide solution. However, the worker neglected to label each bottle. After a few days, the worker could not remember which liquid was in each bottle.

The worker needed to identify the liquid in each bottle. The bottles were labeled *A*, *B*, and *C*. Using materials found in the lab (indicators, conductivity apparatus, and pieces of Mg metal), the worker tested samples of liquid from each bottle. The test results are shown in the table below.

Table of Tests and Results

Test	Test Results		
	Bottle A	Bottle B	Bottle C
methyl orange indicator	yellow	yellow	yellow
bromthymol blue indicator	blue	green	yellow
electrical conductivity	conductor	nonconductor	conductor
reactivity with Mg metal	no reaction	no reaction	reaction

22. Using the test results, state how the worker differentiated the bottle that contained methanol from the other two bottles. [1]

23. The worker concluded that bottle *C* contained hydrochloric acid. Identify *one* test and state the corresponding test result that supports this conclusion. [1]

24. Explain, in terms of pH, why the methyl orange indicator test results were the same for each of the three liquids. [1]

Base your answers to questions 25 through 27 on the information below.

In a laboratory activity, 0.500 mole of NaOH(s) is completely dissolved in distilled water to form 400. milliliters of NaOH(aq). This solution is then used to titrate a solution of HNO_3(aq).

25. Identify the negative ion produced when the NaOH(s) is dissolved in distilled water. [1] _____

26. Calculate the molarity of the NaOH(aq). Your response must include *both* a correct numerical setup and the calculated result. [2]

28. Complete the equation representing this titration reaction by writing the formulas of the products. [1]

HNO_3 + NaOH → _____ + _____

Base your answers to questions 29 through 31 on the information below.

In preparing to titrate an acid with a base, a student puts on goggles and an apron. The student uses burets to dispense and measure the acid and the base in the titration. In each of two trials, a 0.500 M NaOH(aq) solution is added to a flask containing a volume of HCl(aq) solution of unknown concentration. Phenolphthalein is the indicator used in the titration. The calculated volumes used for the two trials are recorded in the table below.

Volumes of Base and Acid Used in Titration Trials

Solution (aq)	Molarity (M)	Trial 1 Volume Used (mL)	Trial 2 Volume Used (mL)
NaOH	0.500	17.03	16.87
HCl	?	10.22	10.12

29. Write a chemical name for the acid used in the titration. [1] _____

30. Using the volumes from trial 1, determine the molarity of the HCl(aq) solution. [1]

31. Based on the information given in the table, how many significant figures should beshown in the calculated molarity of the HCl(aq) solution used in trial 2? [1] _____

Topic 9: Organic Chemistry

1. Organic compounds consist of carbon atoms which bond to each other in chains, rings and networks to form a variety of structures.
- ✓ The source of most hydrocarbons (Table Q) is petroleum, which is a mixture of many hydrocarbons.
- ✓ The hydrocarbons in petroleum are separated from each other by distillation in a "cracking tower," on the basis of boiling points.
- ✓ The greater the molar mass, the higher the intermolecular forces of attraction between molecules. As a result, melting points and boiling points are higher. EX: Octane is a liquid at room temperature, whereas propane (the smaller molecule) is a gas, showing that it has weaker forces of attractions.

2. Organic compounds can be named with the IUPAC system.
- ✓ You should know this system! Use Tables P and Q and R for help!

3. Hydrocarbons are compounds that contain only carbon and hydrogen.
- ✓ Saturated hydrocarbons contain only single carbon-carbon bonds.
- ✓ Unsaturated hydrocarbons contain at least one multiple carbon-carbon bond (double or triple bond).
- ✓ Hydrocarbons tend to be nonpolar molecules, and therefore do not dissolve in water.
- ✓ Hydrocarbons are molecular compounds that do not ionize in water, and are therefore "Non-electrolytes."

4. Organic acids, alcohols, esters, aldehydes, ketones, ethers, halides, amines, amides, and amino acids are categories of organic molecules that differ in their structures.
- ✓ Use Table R for help!
- ✓ Esters are the trickiest to name/draw. Review that one especially!

5. Functional groups give organic molecules distinct physical and chemical properties.

6. Isomers of organic compounds have the same molecular formula but different structures and properties.

7. In a multiple covalent bond, more than one pair of electrons are shared between two atoms. Unsaturated organic compounds contain at least one double or triple bond.

8. Types of organic reactions include: addition, substitution, polymerization, esterification, fermentation, saponification, and combustion.
- ✓ You need to memorize the details of these reactions in order to be able to identify them. The Regents exam likes to go after the esterification reaction especially.

9. Empirical formulas express the simplest ration of elements in a comnpound.
- ✓ EX: Hexane – Molecular formula = C_6H_{14} & has an empirical formula of C_3H_7
  - ▪ Propane – Molecular formula = C_3H_8 but has no simpler formula, so C_3H_8 is also its empirical formula

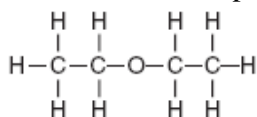# Organic Review – Cut from Jan 2007 – Jan 2008 Exams

1. What is the total number of carbon atoms in a molecule of ethanoic acid?
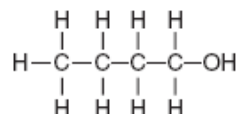
   (1) 1   (3) 3
   (2) 2   (4) 4

2. A double carbon-carbon bond is found in a molecule of

   (1) pentane   (3) pentyne
   (2) pentene   (4) pentanol

3. Given the formulas for two compounds:

and

These compounds differ in
(1) gram-formula mass
(2) molecular formula
(3) percent composition by mass
(4) physical properties at STP

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

   (1) C_2H_2 and CH_3CH_3 (3) P_4O_{10} and P_2O_5
   (2) C_6H_6 and C_2H_2 (4) SO_2 and SO_3

5. The organic compound represented by the condensed structural formula $CH_3CH_2CH_2CHO$ is classified as an

   (1) alcohol   (3) ester
   (2) aldehyde   (4) ether

6. Which two formulas represent Arrhenius acids?
(1) CH_3COOH and CH_3CH_2OH
(2) $HC_2H_3O_2$ and H_3PO_4
(3) $KHCO_3$ and $KHSO_4$
(4) $NaSCN$ and $Na_2S_2O_3$

7. Which compound is an unsaturated hydrocarbon?

   (1) hexanal   (3) hexanoic acid
   (3) hexane   (4) hexene

8. Which formula represents an alkene?

   (1) C_2H_6 (3) C_4H_{10}
   (2) C_3H_6 (4) C_5H_{12}

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

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

10. Which compound is a saturated hydrocarbon?

   (1) CH_2CH_2 (3) CH_3CHO
   (2) CH_3CH_3 (4) CH_3CH_2OH

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

   (1) but-   (3) oct-
   (2) dec-   (4) pent-

12. A molecule of butane and a molecule of 2-butene both have the same total number of

   (1) carbon atoms   (3) single bonds
   (2) hydrogen atoms   (4) double bonds

13. Which general formula represents the homologous series of hydrocarbons that includes the compound l-heptyne?

   (1) C_nH_{2n-6} (3) C_nH_{2n}
   (2) C_nH_{2n-2} (4) C_nH_{2n+2}

14. Which two compounds are isomers of each other?
(1) CH_3CH_2COOH and $CH_3COOCH_2CH_3$
(2) CH_3CH_2CHO and CH_3COCH_3
(3) $CH_3CHBrCH_3$ and $CH_2BrCHBrCH_3$
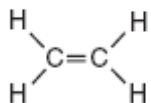(4) $CH_3CHOHCH_3$ and $CH_3CHOHCH_2OH$

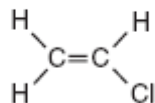15. Which formula represents an unsaturated hydrocarbon?
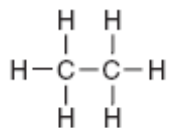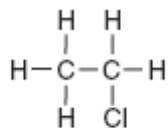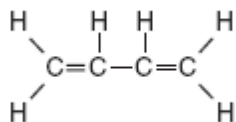

(1)   (3)

(2)   (4)

16. A compound has a molar mass of 90. grams per mole and the empirical formula CH_2O. What is the molecular formula of this compound?
      (1) CH_2O (3) $C_3H_6O_3$
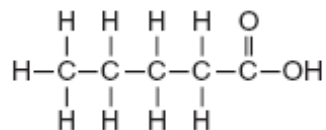   (2) $C_2H_4O_2$ (4) $C_4H_8O_4$

17. Given the formula of a substance:

What is the total number of shared electrons in a molecule of this substance?
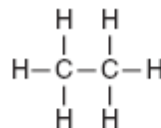   (1) 22   (3) 9
      (2) 11   (4) 6

18. Given the structural formula:

What is the IUPAC name of this compound?
      (1) pentanal   (3) methyl pentanoate
      (2) pentanol   (4) pentanoic acid

19. Which structural formula represents an unsaturated hydrocarbon?
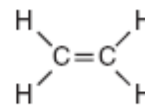
(1)   (3)
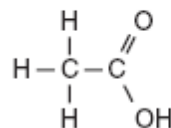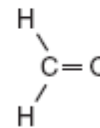
(2)   (4)

20. Two substances have different physical and chemical properties. Both substances have molecules that contain two carbon atoms, one oxygen atom, and six hydrogen atoms. These two substances must be
(1) isomers of each other
(2) isotopes of each other
(3) the same compound
(4) the same hydrocarbon
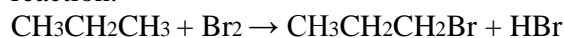
21. Given the balanced equation representing a reaction:
$CH_3CH_2CH_3 + Br_2 \rightarrow CH_3CH_2CH_2Br + HBr$
This organic reaction is best classified as
(1) an addition reaction
(2) an esterification reaction
(3) a polymerization reaction
(4) a substitution reaction

22. Given the structural formula:

$$H-C\equiv C-H$$

What is the total number of electrons shared in the bond between the two carbon atoms?
      (1) 6   (3) 3
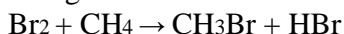   (2) 2   (4) 4

Base your answers to questions 23 and 24 on the information below.

Ozone gas, O_3, can be used to kill adult insects in storage bins for grain without damaging the grain. The ozone is produced from oxygen gas, O_2, in portable ozone generators located near the storage bins. The concentrations of ozone used are so low that they do not cause any environmental damage. This use of ozone is safer and more environmentally friendly than a method that used bromomethane, CH_3Br. However, bromomethane was more effective than ozone because CH_3Br killed immature insects as well as adult insects.

Adapted From: *The Sunday Gazette* (Schenectady, NY) 3/9/03

23. Determine the total number of moles of CH_3Br in 19 grams of CH_3Br (gram-formula mass = 95 grams/mol). [1]

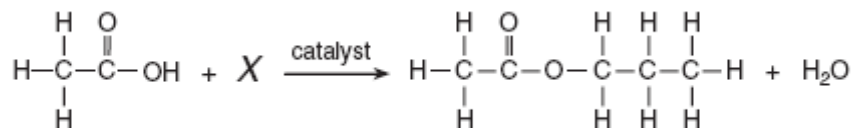24. Given the balanced equation for producing bromomethane:
$$Br_2 + CH_4 \rightarrow CH_3Br + HBr$$
Identify the type of organic reaction shown. [1]

25. Write the empirical formula for the compound C_8H_{18}. _____

Base your answers to questions 26 through 28 on the information below.

The incomplete equation below represents an esterification reaction.
The alcohol reactant is represented by X.

26. On the reaction above, circle the acid functional group.[1]

27. Write an IUPAC name for the reactant represented by its structural formula in this equation. [1]
_____

28. Draw the structural formula for the alcohol represented by X.

29. Explain, in terms of molecular structure or distribution of charge, why a molecule of methane is nonpolar. [1]

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

    The graph below shows the relationship between boiling point and molar mass at standard pressure for pentane, hexane, heptane, and nonane.

Boiling Point Versus Molar Mass of Some Alkanes
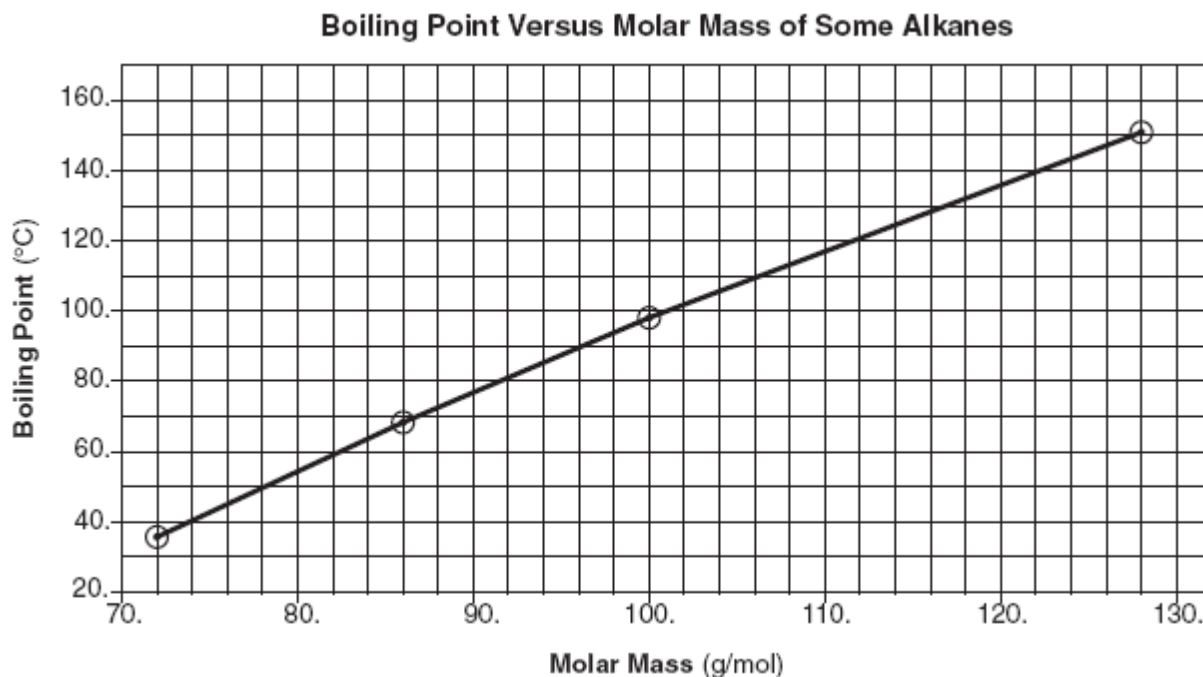
30. Octane has a molar mass of 114 grams per mole. According to this graph, what is the boiling point of octane at standard pressure? [1]

31. State the relationship between molar mass and the strength of intermolecular forces for the selected alkanes. [1]

Topic 10: Phases & Gases

1. The three phases of matter are solid, liquid and gas. Each has its own properties.
✓ Solids have a constant volume and shape. Particles are held in a rigid, crystalline structure.
✓ Liquids have a constant volume but a changing shape. Particles are mobile but still held together by strong attraction.
✓ Gases have no set volume or shape. They will completely fill any closed contained. Particles have largely broken free of the forces holding them together.

2. A heating curve (or cooling curve) traces the changes in temperature of a substance as it changes from solid to liquid to gas (or gas to liquid to solid).
✓ When the substance undergoes a phase change, there is no change in temperature. The line "flattens" until the phase change is complete.
✓ When a phase change is occurring, the potential energy of the substance changes while kinetic energy remains the same.
✓ As temperature increases, kinetic energy increases.

3. <u>Heat of fusion</u> (H_f) is the energy needed to convert one gram of a substance from solid to liquid.

4. <u>Heat of vaporization</u> (H_v) is the energy needed to convert one gram of a substance from liquid to gas.

5. <u>Specific heat</u> (C) is the energy required to raise one gram of a substance 1 degree (Celcius or Kelvin).
✓ The specific heat of liquid water is 1 cal/g*J or 4.2 J/g*K.

6. The combined gas law states the relationship between pressure, temperature and volume in a sample of gas.
✓ Increasing pressure causes a decrease in volume (inverse relationship).
✓ Increasing temperature causes an increase in volume (direct relationship).
✓ Increasing temperature causes an increase in pressure.(direct relationship).
✓ $(PV/T)_1 = (PV/T)_2$ reminder!! All temperatures must be in the Kelvin scale when using this equation

7. An ideal gas model is used to explain the behavior of gases. An "ideal gas" would perfectly obey the combined gas law (PV/T) at all conditions of temperature and pressure.
✓ A real gas is most like an ideal gas when it is at <u>high temperature and low pressure.</u> In other words, at conditions which would favor the gas staying in the gas phase (and not liquefying).
✓ The real gases that most resemble ideal gases are H_2 and He.

8. The Kinetic Molecular Theory (KMT) for an ideal gas states that all gas particles:
✓ are in random motion.
✓ have no forces of attraction between them.
✓ have a negligible volume compared to the distances between them.
✓ have collisions that result in the transfer of energy from one particle to another, with no net loss of energy from the collision.

9. Equal volumes of gases at the same temp and pressure have an equal number of particles.

Phases & Gases – Cut from Jan 2007 – Jan 2008 Exams

1. The boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the pressure on the surface of the liquid. What is the boiling point of propanone if the pressure on its surface is 48 kilopascals?

      (1) 25°C   (3) 35°C

      (2) 30.°C   (4) 40.°C

2. At which Celsius temperature does lead change from a solid to a liquid?

      (1) 874°C   (3) 328°C

      (2) 601°C   (4) 0°C

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

Data for Four Gas Samples

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

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

      (1) *A* and *B*   (3) *B* and *C*

      (2) *A* and *C*   (4) *B* and *D*

4. At which temperature is the vapor pressure of ethanol equal to the vapor pressure of propanone at 35°C?

      (1) 35°C   (3) 82°C

    q(2) 60.°C   (4) 95°C

5. A rigid cylinder with a movable piston contains a 2.0-liter sample of neon gas at STP. What is the volume of this sample when its temperature is increased to 30.°C while its pressure is decreased to 90. kilopascals?

      (1) 2.5 L   (3) 1.6 L

      (2) 2.0 L   (4) 0.22 L

6. Which kelvin temperature is equal to 56°C?

      (1) –329 K   (3) 217 K

      (2) –217 K   (4) 329 K

7. A sample of gas is held at constant pressure. Increasing the kelvin temperature of this gas sample causes the average kinetic energy of its molecules to

(1) decrease and the volume of the gas sample to decrease

(2) decrease and the volume of the gas sample to increase

(3) increase and the volume of the gas sample to decrease

(4) increase and the volume of the gas sample to increase

8. At STP, which sample contains the same number of molecules as 11.2 liters of $CO_2(g)$ at STP?

  (1) 5.6 L of $NO_2(g)$ (3) 11.2 L of $N_2(g)$

 (2) 7.5 L of $H_2(g)$ (4) 22.4 L of CO(g)

9. At which temperature would atoms of a He(g) sample have the greatest average kinetic energy?

      (1) 25°C   (3) 273 K

      (2) 37°C   (4) 298

10. A 1.00-mole sample of neon gas occupies a volume of 24.4 liters at 298 K and 101.3 kilopascals. In the space *in your answer booklet*, calculate the density of this sample. Your response must include *both* a correct numerical setup and the calculated result. [2]

Base your answers to questions 11 through 14 on the information below.

The temperature of a sample of a substance is increased from 20.°C to 160.°C as the sample absorbs heat at a constant rate of 15 kilojoules per minute at standard pressure. The graph below represents the relationship between temperature and time as the sample is heated.

Temperature Versus Time
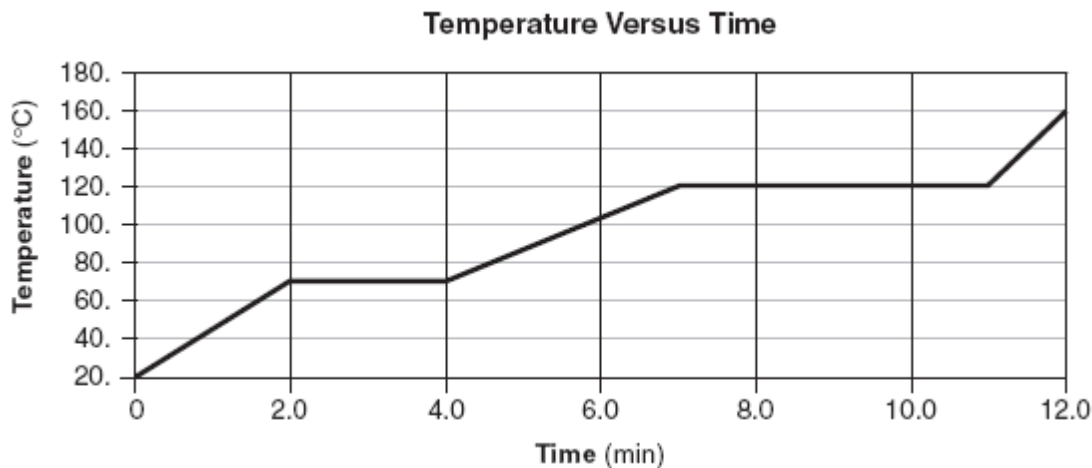
11. What is the boiling point of this sample? [1] _____

12. Draw *at least nine* particles in the box, showing the correct particle arrangement of this sample during the first minute of heating. [1]

13. What is the total time this sample is in the liquid phase, only? [1] _____

14. Determine the total amount of heat required to completely melt this sample at its melting point. [1]

Base your answers to questions 15 through 17 on the information below.

A rigid cylinder is fitted with a movable piston. The cylinder contains a sample of helium gas, He(g), which has an initial volume of 125.0 milliliters and an initial pressure of 1.0 atmosphere, as shown below. The temperature of the helium gas sample is 20.0°C.
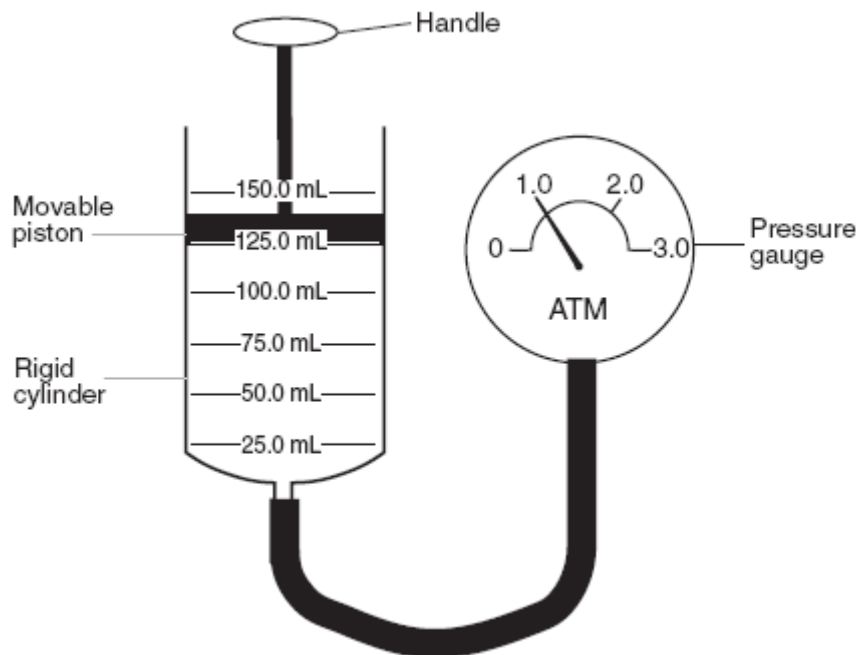
15. Express the initial volume of the helium gas sample, in liters. [1] _____ Liters

16. The piston is pushed further into the cylinder. Show the correct numerical setup for calculating the volume of the helium gas that is anticipated when the reading on the pressure gauge is 1.5 atmospheres. The temperature of the helium gas remains constant. [1]

17. Helium gas is removed from the cylinder and a sample of nitrogen gas, $N_2(g)$, is added to the cylinder. The nitrogen gas has a volume of 125.0 milliliters and a pressure of 1.0 atmosphere at 20.0°C. Compare the number of particles in this nitrogen gas sample to the number of particles in the original helium gas sample. [1]

Topic 11: Electrochemistry (Oxidation-Reduction Reactions) Outline

1. In all chemical reactions there is a conservation of mass, energy, and charge.

2. An oxidation-reduction (redox) reaction involves the transfer of electrons (e-).

3. Reduction is the gain of electrons.

✓ A half-reaction can be written to represent reduction.

For example: $Cl_2 + 2\ e^- \rightarrow 2\ Cl^{1-}$

4. Oxidation is the loss of electrons.

✓ A half-reaction can be written to represent oxidation.

For example: $Na \rightarrow Na^+ + 1\ e^-$

5. Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred.

✓ Be able to use an activity series (Reference Table J) to determine whether a redox reaction is spontaneous.

6. In a redox reaction the number of electrons lost is equal to the number of electrons gained.

✓ This supports the fact that charge is always conserved!

7. An electrochemical cell can be either voltaic or electrolytic. In an electrochemical cell, oxidation occurs at the anode and reduction at the cathode.

✓ Be able to compare and contrast voltaic and electrolytic cells.

8. A voltaic cell spontaneously converts chemical energy to electrical energy.

✓ Identify and label the parts of a voltaic cell (cathode, anode, salt bridge) and direction of electron flow, given the reaction equation.

✓ Since this reaction is spontaneous, use Ref. Table J to help you determine what is the anode (oxidation) and the cathode (reduction). The most reactive metal will oxidize, the most reactive non-metal will reduce.

9. An electrolytic cell requires electrical energy to produce a chemical change. This process is known as electrolysis.

✓ Identify and label the parts of an electrolytic cell (cathode, anode) and direction of electron flow, given the reaction equation

✓ Since this reaction is non-spontaneous, use Ref. Table J to help you determine what is the anode (oxidation) and the cathode (reduction). The most reactive metal will reduce, the most reactive non-metal will oxidize. (NOTE: This is the opposite of what metals/non-metals want to do!)
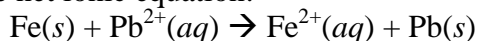
Electrochemistry Practice Questions

1. Given the equation:

$$\_\_Ca^{2+}(aq) + \_\_PO_4^{3-}(aq) \rightarrow \_\_Ca_3(PO_4)_2(s)$$
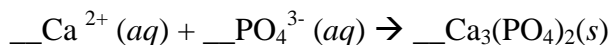
When the equation is correctly balanced, the sum of the total charge of the reactants is

(1) 0   (3) -3
(2) +2   (4) +6

2. The net ionic equation:
$$Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$$
illustrates conservation of

(1) mass and charge
(2) charge but not mass
(3) mass but not charge
(4) neither mass nor charge

3. As an atom of nitrogen gains electrons, its oxidation number

(1) decreases
(2) increases
(3) remains the same

4. When a neutral atom undergoes oxidation, the atom's oxidation state

(1) decreases as it gains electrons
(2) decreases as it loses electrons
(3) increases as it gains electrons
(4) increases as it loses electrons

5. In which substance does hydrogen have an oxidation number of zero?

(1) LiH   (3) H_2S
(2) H_2O (4) H_2

6. What is the oxidation number of chlorine in $HClO_4$?

(1) +1   (3) +3
(2) +5   (4) +7

7. In which substance is the oxidation number of Cl equal to +1?

(1) Cl_2 (3) $AlCl_3$
(2) Cl_2O (4) $HClO_2$

8. What is the oxidation number of chromium in $K_2Cr_2O_7$?

(1) +12   (3) +3
(2) +2   (4) +6

9. Oxygen has an oxidation number of -2 in

(1) O_2 (3) Na_2O_2
(2) NO_2 (4) OF_2

10. In which compound does chlorine have the highest oxidation number?

(1) NaClO   (3) $NaClO_3$
(2) $NaClO_2$ (4) $NaClO_4$

11. In which compound does carbon have an oxidation state of -4?

(1) CO   (3) CCl_4
(2) CO_2 (4) CH_4

12. What is the oxidation number of carbon in $NaHCO_3$?

(1) +6   (3) -4
(2) +2   (4) +4

13. Given the reaction:

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_3(g) + 2H_2O(l)$$

As the reaction occurs, what happens to copper?

(1) It undergoes reduction and its oxidation number decreases.
(2) It undergoes reduction and its oxidation number increases.
(3) It undergoes oxidation and its oxidation number decreases.
(4) It undergoes oxidation and its oxidation number increases.

14. Which component of a voltaic cell is correctly paired with its function?

(1) external conductor — allows the solutions to mix
(2) external conductor — permits the migration of ions
(3) salt bridge — allows the solutions to mix
(4) salt bridge — permits the migration of ions

15. The diagram represents a chemical cell at 298 K.

$$2Al(s) + 3Ni^{2+}(aq) \longrightarrow 2Al^{3+} + 3Ni(s)$$
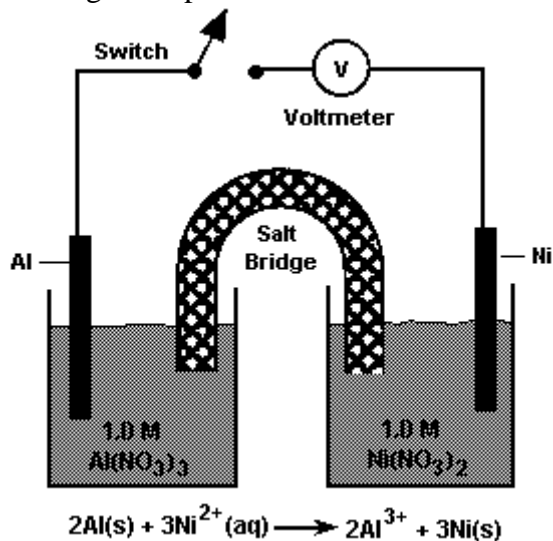
When the switch is closed, electrons flow from
(1) Al(s) to Ni(s)
(2) Ni(s) to Al(s)
(3) $Al^{3+}(aq)$ to $Ni^{2+}(aq)$
(4) $Ni^{2+}(aq)$ to $Al^{3+}(aq)$

16. Which redox equation is correctly balanced?
(1) $Cr^{3+} + Mg \rightarrow Cr + Mg^{2+}$
(2) $Al^{3+} + K \rightarrow Al + K^{+}$
(3) $Sn^{4+} + H_2 \rightarrow Sn + 2H^{+}$
(4) $Br_2 + Hg \rightarrow Hg^{2+} + 2Br^{-}$

17. Which statement best describes how a salt bridge maintains electrical neutrality in the half-cells of a voltaic cell?

(1) It prevents the migration of electrons.
(2) It permits the migration of ions.
(3) It permits the two solutions to mix completely.
(4) It prevents the reaction from occuring spontaneously.

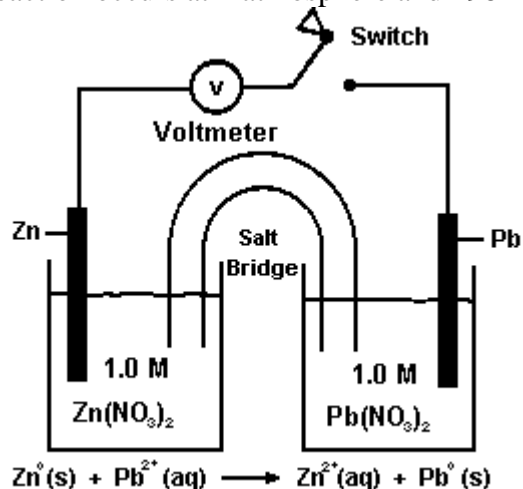18. Given the reaction:

$$\_Hg^{2+} + \_Ag^{0} \rightarrow \_Hg^{0} + \_Ag^{1+}$$
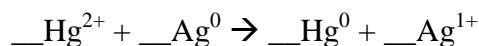
When the equation is completely balanced using the smallest whole-number coefficients, the coefficient of Hg will be

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

19. The diagram shows a voltaic cell. The reaction occurs at 1 atmosphere and 298 K.

$$Zn^{0}(s) + Pb^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Pb^{0}(s)$$

When the switch is closed, what occurs?

(1) Pb is oxidized and electrons flow to the Zn electrode.
(2) Pb is reduced and electrons flow to the Zn electrode.
(3) Zn is oxidized and electrons flow to the Pb electrode.
(4) Zn is reduced and electrons flow to the Pb electrode.

20. Which metal can replace Cr in Cr_2O_3?

(1) nickel   (3) copper
(2) lead   (4) aluminum

21. Given the reaction:

$2Cr(s) + \underline{\quad}Sn^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + \underline{\quad}Sn(s)$
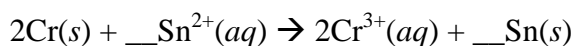
When the reaction is correctly balanced using the smallest whole numbers, the coefficient of $Sn^{2+}(aq)$ is

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

22. An electrochemical cell that generates electricity contains half-cells that produce

(1) oxidation half-reactions, only
(2) reduction half-reactions, only
(3) spontaneous redox reactions
(4) nonspontaneous redox reactions

23. Given the reaction:

$$2Li(s) + Cl_2(g) \rightarrow 2LiCl(s)$$
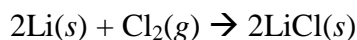
As the reaction takes place, the $Cl_2(g)$ will

(1) gain electrons   (3) gain protons
(2) lose electrons   (4) lose protons

24. In the reaction $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$, the oxidizing agent is

(1) Cu   (3) Ag^+
(2) Cu^{2+} (4) Ag

25. Which procedure requires the use of an external electric current to force a redox reaction to occur?

(1) polymerization
(2) distillation
(3) electrolysis
(4) saponification

26. An electrolytic cell is different from a voltaic cell because in an electrolytic cell

(1) a redox reaction occurs
(2) a spontaneous reaction occurs
(3) an electric current is produced
(4) an electric current causes a chemical reaction

27. The diagram shows an electrolytic cell in which the electrodes are tin and copper.
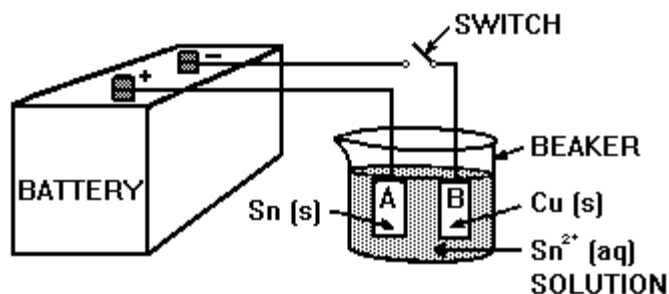
When the switch is closed, what will happen to the two electrodes?

(1) B will dissolve and A will become coated with tin.
(2) A will dissolve and B will become coated with tin.
(3) B will dissolve and A will become coated with copper.
(4) A will dissolve and B will become coated with copper.

28. Which statement best describes the reaction represented by the equation below?

$2NaCl + 2H_2O + electricity \rightarrow Cl_2 + H_2 + 2NaOH$
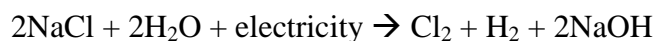
(1) The reaction occurs in a voltaic cell and releases energy.
(2) The reaction occurs in a voltaic cell and absorbs energy.
(3) The reaction occurs in an electrolytic cell and releases energy.
(4) The reaction occurs in an electrolytic cell and absorbs energy.

29. What is the oxidation state of nitrogen in the compound NH_4Br?

(1) –1   (3) –3
(2) +2   (4) +4

30. Given the unbalanced ionic equation:

$$3Mg + \underline{\quad} Fe^{3+} \rightarrow 3Mg^{2+} + \underline{\quad} Fe$$
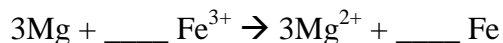
When this equation is balanced, both Fe^{3+} and Fe have a coefficient of

(1) 1, because a total of 6 electrons is transferred
(2) 2, because a total of 6 electrons is transferred
(3) 1, because a total of 3 electrons is transferred
(4) 2, because a total of 3 electrons is transferred

31. A student collects the materials and equipment below to construct a voltaic cell.

• two 250-mL beakers
• wire and a switch
• one strip of magnesium
• one strip of copper
• 125 mL of 0.20 M $Mg(NO_3)_2(aq)$
• 125 mL of 0.20 M $Cu(NO_3)_2(aq)$

Which additional item is required for the construction of the voltaic cell?

(1) an anode   (3) a cathode
(2) a battery   (4) a salt bridge

32. The diagram below represents an operating electrochemical cell and the balanced ionic equation for the reaction occurring in the cell.


$$Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$$
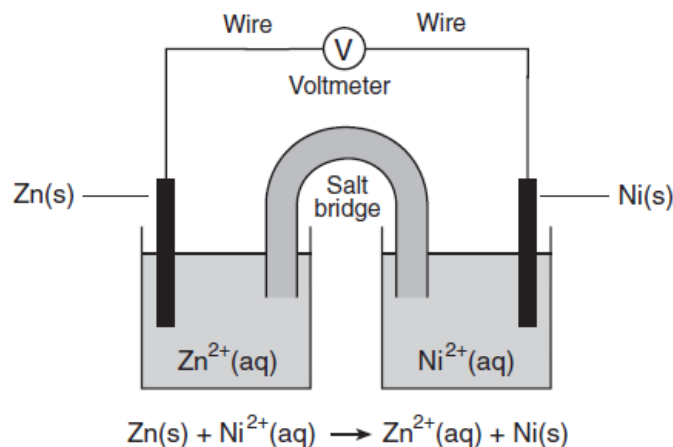
Which statement identifies the part of the cell that conducts electrons and describes the direction of electron flow as the cell operates?

(1) Electrons flow through the salt bridge from the Ni(s) to the Zn(s).
(2) Electrons flow through the salt bridge from the Zn(s) to the Ni(s).
(3) Electrons flow through the wire from the Ni(s) to the Zn(s).
(4) Electrons flow through the wire from the Zn(s) to the Ni(s).

Base your answers to questions 33 through 35 on the information below.

In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.

$$\underline{\quad} Mg(s) + \underline{\quad} HCl(aq) \rightarrow \underline{\quad} H_2(g) + \underline{\quad} MgCl_2(aq)$$
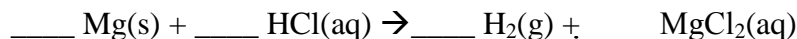
33. State, in terms of the relative activity of elements, why this reaction is spontaneous. [1]

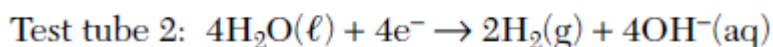34. Balance the equation above, using the smallest whole-number coefficients. [1]

35. Write a balanced half-reaction equation for the oxidation that occurs. [1]

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

The diagram below shows a system in which water is being decomposed into oxygen gas and hydrogen gas. Litmus is used as an indicator in the water. The litmus turns red in test tube 1 and blue in test tube 2.
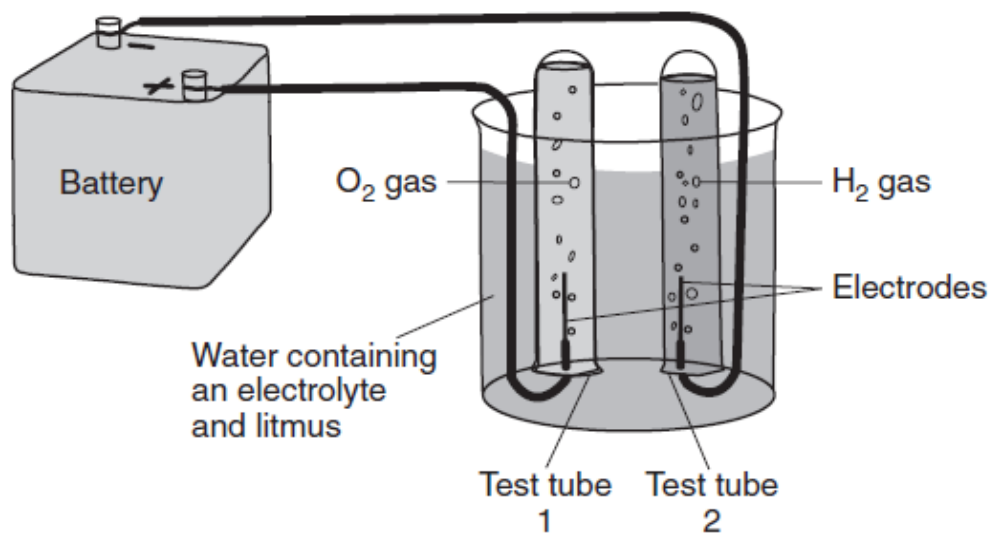
The oxidation and reduction occurring in the test tubes are represented by the balanced equations below.

$$\text{Test tube 1: } 2H_2O(\ell) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

$$\text{Test tube 2: } 4H_2O(\ell) + 4e^- \rightarrow 2H_2(g) + 4OH^-(aq)$$

36. Identify the information in the diagram that indicates this system is an electrolytic cell. [1]

37. Determine the change in oxidation number of oxygen during the reaction in test tube 1. [1]

Base your answers to questions 38 through 40 on the diagram below.

The diagram shows a voltaic cell with copper and aluminum electrodes immediately after the external circuit is completed.
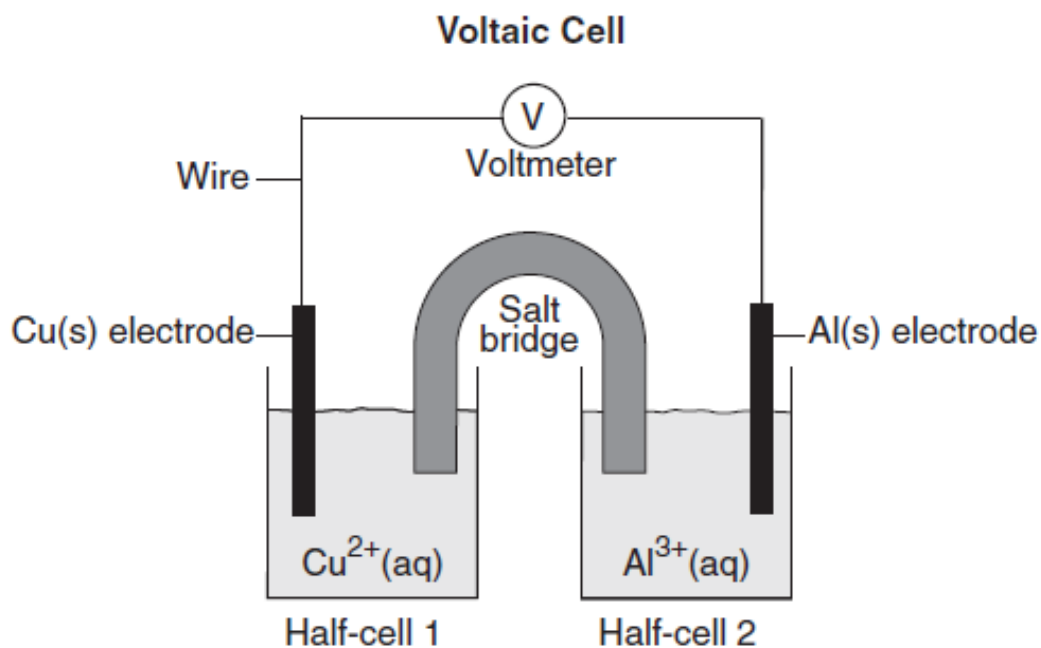
Voltaic Cell

38. Balance the redox equation below using the smallest whole-number coefficients. [1]

_____ $Cu^{2+}(aq)$ + _____ $Al(s)$ → _____ $Cu(s)$ + _____ $Al^{3+}(aq)$
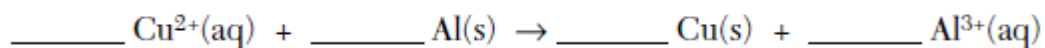
39. As this voltaic cell operates, the mass of the Al(s) electrode decreases. Explain, in terms of particles, why this decrease in mass occurs. [1]

40. Explain the function of the salt bridge. [1]

Base your answers to questions 41 through 44 on the information below.

In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.

Cell with iron and copper electrodes: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$
Cell with zinc and iron electrodes: $Fe^{2+}(aq) + Zn(s) \rightarrow Fe(s) + Zn^{2+}(aq)$
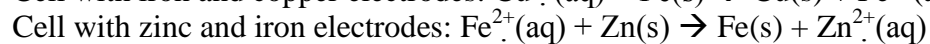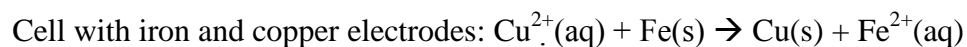
41. State evidence from the balanced equation for the cell with iron and copper electrodes that indicates the reaction in the cell is an oxidation-reduction reaction. [1]

42. Identify the particles transferred between Fe^{2+} and Zn during the reaction in the cell with zinc and iron electrodes. [1]

43. Write a balanced half-reaction equation for the reduction that takes place in the cell with zinc and iron electrodes. [1]

44. State the relative activity of the three metals used in these two voltaic cells. [1]

Answer Key

p. 5-7 Matter

1.	3	8.	3
2.	2	9.	2
3.	3	10.	3
4.	1	11.	1
5.	4	12.	4
6.	3	13.	3
7.	1	14.	4

15. density of neon gas = 0.827 grams/Liter
16. one physical property is the melting of the wax
17. one indication of chemical change is the statement "the candle burns"… burning (combustion) is always a chemical change
18. for the solid diagram, show all the particles touching and with a regular pattern to their arrangement, for the liquid, still show all the particles touching, but do not show them having a regular arrangement
19. Sample 3
20. Two molecules can be made, leaving one particle of "y" and 4 particles of "z" left over.
21. This symbol does not represent a compound because only one type of element is shown.

p. 10-12 Atomic Structure

1.	4	12.	3
2.	2	13.	2
3.	2	14.	2
4.	1	15.	3
5.	3	16.	3
6.	1	17.	4
7.	1	18.	3
8.	1	19.	3
9.	4	20.	1
10.	2	21.	2
11.	4	22.	4

23. Gas A & D. This is determined by noticing that the spectral lines in the mixture includes all the spectral lines of gas A and also those of gas D.
24. An emission spectrum is produced when electrons in the excited state emit energy as they move from higher to lower energy levels.

p. 13-16 Atomic Concepts

1.	3	11.	2
2.	1	12.	2
3.	1	13.	1
4.	2	14.	2
5.	4	15.	1
6.	2	16.	2
7.	4	17.	3
8.	3	18.	2
9.	3	19.	1
10.	4		

20. Si-29 contains 15 neutrons
21. Atomic Mass calculation set up:

$$\begin{aligned}
&(27.98 \text{ amu}) \times 0.9222 \\
&(28.98 \text{ amu}) \times 0.0469 \\
+\ &\underline{(29.97 \text{ amu}) \times 0.0309}
\end{aligned}$$

22. Percent error = (3.29 -3.09)/3.09 X 100% = 6.47%
23. Silicon in the ground state is 2-8-4, so in the excited state, there still need to be 14 electrons shown, with one electron shown at a higher than normal level. Examples include: 2-7-5, 2-8-3-1
24. One piece of info shown by the electron – shell diagram that is not shown on a Lewis dot diagram is that there is info shown about the non-valence electrons. Another piece of info shown is that about the number of protons and neutrons in the nucleus.
25. The magnesium atom has a mass number of 23.
26. For showing bonding, Lewis diagrams are generally more suitable because bonding only involves valence electrons, and this is what is shown on Lewis diagrams.

p. 19-20 Periodic Table

1. 1
2. 4
3. 2
4. 4
5. 2
6. 1
7. 2
8. 1
9. 3
10. 2
11. 2
12. 2
13. 3
14. 4
15. 3

p. 21-24 Periodic Table Cut from Jan 2007-2008 Exams

1. 3
2. 3
3. 4
4. 3
5. 3
6. 1
7. 1
8. 4
9. 4
10. 1
11. 2
12. 1
13. 1
14. 1
15. 3
16. 4
17. 2
18. 4
19. 1

20. The Lewis dot diagram for Uut should show 3 dots since the element is placed in Group 13.
21. The charge on a nucleus is due to the number of protons in it, which is identified by the atomic number. Since Uut is element 113, the answer is +113 (positive charge because protons are positive in charge).
22. Any other Group 13 element would have similar chemical properties as Uut, especially the other metals in that Group, elements In and Tl.
23. Element 118 would be placed in Group 18.
24. The element Uuo has 118 protons, and the isotope Uuo-291 has 291 protons and neutrons combined. So the number of neutrons in Uuo-291 is (291-118) = 173.
25. The number of electrons in an atom of Uuo would be equal to the number of protons, so 118 electrons.
26. Reproducing scientific results is important so that those results can be verified, or confirmed, by other scientists. This has to happen before the discovery is considered factual.

p. 26-27 Equations & Stoichiometry

1. 4	8. 3
2. 4	9. 2
3. 4	10. 3
4. 2	11. 1
5. 2	12. 3
6. 2	13. 3
7. 2	14. 4

p. 28-32 Formulas, Equations & Stoichiometry Review

1. 2	10. 1
2. 3	11. 4
3. 3	12. 3
4. 1	13. 2
5. 1	14. 4
6. 4	15. 2
7. 2	16. 4
8. 4	17. 3
9. 3	18. 2

19.   Percent Error = [(21.4-20.9)/20.9] x 100% = **2.4%**

20.   The empirical formula of C_8H_{18} is **C_4H_9**

21. Na: 23 x 1 =   23
      C: 12 x 1 =   12
      O: 16 x 3 = <u>48</u>
                      106 grams/mole

      %C = 12/106 x 100% = **11.3%**

22. The reaction is **endothermic** because "heat" is written on the reactants side.

23. To do this question, use the mole ratio between the two substances:

$NaHCO_3$: CO_2
2: 1
7: x   2x = 7   **x = 3.5 moles of CO_2**

24. **3** S + **2** $KClO_3$ → **3** SO_2 + **2** KCl + energy

25. The reaction is a **synthesis** reaction.

26. Fe: 55.8 x 2 = 111.6
     O: 16 x 3 =   <u>48.0</u>
                      159.6 grams/mole

27. The IUPAC (systematic) name for this compound is **Iron III oxide**. The roman numeral 3 is needed because Iron ions can be charged +2 or +3, and is chosen in this case because the formula includes iron with the +3 charge.

28. moles = 19 / 95 = **0.20 moles**

29. The type of ORGANIC reaction is **substitution.**

30. One advantage of using ozone is that it is **safer to use.** Another is that it is more **environmentally friendly.**

31. **Three** significant figures are shown.

32. **%** water mass = (0.76 g of water/ 2.13 g of hydrate) x 100% = 35.7%

33. The crucible containing the sample must be heated until a constant mass is achieved in order **to insure that all the water has been driven out of the hydrate**

p. 35-36 Bonding

1. 3		7. 1	
2. 2		8. 1	
3. 1		9. 2	
4. 1		10. 2	
5. 2		11. 4	
6. 3		12. 4	

p. 37-41 Bonding Review Cut from Jan 2007-Jan 2008 Exams

1. 2		10. 2	
2. 4		11. 4	
3. 2		12. 3	
4. 2		13. 1	
5. 3		14. 3	
6. 3		15. 1	
7. 4		16. 2	
8. 2		17. 4	
9. 3		18. 1	

19. The P-Cl bond is more polar than the P-S bond because the electronegativity difference between P and Cl is 1.0, whereas it is only 0.6 between P and S. The greater the electronegativity difference, the more polar the bond.
20. The boiling point of octane according to the graph is about **125°C**.
21. The relationship is as the **molar mass increases**, the **strength of the intermolecular forces between molecules also increases.**
22. **AQ$_2$**
23. The bonds between nonmetals is always **covalent**.
24. The total mass before and after a reaction is always the **same** due to conservation of mass.
25. Methane is a nonpolar molecule because the **electrons are spread out evenly** in the molecule.
26. The boiling point of water under 90 kPa of atmospheric pressure is **97°C**.
27. The graph of temperature vs. surface tension is linear, and sloping negatively
28. At 60°C, one can read the graph, and or interpolate from the data, to estimate the surface tension value to be about **65 mN/m**
29. As temperature increases, the surface tension decreases.
30. The lower surface tension value for CCl$_4$ indicates that the intermolecular forces between CCl$_4$ molecules are weaker than those between H$_2$O molecules.
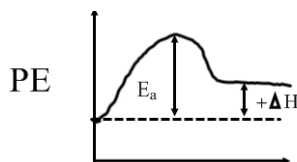
p. 45-49 Heat and Temperature Cut from Jan 2007-Jan 2008 Exams

1. 2
2. 1
3. 3
4. 4
5. 2
6. The diagram should look like

PE
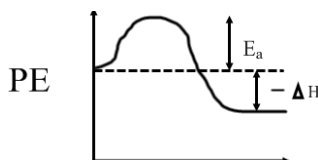
7. According to the chemical equation, reacting 1 mole of propane (C$_3$H$_8$) releases 2219.2 kJ , so 2.5 moles x 2219.2 kJ = **5548.0 kJ**

8. The diagram should look like

PE   (Table T)

9. Use the q = mcΔT equation

$q = (5.00 \text{ g}) (4.71 \text{ J/g }^{\circ}\text{C}) (30 \,^{\circ}\text{C})$

q = 707 J

9. average **KE = remains the same** since Temperature is not changing along BC.
   PE = increasing as the sample vaporizes

10. Use the q = mH_v equation from Table T

$q = 5.00 \text{ g} (1370 \text{ J/g})$

q = 6850 J

11. The entire sample of NaCl is in the liquid phase along segment CD, where the temperature changes from 801 to 1465^{0}C, so the range over which the entire sample is in the liquid phase is (1465-801) = **664°C**

12. The average KE is changing anywhere where the temperature is changing, so correct answers are **segments AB or CD**

13. The segment where NaCl is in one phase and able to conduct electricity would be where it is all and only in the liquid phase (it then has mobile ions capable of conducting electricity). So the **segment CD** applies.

14. Use the q = mcΔT equation from Table T

Obtain the specific heat of water value from Table B, and the change in temperature value from the paragraph above the picture (22°C to 57°C)

$q = (100.0 \text{ g}) (4.18 \text{ J/g }^{\circ}\text{C}) (35^{o}\text{C})$
q = 14630 J

15. Percent error = (30.2 – 25.9/30.2) x 100% = **14.2%**
16. The boiling point is **120°C.**
17. During the first minute of heating the sample is in the solid phase, so draw the particles touching each other and arranged in a definite pattern.
18. The substance is in the liquid phase only from 4 minutes to 7, so the answer is **3 minutes**.
19. Heat is being added at a constant rate of 15kJ per minute according to the paragraph above the graph, and the substance takes from 2 minutes to 4 minutes to melt, so the heat needed to melt is 15 x 2 = **30 kJ**

p. 51-54 Rate of Reaction & Equilibrium Cut from Jan 2007-Jan 2008 Ex.

1. 3   7. 4
2. 4   8. 1
3. 2   9. 3
4. 3   10. 4
5. 2   11. 3
6. 3

12. As molecules increase in temperature, they move faster and collide harder. As a result, more of them will break each other's bonds and go on to react.
13. The rates must be **equal**.
14. Adding more $HC_2H_3O_2$ will cause the equilibrium reaction to "shift right" and in so doing, produce more H^+ ions, so the **concentration of H^+ increases**
15. This one is given to you on Table R!
16. One source of the activation energy would be the hard contact between the two ceramic balls.
17. The fact that the "571.6 kJ" value is located on the products side.
18. The system coverts reactants that are all in the gas phase to one product in the liquid phase. SO the entropy decreases because:
    - o **The product particles are in a less energetic phase, OR**
    - o **There are fewer product particles than reactant particles**
19. Adding Cl_2 causes the reaction to "shift to the right" meaning the reaction speeds up in the forward direction, producing more OCl^- ion.
20.
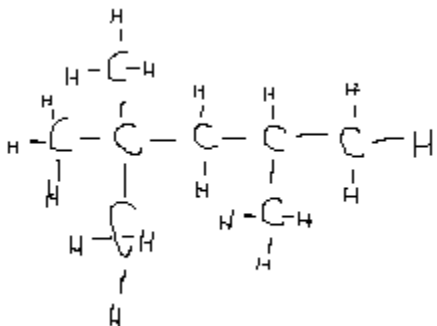
21. In the vapor phase, the molecules of octane are no longer in contact with each other, so they have a less orderly arrangement, whereas in the liquid state, the intermolecular forces between molecules still exists and causes the molecules to be in a more orderly arrangement.

p. 56-60 Water & Solutions Cut from Jan 2007-Jan 2008 Exams
1. 3
2. 2
3. 3
4. 4
5. 1
6. 3
7. 4
8. 2
9. 1
10. 3
11. 4
12. 3
13. 2
14. 3
15. 4
16. 4
17. 4
18. 2
19. The water molecules should be shown with their positive ends near the negatively charged Cl^- ion, or in other words, the hydrogen ends should be closest to the Cl^- ion.
20. This point when plotted falls below the solubility curve, so the solution is **unsaturated**.
21. O_2 has a low solubility due to the fact **that O_2 molecules are nonpolar and water molecules are polar**, hence they have little ability to attract each other, so O_2 does not dissolve well.
22. Use the parts per million equation from Table T:
    ppm = (0.0070 g/1000.0070 g) X 1,000,000 = **7 ppm**
23. Pressure has **no effect** on the solubility of KNO_3 (solid). Pressure over a solution only affects the solubility of gaseous solutes like CO_2.

24. At 15°C, the solubility of KNO₃ is 30 g per 100 g of water. Since 65 grams of KNO₃ are present in the system, (65-30) = **35 g** of KNO₃ will settle to the bottom.
25. The "-ol" suffix on the name "1,2-ethanediol" tips you off that this compound is an **alcohol**.
26. Since water is polar, the fact that 1,2 ethanediol mixes well with water implies that the **molecules of 1,2 ethanediol are also polar**.
27. Use the "moles" equation from Table T;

$$Moles = 6690/62 = \textbf{108 moles}$$

28. Use the parts per million equation from Table T:

300 ppm = (mass of solute/1000 g of solution) x 1,000,000
Mass of KOH = 0.300 grams

29. **97°C**

p. 62-65 Acids & Bases Cut from Jan 2007-Jan 2008 Exams

1. 4
2. 3
3. 2
4. 3
5. 4
6. 2
7. 2
8. 4
9. 2
10. 4
11. 1
12. 1
13. 2
14. 3
15. 2
16. 2
17. 4
18. **blue**
19. **Thymol blue and bromothymol blue**
20. Since methyl orange is only yellow if the pH is greater than 4.4, and litmus is only red if the pH is less than 5.5, any **pH value between 4.4 and 5.5** is an acceptable answer.
21. **Any base** would work, so look on Table L for ideas,if you need to.   Acceptable answers include NaOH, Ca(OH)₂ etc.
22. **Sample B is the methanol.** This can be concluded because alcohols are **molecular substances** and do not dissolve in water to produce ions, and hence are **nonconductors** when dissolved.
23. Reactivity with Mg metal – a reaction indicates an acid
     Bromothymol blue indicator – yellow color indicates that the pH was   6.0 or lower, which is an acidic pH
24. The explanation must be that all 3 solutions have **pH values greater than 4.4**.
25. **OH⁻¹ ion** (hydroxide)
26. Molarity = moles / Liter of solution = 0.500 moles/0.400 L =**1.2 Molar or 1.2 M**
27. no question
28. This is a neutralization equation so the products of a reaction between an acid and a base are always water and some salt. In this case, **H₂O + NaNO₃**
29. Use Table K… "**hydrochloric acid**"
30. Use the titration equation from Table T:
$$M_A V_A = M_B V_B$$
$$M_A (10.22mL) = (0.500 M) (17.03 mL)$$
$$\textbf{M}_A = \textbf{0.833 Molar}$$
31. The volumes each have 4 significant figures, and the Molarity value has 3, so the answer for the calculated HCl Molarity should have **3 significant figures**.

1. 2
2. 2
3. 4
4. 3
5. 2
6. 2
7. 4
8. 2
9. 2
10. 2
11. 1
12. 1
13. 2
14. 2
15. 2
16. 3
17. 1
18. 4
19. 3
20. 1
21. 4
22. 1

23. moles = 19/95 = **0.20 moles**
24. The type of ORGANIC reaction is **substitution**
25. The empirical formula is C_4H_9
26. The – COOH group on the reactant molecule should be circled
27. ethanoic acid
28. You should draw 1 – propanol
29. Methane is a nonpolar molecule because its valence electrons are distributed evenly throughout.
30. 125°C
31. As molar mass increases, so does the boiling point temperature.

p. 72-74 Phases & Gases

1. 3
2. 3
3. 2
4. 2
5. 1
6. 4
7. 4
8. 3
9. 2

10. Neon has an atomic mass of 20.1 so one mole is equal to 20.1 grams.

$$\text{Density} = m/V = 20.1 \text{ g} / 24.4 \text{ L} = \textbf{0.824 g/L}$$

11. The boiling point is **120°C.**
12. The particles should all be shown touching each other, and should be arranged geometrically, or with a specific pattern shown.
13. The sample becomes completely melted at 4 minutes,and then warms up and starts to vaporize at 7 minutes. So it is in the liquid ONLY phase from 4 to 7 minutes, a total of **3 minutes**.
14. The paragraph above the graph states that heat is added at "a constant rate of 15 kJ per minute." The sample takes from 2 minutes to 4 minutes to melt at its melting point (70°C). SO multiply 15 x 2 = **30 kJ**
15. The initial volume of the cylinder is 125 mL, which is equivalent to **0.125 Liters**
$$125 \text{ mL } \times \frac{1.0 \text{ Liter}}{1000 \text{ mL}} = 0.125 \text{ Liters}$$

16. Use the combined gas law:
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
** Since Temp remains the same, you may drop "T_1" and "T_2" out of the equation:

$$(1.0 \text{ atm}) (125 \text{ mL}) = (1.5 \text{ atm}) V_2$$

$$V_2 = \textbf{83 mL}$$

17. This is a way of asking Avogadro's hypothesis. For gases, there is a predictable relationship between the V, T, P and number of molecules for the sample. If two gas samples have the same P, T, and V values, they must **contain the same number of molecules.**

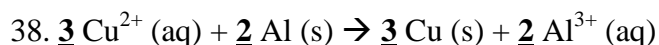p. 76-82 Electrochemistry (Oxidation-Reduction Reactions)

1. 1
2. 1
3. 1
4. 4
5. 4
6. 4
7. 2
8. 4
9. 2
10. 4
11. 4
12. 4
13. 4
14. 4
15. 1
16. 4

17. 2
18. 1
19. 3
20. 4
21. 3
22. 3
23. 1
24. 3
25. 3
26. 4
27. 2
28. 4
29. 3
30. 2
31. 4
32. 4

33. **Mg is more reactive** and correctly oxidizes, while **H⁺ is less reactive** and correctly reduces.

34. **1** Mg (s) + **2** HCl(aq) → **1** H₂ (g) + **1** MgCl₂ (aq)

35. Mg → Mg^{2+} + 2 e⁻
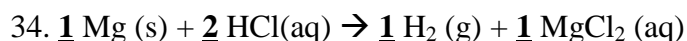
36. The cell is connected to a **battery**

37. From **-2 to 0**.

38. **3** Cu^{2+} (aq) + **2** Al (s) → **3** Cu (s) + **2** Al^{3+} (aq)
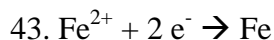
39. The Al(s) is oxidized, and therefore **loses electrons** to form Al^{3+} (aq) in solution.

40. The salt bridge permits the **migration of ions** between the cells (and helps to close the circuit!)

41. **Oxidation numbers are changing** from the reactant to product side (Cu^{2+}→Cu; Fe→Fe^{2+})

42. Electrons

43. Fe^{2+} + 2 e⁻ → Fe

44. Zn is the most reactive, followed by Fe, then Cu being the least reactive.