Grading Sheet: Quarter 1

Record your scores on the sheet below as the assignments instruct you to do so.

Date	Lesson	Assignment	My Score	Possible Score
	2	Assignment		5
	3	Assignment		10
	4	Self-Assessment (pot. extra credit)		28
	5	Daily Assignment		5
	6	Lab		40
	7	Quiz Me		6
	8	Quiz Me		4
	10	Lab		20
	11	Quiz Me		6
	12	Unit 1 Test (potential extra credit)		45
	16	Daily Assignment		17
	19	Lab		20
	20	Quiz		12
	21	Lab Report		20
	22	Lab		15
	24	Test		32
	27	Assignment (potential extra credit)		7
	29	Self-Assessment		8
	30	Quiz Me		7
	35	Self-Assessment		14
	35	Quiz Me		7
	36	Lab Report		20
	38	Test		74
	40	Quiz Me		9
	43	Quiz Me		13
	44	Chart		20
	45	Quiz Me		22
		TOTAL		

Grading Sheet: Quarter 2

Record your scores on the sheet below as the assignments instruct you to do so.

Date	Lesson	Assignment	My Score	Possible Score
	47	Lab		50
	49	Test (potential extra credit)		90
	52	Lab		50
	52	Assignment		20
	52	Assignment		10
	54	Assignment		20
	56	Test		44
	61	Assignment		10
	63	Assignment		20
	64	Assignment		10
	68	Test		100
	70	Quiz		5
	72	Lab		50
	74	Assessment		16
	76	Test		21
	76	Test		46
	78	Self Assessment		10
	79	Assignment		10
	83	Assignment		50
	87	Lab		50
	90	Lab		50
		TOTAL		

Grading Sheet: Quarter 3

Record your scores on the sheet below as the assignments instruct you to do so.

Date	Lesson	Assignment	My Score	Possible Score
	92	Test		46
	94	Assignment		20
	95	Assignment		15
	97	Self-assessment		20
	100	Lab		50
	101	Assignment		10
	107	Test		100
	111	Assignment		10
	116	Test		80
	119	Midterm		20
	124	Lab		50
	126	Assignment		20
	127	Assignment		20
	128	Assignment		5
	130	Test		100
	132	Assignment		10
		TOTAL		

Grading Sheet: Quarter 4

Record your scores on the sheet below as the assignments instruct you to do so.

Date	Lesson	Assignment	My Score	Possible Score
	139	Lab		64
	140	Lab		25
	141	Assignment		10
	143	Lab		31
	145	Assignment		100
	149	Assignment		20
	151	Assignment		10
	153	Lab		50
	157	Test (potential extra credit)		65
	160	Assignment		10
	163	Assignment		10
	165	Lab		50
	167	Test		100
	170	Assignment		10
	174	Lab		20
	174	Assignment		25
	176	Test (potential extra credit)		46
	179	Final		120
	180	Essay		20
		TOTAL		

Scientific Method:

1. **Scientific method** – a method of research in which a problem is identified, relevant data are gathered, a hypothesis is formulated from these data, and the hypothesis is empirically tested.
2. **Observation** – the act of noting and recording something with instruments.
3. **Hypothesis** – a tentative explanation for an observation or scientific problem that can be tested by further investigation.
4. **Experiment** – a test under controlled conditions that is made to demonstrate a known truth, examine the validity of a hypothesis, or determine the efficacy of something previously untried.
5. **Independent variable (manipulated variable)** – variable whose value is being altered to bring a change in some condition.
6. **Dependent variable (responding variable)** – the observed variable in an experiment or study whose changes are determined by the presence or degree of one or more independent variables.
7. **Controlled variable** – a sample in which a factor whose effect is being estimated is absent or is held constant, in order to provide a comparison.
8. **Conclusion** – a position reached after consideration of data obtained from an experiment.

Measurements and Calculations:

9. **Qualitative data** – data described in terms of some quality or categorization that may be informal or may use ill-defined characteristics such as warmth and flavor.
10. **Quantitative data** – data described in terms of quantity and in which numerical values are used.
11. **Mass** – a property of matter equal to the measure of an object’s resistance to changes in either its speed or direction of its motion. The mass of an object is not dependent on gravity and therefore is different from but proportional to its weight.
12. **Volume** – the amount of space occupied by a three-dimensional object or region of space.

13. **Length** – the measurement of the extent of something along its greatest dimension.
14. **Temperature** – a measure of the average kinetic energy of the particles in a sample of matter, expressed in terms of units or degrees designated on a standard scale.
15. **Metric system** – a decimal system of units based on the meter as a unit length, the kilogram as a unit mass, and the second as a unit time.
16. **Scientific notation** – a method of writing or displaying numbers in terms of a decimal number between 1 and 10 multiplied by a power of 10.
17. **Dimensional analysis** – a technique that involves the study of dimensions of physical quantities.
18. **Significant figures** – all of the numbers in a measurement that are known to be accurate plus one that is uncertain.

Lesson 4: Scientific Method

A ___________________ ___________________ of ______________________

___________________.

Steps of the Scientific Method:

1. Making ________________ that lead to a __________________.
2. Forming a __________________ to answer the __________________.
3. Testing the __________________ by _________________.
4. Making a ______________ based on the results of the _________________.

Definitions:

Observing: use of the ________________ and ________________ in the lab to

____________________ ___________________.

Conclusion: a statement based on _______________ and prior ____________.

Observations vs. Conclusions:

1. _____________ 2. ____________ 3. ____________ 4. ____________

5. _____________ 6. ____________ 7. ____________ 8. ____________

Observations:

Quantitative	**Qualitative**
Includes _____________	Does ____________ include
and ____________.	________________ and units.

Qualitative vs. Quantitative:

1. ______________ 2. _____________ 3. _____________ 4. _____________

5. ______________ 6. _____________ 7. _____________ 8. _____________

Hypothesis: an _________________ ____________________.

A good hypothesis:

1. ___________________ an __________________

2. can be __________________

3. will _____________________ an __________________

Practice forming a hypothesis:

1. What is the ______________ of the _______________ _______________?

2. As the candle _________________, it gets ___________________. Where does the _____________________ go?

Experiment:

1. is _________________ to ___________________ a ___________________

2. involves _________________

3. is performed under _________________ ____________________

Variables: factors that can be ____________________.

Control: a ________________ that is held _____________________.

Lesson 4: Scientific Method (cont.)

In a good experiment:

1. only ___________________ ___________________ are allowed to _____________.
2. the ____________________ (or _________________) ___________________ is changed by the ___________________.
3. the ____________________ (or _________________) ___________________ changes as a result of the ________________.
4. all other ________________ are _________________ because they are held __________________.

The Chemistry Quiz:

1. ________ 2. _______ 3. _______ 4. _______ 5. _______

$M \times 10^n$
$1 \leq M < 10$

Are the following in correct form? Write C for correct or I for incorrect.

9.2×10^2 __________

0.4×10^{-14} __________

1.8×2^5 __________

860×10^{-3} __________

9.898989887×10^1 __________

Problem Set One – Change from scientific notation to decimal notation:

a.

b.

c.

d.

e.

Problem Set Two – Change from decimal notation to scientific notation:

a.

b.

c.

d.

e.

The Chemistry Quiz:

CR1. ______ CR2. _______ 1. _______ 2. _______ 3. ________ 4. _________

International System (________)

____________________ system used in __________________

AKA _______________ ___________________

<u>Length</u> - ________________ between ________________ ______________

Base unit - _______________ (________)

Instrument - ___________________

<u>Mass</u> - ___________________ of ________________ in a ________________

Base unit - _______________ (________)

Instrument - ___________________

<u>Volume</u> - ________________ of ____________ an object ________________

Base unit - _______________ (________)

Instrument - ___________________ ruler or __________________

____

<u>Temperature</u> – average ___________________ ___________________

Base unit - _______________ (___) or ________________ (___)

Instrument - ___________________

Measuring example:

Width of a textbook = ________________

A measurement includes both a _____________ and a ______________!

Significant Figures - ________________ numbers plus one ______________ number

Examples of volume measurements:

Volume of liquid = ___________________

Volume of solid

Rectangular solid – V= ________________

Volume of cube = _______________

Volume of rectangular solid = _________________________

Irregular solid - _________________ ___________________

Practice Problems:

Volume of the rock

Volume of rectangular solid

Volume of rectangular solid using water displacement

__________cm^3 = _________mL

The Chemistry Quiz:

CR1. ______ CR2. _______ 1. _______ 2. _______ 3. _______ 4. ________

Metric Prefixes

One ________________thing equals ____________ ______________ things

Prefix	Abbreviation	One smaller =	Many larger
centi-		**1 g =**	
kilo-		**1 kg =**	
Mega-		**1 Mg=**	
deci-		**1 g =**	
milli-		**1 g =**	
micro-		**1 g =**	

Prefix Questions

1. __________m = 1 km
2. __________km = 1 m
3. Which is larger? milliliter or centiliter
4. Which is larger? kilogram or megagram
5. Which is larger? millimeter or kilometer
6. a penny = 1 ______________ dollar
7. a dime = 1 __________ dollar
8. a megabuck = ____________ dollars
9. a kilopenny = _____________ dollars

Prefix and Base Unit Questions

1. ________________________
2. ________________________
3. ________________________
4. ________________________
5. ________________________

Conversions

Convert 152 cm to m

Conversion factor: a _____________ of fact expressed as a ___________ equaling _____________

Fact: 1 m = 100 cm

$\dfrac{1\text{ m}}{100\text{ cm}}$ $\dfrac{100\text{ cm}}{1\text{ m}}$

Convert ________ kg to grams

Convert ________ μs to ds

_____ mm = ____ km

Problem Set One

1.

2.

3.

4.

5.

How many yards are in __________ km?

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. _______ 5. _______

Chemical and Physical Properties and Changes:

1. **Physical property** – property of a compound that can change without involving a change in chemical composition; examples are the melting point and boiling point.

2. **Chemical property** – any of a chemical's properties that become evident during a chemical reaction; examples are reactivity and flammability.

3. **Physical change** – any change not involving a change in the substance's chemical identity. Includes a change from one state (solid or liquid or gas) to another without a change in chemical composition.

4. **Chemical Change** – any process determined by the atomic and molecular composition and structure of the substances involved.

Classification of Matter:

5. **Element** – a substance composed of atoms having an identical number of protons in each nucleus. Elements cannot be reduced to simpler substances by normal chemical means.

6. **Pure substance** – a sample of matter, either an element or a compound, that consists of only one component with definite physical and chemical properties and a definite composition.

7. **Compound** – a pure, macroscopically homogeneous substance consisting of atoms or ions of two or more different elements in definite proportions that cannot be separated by physical means. A compound usually has properties unlike those of its constituent elements.

8. **Mixture** – a composition of two or more substances that are not chemically combined with each other and are capable of being separated.

9. **Solution** – a homogeneous mixture of two or more substances, which may be solids, liquids, gases, or a combination of these.

10. **Heterogeneous** – consisting of dissimilar parts. Heterogeneous mixtures have distinguishable phases.

11. **Homogeneous** – uniform in structure or composition throughout. Homogeneous mixtures have atoms and molecules interspersed.

12. **Alloy** – a homogeneous mixture or solid solution of two or more metals, the atoms of one replacing or occupying interstitial positions between the atoms of the other: Brass is an alloy of zinc and copper.

13. **Distillation** – the evaporation and subsequent collection of a liquid by condensation as a means of purification.

14. **Density** – the mass per unit volume of a substance. Commonly measured in grams per milliliter (g/mL) or grams per cubic centimeter (g/cm^3).

Definitions of Properties

Physical properties can be _______________ without _________________ _______________ matter.

Chemical properties describe how a _________________ ______________ with other ___________.

Examples of Properties

Physical Properties **Chemical Properties**

Phases of Matter

	Shape	Volume	Particles
Solids			
Liquids			
Gasses			

Types of Changes

Physical change - ________ ______________ in ____________ of substance

Chemical change - _________ or more _____________ substances produced

Phase changes are ______________ ______________

Freezing point - _________ to ____________

Melting point - __________ to ____________

_______________ point = _____________ point

Boiling point - ___________ to ____________

Condensation point - _______________ to ____________

Sublimation point - ___________ to ____________

Examples of Changes

Physical Changes **Chemical Changes**

Law of Conservation of Mass: Mass ____________ be _________________

_______________ ______________.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Types of Matter

Element – _______________ substance

- _____________ kind of _____________

- represented by __________ or ____________ letter

  symbol

- example: _______________ (____)

Compound – _________ or more ____________ _____________ combined

- represented by __________ ______________

- example: _______________ (____)

- H_2O ________________ into _______________ ______________ and

  ______________ ______________

- _____________ in same ____________

Mixture – ____________ or more substances ___________ ______________

- do not _____________ in same ____________

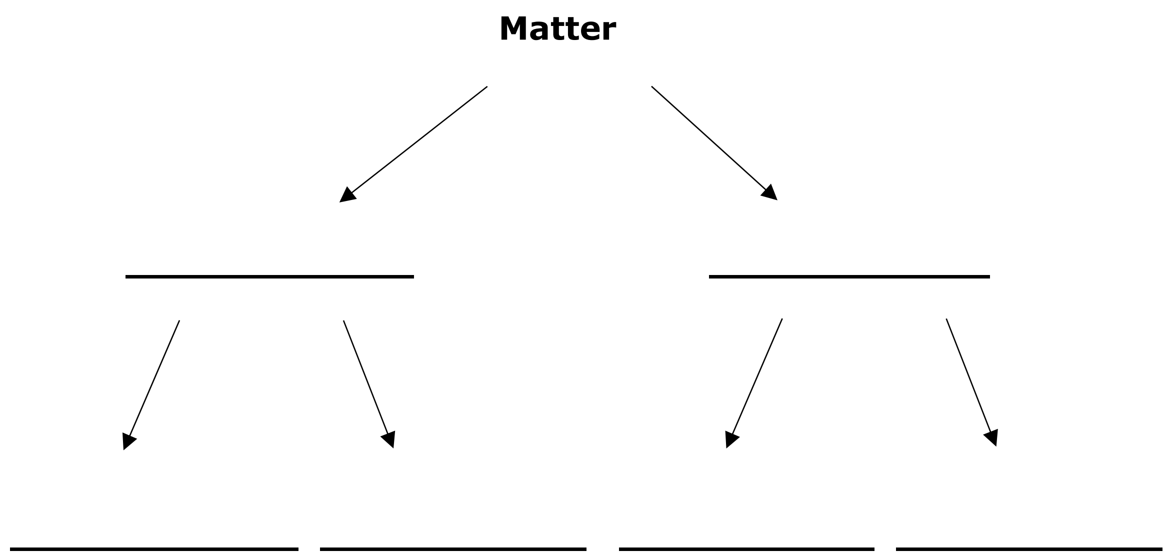
- heterogeneous mixture – not ______________

- homogeneous mixture (solution) – ___________________

  - alloy – _____________ of two or more _________________

    - Pure ____________________ are _________________

Lesson 15: Classification of Matter (cont.)



E – element S – homogeneous mixture/solution
C – compound HE – heterogeneous mixture

hydrogen peroxide _____

carbon dioxide _____

pizza _____

salad dressing _____

rust _____

apple juice _____

carbon ____

carbon monoxide _____

steel _____

calcium _____

lead ______

The Chemistry Quiz

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 17: Separation of Mixtures

The components of a mixture may be separated based on the physical

properties of ___________________

_______________ (An instrument that uses this property to separate mixtures when tiny solid particles are dissolved in a liquid is called a

____________________.)

Some techniques used to separate the components of a mixture are:

_______________ – used to separate liquids and solids

$\left.\begin{array}{l}\text{\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_}\\ \text{\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_}\end{array}\right\}$ used to separate solutions

	Observation	**Explanation**
1.	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________
2.	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________
3.	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________
	_________________________	______________________________

Lesson 17: Separation of Mixtures (cont.)

Density - ___________ per unit of __________

Density =

(D= ______)

common units for density: _______, _________

Example problem: Suppose we have an object with a mass of ________ grams and a volume of ______ mL. What would be the density of this object?

Problem Set One

1.

2.

3.

Problem Set Two

1.

2.

3.

The Chemistry Quiz

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

DATA TABLE:

	Metal A	Metal B	Metal C
Mass of Empty Graduated Cylinder (g)			
Mass of Graduated Cylinder + Metal (g)			
Mass of Metal (g)			
Volume of Metal Sample (mL)			
Density (g/mL)			

Calculations:
Mass of Metal = Mass of Graduated Cylinder + Metal – Mass of Empty Graduated Cylinder

Metal A:

Metal B:

Metal C:

Density = Mass of Metal / Volume of Metal Sample

Metal A:

Metal B:

Metal C:

Questions:

Using the values given in the Problem Description, and comparing those values to the density you calculated in your data table, what is the identity of each metal?

1. Metal A:
2. Metal B:
3. Metal C:

4. If the metals were not in powdered form (perhaps, pellets), what would be an alternative way of finding volume?

Lesson 25: Atomic Theory Key Terms

1. **Democritus** – (born c. 460 — died c. 370 BC) Greek philosopher. Though only a few fragments of his work survive, he was apparently the first to describe invisible "atoms" as the basis of all matter. His atoms — indestructible, indivisible, incompressible, uniform, and differing only in size, shape, and motion — anticipated with surprising accuracy those discovered by 20th-century scientists.
2. **Atom** – A unit of matter, the smallest unit of an element having all the characteristics of that element and consisting of a dense, central, positively charged nucleus surrounded by a system of electrons.
3. ***Atomos*** – Greek word meaning indivisible or not able to be cut.
4. **Aristotle** – Greek philosopher who did not follow the teachings of Democritus. He believed in the four earthly elements (fire, earth, air, and water), and proposed a fifth heavenly element (aether).
5. **Dalton** – (born 1766 - died 1844) English scientist who was one of the first to decide that all matter is made up of small particles, or atoms.
6. **Law of Conservation of Mass** – The notion that mass, or matter, can be neither created nor destroyed. According to conservation of mass, reactions and interactions which change the properties of substances leave unchanged their total mass; for instance, when charcoal burns, the mass of all of the products of combustion, such as ashes, soot, and gases, equals the original mass of charcoal and the oxygen with which it reacted.
7. **Law of Definite Proportions** – elements composing a compound are always present in the same proportions by mass.
8. **Law of Multiple Proportions** – law stating that when two elements can combine to form more than one compound the amounts of one of them that combines with a fixed amount of the other will exhibit a simple multiple relation.
9. **Rutherford** – (born 1871 - died 1937) New Zealand physicist who named the nucleus and proton and won the 1908 Nobel prize in chemistry.
10. **Nucleon** – A proton or a neutron, especially as part of an atomic nucleus.
11. **Proton** – A stable, positively charged subatomic particle having a mass 1,836 times that of the electron.
12. **Neutron** – An electrically neutral subatomic particle having a mass 1,839 times that of the electron.

Lesson 25: Atomic Theory Key Terms (cont.)

13. **Electron** – A stable subatomic particle having a rest mass of 9.1066×10^{-28} grams and a unit negative electric charge of approximately 1.602×10^{-19} coulombs.
14. **Isotope** – One of two or more atoms having the same atomic number, but different mass numbers, due to a different number of neutrons in the nucleus.
15. **Mass number** – The sum of the number of neutrons and protons in an atomic nucleus.
16. **Photon** – The quantum of electromagnetic energy, regarded as a discrete particle having zero mass, no electric charge, and an indefinitely long lifetime.
17. **Spectroscopy** – Study of spectra, especially experimental observation of optical spectra.
18. **Quantum number** – A set of real numbers assigned to a physical system that individually characterize the properties and collectively specify the state of a particle or of the system.
19. **Orbital** – The wave function of an electron in an atom or molecule, indicating the electron's probable location.
20. **Sublevel** – One or more orbitals in the electron shell of an atom.
21. **Hund’s Rule** – If two or more orbitals of equal energy are available, electrons will occupy them singly before filling them in pairs.
22. **Pauli Exclusion Principle** – No two electrons may simultaneously occupy the same quantum state.
23. **Aufbau Principle** – Used to determine the electron configuration of an atom, molecule, or ion. The principle postulates a hypothetical process in which an atom is "built up" by progressively adding electrons. As they are added, they assume their most stable conditions (electron orbitals) with respect to the nucleus and those electrons already there. According to the principle, electrons fill orbitals starting at the lowest available (possible) energy states before filling higher states (e.g. 1s before 2s).

Lesson 25: Atomic Theory Notes

Model: A __________ idea used to explain ___________ facts in ___________.

Theory: An_______________ of _____________ facts and ____________.

To remain valid, models and theories must:

- _____________ all known ___________
- enable ____________ to make correct ___________

Democritus:

- proposed the ___________ of an _______________
- word comes from the _________ word __________ which means

  _____________ _______________ _______________ or "indivisible"

Aristotle:

- _____________ the ____________ of the _______________
- said _______________ could be _____________ _______________

Dalton's theory proposed that atoms:

- are _______________ _______________ of _____________
- are _______________
- of the _____________ _______________ are ____________
- of ___________ _______________ are ____________
- _____________ in _____________, ______________ _______________

  ratios to form _______________

J.J. Thomson

- credited with ______ of _____________: a blow to ___________ _____________ atom
- proposed the ____________ _______________ model of the __________: _____________ charged _____________ embedded in a ball of _____________ ______________

Rutherford's Gold-foil Experiment:

- aimed _____________ _______________ at ___________ ____________
- _____________ passed ______________
- a _________________ particles were _____________
- _____________ particles even _____________ _______________

Conclusions of Rutherford's experiment:

- _____________ of the _______________ is ___________ ____________
- _____________ _______________ charged _____________

The Chemistry Quiz

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Periodic Table of the Elements

Group→ ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Lesson 29: Structure of the Atom Notes

Nucleons: ______________ in the ____________ of _________________.

Atomic Number: ____________________ of _____________ in the _________ of an __________.

Neutral atom: # of _____________ (+) = # of ________________ (-)

Isotope: ____________ of an ________________ that have ___________ ________________ of _____________.

Isotopes of Hydrogen:

Hydrogen-1

- _____________ proton and __________ neutrons

Hydrogen-2

- _____________ proton and __________ neutrons

Hydrogen-3

- _____________ proton and __________ neutrons

Mass Number: _________________ number of ___________ and ___________ in an ________________.

Example: Carbon-14 _______________ Neon-20 _________________

Particle	Charge	Mass	Location
Proton			
Neutron			
Electron			

Lesson 29: Structure of the Atom Notes (cont.) Chemistry with Lab

Atomic Mass: __________ of the __________ of all the element's _________.

Subatomic Particles

of protons = ____________ ______________

# of electrons = ___________ ______________

# of neutrons = __________ __________ - ___________ ____________

iron: | 26 / *Fe* / 55.8 |

__________ protons

__________ electrons

__________ neutrons

oxygen-17: | 8 / *O* / 15.994 |

__________ protons

__________ electrons

__________ neutrons

helium: | 4 / *He* / 2 |

__________ protons

__________ electrons

__________ neutrons

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 29: Atomic Structure

1. The atomic number tells the number of positively charged ____________ in the nucleus of an atom. The atom is ______________ because this is also the number of _______________ charged ____________ in the atom.

2. The mass number tells the total number of ___________ and ___________ in the nucleus of an atom. These particles collectively are called ____________ since both are located in the nucleus.

3. Isotopes are atoms of the same element with different numbers of ____________ which results in different _____________ numbers.

4. Write the hyphen notation of the three isotopes in hydrogen:
   a.
   b.
   c.

5. Fill in the following chart with the missing numbers or symbols. (Hint: Isotopes may have different masses than the periodic table lists!):

Symbol	**Atomic**	**Mass**	**Protons**	**Electrons**	**Neutrons**
Hg	80	201			
Mo	42	96			
	29			29	35
C		14			
	15	31			
Pb		208	82		
Na	11				12

Lesson 30: Isotope Problems

1. A student looked up the naturally occurring isotopes of bromine and found the following information:

   50.54% of the naturally occurring isotopes of bromine have an atomic mass of 78.92 u while 49.46% of the naturally occurring isotopes of bromine have an atomic mass of 80.92 u.

   Calculate the average atomic mass of bromine, showing all work:

2. Using the following data, calculate the average atomic mass of magnesium (give your answer to the nearest .01 u). Show all work!

   Isotope: ${}^{24}_{12}Mg$ Percent abundance: 78.70%

   Isotope: ${}^{25}_{12}Mg$ Percent abundance: 10.13%

   Isotope: ${}^{26}_{12}Mg$ Percent abundance: 11.17%

3. Using the periodic table:

   What is the average atomic mass of bromine? ____________

   What is the average atomic mass of magnesium? __________

   How do your calculated answers in #1 and #2 compare to those on the periodic table?

Lesson 31: Electrons in the Atom Notes
Chemistry with Lab

Bohr's Energy Levels

- Electrons in _____________ ________________ __________________
- _______________ energy levels: ____________ to ____________
- _______________ energy levels: ____________ from _______________
- Ground State: ____________ in _____________ __________________

  _______________ possible

Excited Atom

- Atom has ____________ __________________
- __________ state is ____________
- __________ soon __________ same amount of __________ _________
- __________ seen as _____________ __________________

Wave Description of Light:

Wavelength (_________): _____________ between ____________ _________ on ____________ waves

Frequency (__________): the __________ of ____________ passing a given _______________ in a given ________________

c = ____________

c = _________________: speed of ________________

Lesson 31: Electrons in the Atom Notes (cont.) Chemistry with Lab

Sample problem 1:
What is the frequency of light if the wavelength is 6.0×10^{-7}m?

Sample problem 2:
What is the wavelength of light if its frequency is 5.0×10^{14} Hz?

Particle Description of Light

_______________ exists as _________________ called _______________

E = _________________

The Modern View of Light

________________ has a _______________ _______________

- Light may _____________ as a ________________
- Light may _____________ as a ________________ of _____________ called ____________ or _____________

Spectroscopy

- ____________ lines represent _____________ ___________ as ___________ returns to _____________ _______________ _______________

- ___________ lines _______________ an ____________
- Called the ____________ ___________ ____________ of an __________

Orbital

_______________ of ____________ where an ___________ is ____________

to be _______________

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Quantum Numbers

- ______________________________
- Used to _____________ an ____________ in an _______________

n

- _____________ ______________ ______________
- Represents _____________ energy level of ____________

  _______________ # of _______________ in an

  _______________ ______________ = __________________

Example: What is the maximum number of electrons that can be in the __________ main energy level?

l

- The ______________ ______________ ______________
- Describes the ____________ _____________ within an ____________ _____________
- _____________ of orbital ____________ possible in _______________ _____________ = _____________

Orbital Shapes

designated _______________

- level 1: _____________
- level 2: _____________
- level 3: _____________
- level 4: _____________

How many electrons can each sublevel hold?

s = 1 orbital x 2 e/orbital = ___________ e

p = 3 orbitals x 2 e/orbital = __________ e

d = 5 orbitals x 2 e/orbital = __________ e

f = 7 orbital x 2 e/orbital = ___________ e

m

- The ______________ ______________ ______________
- describes ______________ of _________________ in ____________

s

- The ______________ ______________ ______________
- describes ______________ of _________________ in ____________

Ground State: ___________ energy arrangement of ______________

Diagonal Rule:

Examples:

hydrogen ________________ lithium __________________

nitrogen _________________

Orbital Notation:

Examples:

hydrogen

nitrogen

Lesson 33: Electron Distribution Notes (cont.) Chemistry with Lab

Hund's Rule:

_______________ of ____________ ______________ are each ____________

by one ______________ before any _______________ is occupied by a

_______________ ___________________.

Pauli Exclusion Principle:

No two ______________ in the ______________ ______________ can have the

_______________ ______________ of ____________ ______________

_______________.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 33: Energy Levels Chart

Energy Level n	E Sublevel (type of orbital)	# of Orbitals in Sublevel	# of e^- in Sublevel	Total # of e^- in E level ($2n^2$)
1				
2				
3				
4				

Each orbital can hold a maximum of _____________ electrons.

Since both electrons have a ___________ charge, they _____________.

What keeps them from flying apart? Each electron ____________ on its axis.

One spins ________________ and the other spins counter-clockwise. When charged particles spin, they act like tiny magnets. Since the two electrons spin in _______________ directions, one acts like the north pole of a magnet and the other acts like the south pole.

This makes the electrons _____________.

Since each orbital can hold ____________ electrons:

The "s" sublevel can hold _____________ electrons.

The "p" sublevel can hold _____________ electrons.

The "d" sublevel can hold _____________ electrons.

The "f" sublevel can hold ______________ electrons.

The orbital shaped like a "dumbbell" is the _____________ orbital, while the orbital shaped spherically is the _____________ orbital.

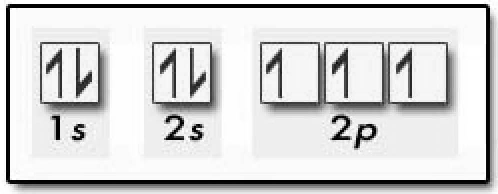
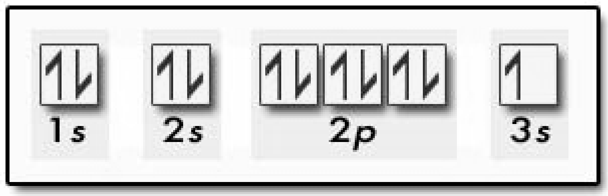
How many sublevels are present in the third main energy level? ___________

What is the maximum number of orbitals in the "d" sublevel? _____________

The maximum number of electrons that can occupy an orbital is___________, provided they have _____________ spins.

The likely location of an electron within the atom is a(n) _____________.

Lesson 35: Orbital Diagram Chart

<table>
<tr><th>Element</th><th>Orbital Diagram</th></tr>
<tr><td></td><td>
</td></tr>
<tr><td></td><td>
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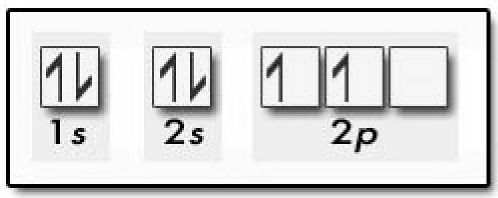
<tr><td></td><td>
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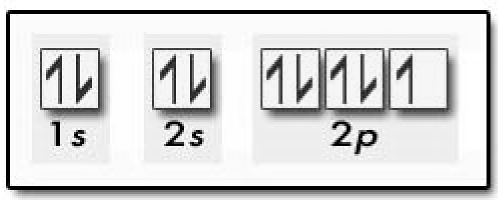
<tr><td></td><td>
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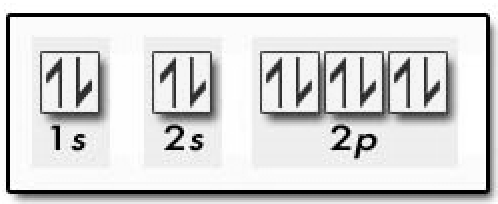
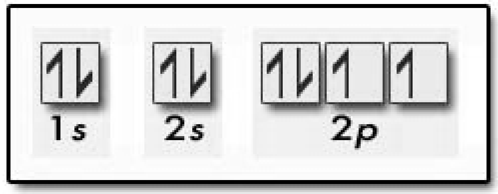
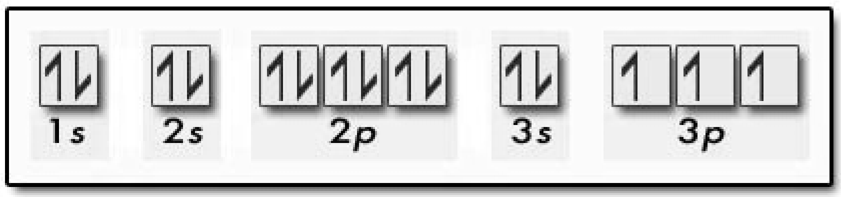
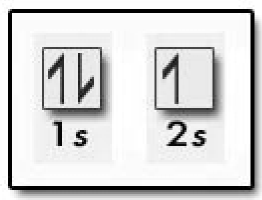
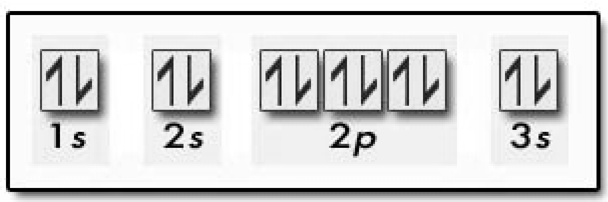
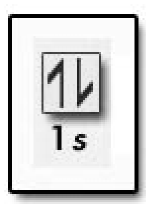
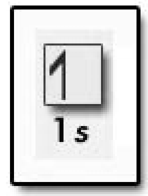
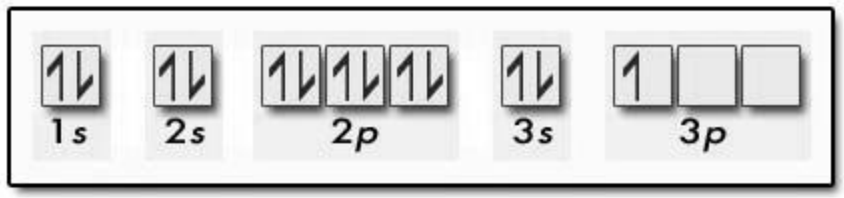
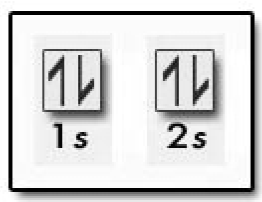
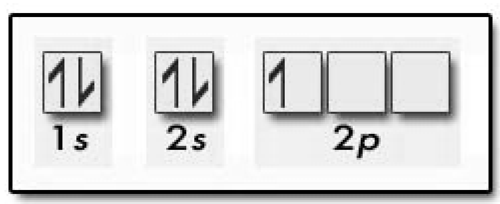
<tr><td></td><td>
</td></tr>
</table>

Element	Orbital Diagram
	
	
	
	
	

Lesson 35: Orbital Diagram Chart (cont.)

Element	Orbital Diagram
	
	
	
	

1. **Mendeleev** – (born 1834— died 1907) Russian chemist who created the first valid periodic table of the elements in 1869
2. **Moseley** – (born 1887— died 1915) English physicist who used X-ray diffraction to show that each element has an atomic number. He arranged the elements by increasing atomic number, which was almost the same as Mendeleev's arrangement in order of increasing atomic weight, but differs slightly. These differences account for the few discrepancies inherent in the Mendeleev system. He was killed in World War I.
3. **Periodic law** – The principle that the properties of the elements recur periodically as their atomic numbers increase.
4. **Family** – A vertical column in the periodic table of elements.
5. **Group** – A vertical column in the periodic table of elements.
6. **Period** – A sequence of elements arranged in order of increasing atomic number and forming one of the horizontal rows in the periodic table.
7. **Electron distribution** – A function which gives the number of electrons per unit volume of phase space.
8. **Valence electron** – An electron in an outer shell of an atom that can participate in forming chemical bonds with other atoms.
9. **Noble gas distribution** – A short hand of electron distribution. Constructed by putting the symbol of the noble gas in the period before the element in brackets and continuing the electron configuration from where the noble gas left off.
10. **Energy level** – One of the allowed values of the internal energy of an isolated physical system.
11. **Stable** – Not easily decomposed or otherwise modified chemically.
12. **Metalloid** – An element which exhibits the external characteristics of a metal but behaves chemically both as a metal and a nonmetal.
13. **Semimetal** – Another name for an element which exhibits the external characteristics of a metal but behaves chemically as a metal and a nonmetal.
14. **Atomic radius** – Half the distance between the nuclei of two like atoms that are covalently bonded.
15. **Electron affinity** – The work needed in removing an electron from a negative ion, thus restoring the neutrality of an atom or molecule.

16. **Ionization energy** – The amount of energy needed to remove an electron from a given kind of atom or molecule to an infinite distance.
17. **Electron configuration** – The orbital and spin arrangement of an atom's electrons, specifying the quantum numbers of the atom's electrons in a given state.

Lesson 39: History of the Periodic Table Notes Chemistry with Lab

Dmitri Mendeleev

- Arranged elements by ___________ ______________ ______________
- predicted ____________ of missing ________________

Henry Moseley

- discovered that each ____________ has a unique _____________ ____________________
- arranged elements by ___________ ___________________ ____________________
- now all elements fit into place based on properties

Modern Periodic Law

- Properties of elements ___________ when elements are arranged by _______________ ______________ ______________

Questions about the extended periodic table

Why do you think these elements get pulled out?

Which element is in a different column than we are used to?

What happens to the number of the energy level as you begin each new row?

What is the similar characteristic in each column?

Lesson 39: Periodic Table Notes (cont.)

Elements in the same column

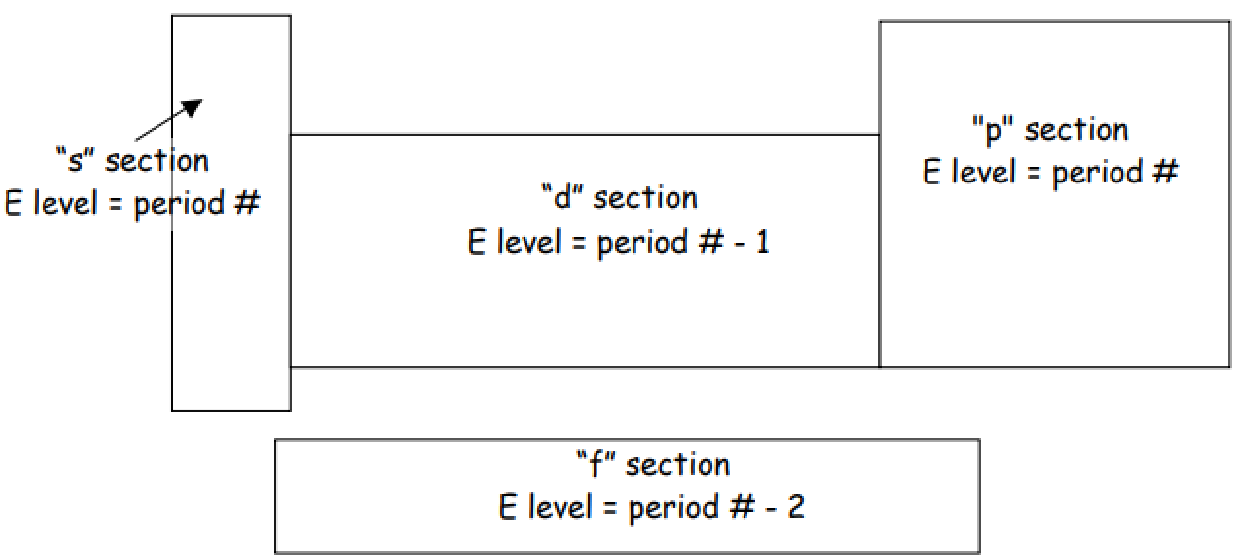
- have similar ________________ _________________________
- have same number of _________________ _________________________
- have similar _______________ _________________________
- Electron distribution for Ne: ____________________________

  Kr: ____________________________

Noble Gases

- have ____________ outer ____________________ ____________________
- have ________________ valence _______________________
- are __________________

Using periodic table to find electron distributions



Noble Gas Distributions – Find noble gas on end of row before the element and put its symbol in brackets.

Use diagonal rule to write electron distribution of Sn (atomic no. 50):

____

Use noble gas distribution to write electron distribution of Sn:

____

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 41: Organization of the Periodic Table Chemistry with Lab

Period

- _______________ on ________________ ___________________
- tells __________________ _____________________

Family

- _______________ on the __________________ ___________________
- for columns 1-2 and 13-18, the number in the one's place represents

  number of ________________ __________________

Metalloids

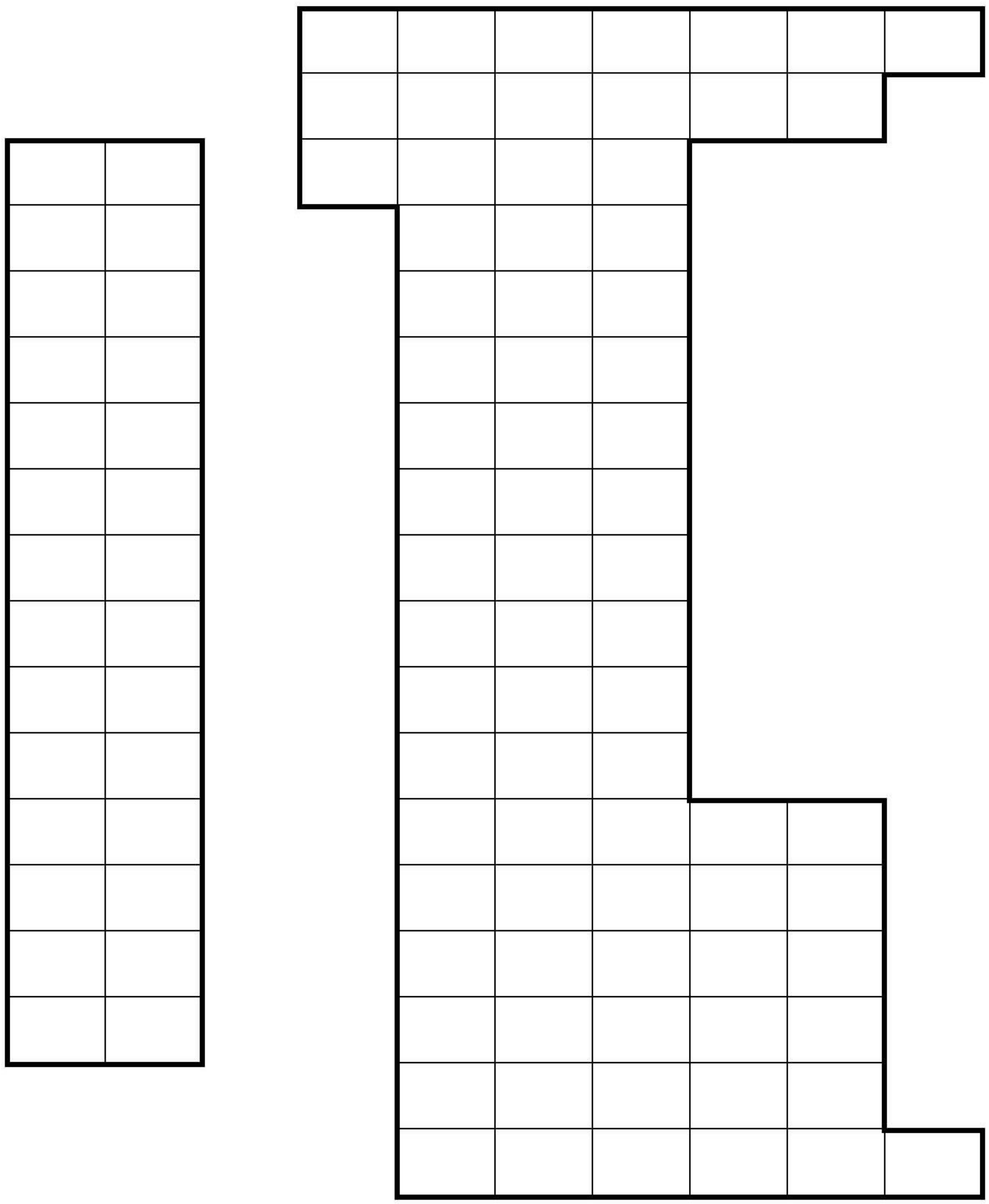
- have properties of both _______________ and ____________________
- border ________________ line (except Al, which is a ________________)

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 41: Blank Periodic Table



Lesson 42: Matching Activity

A family	A row on the periodic table
A period	An element that shares characteristics of both metals and non-metals
A metalloid	A column on the periodic table

Family that contains Bromine	Alkaline earth metals
Family that contains Argon	Actinides
Family that contains Beryllium	Halogens
Family that contains Uranium	Noble gases

Te	8 valence electrons
P	6 valence electrons
Bi	2 valence electrons
Sn	4 valence electrons
S	5 valence electrons
Al	3 valence electrons
Ne	6 valence electrons
Cs	1 valence electron
Br	5 valence electrons
Mg	7 valence electrons

Lesson 44: Trends on the Periodic Table

Atomic Number

- _____________ across a period
- _____________ down a family

Atomic Mass

- _____________ across a period
- _____________ down a family

Atomic Radius

- __________ the distance between ___________ of two like ____________.
- _________________ across a period (due to increased attraction of more positive nuclei to electron clouds)
- _________________ down a family (due to addition of an energy level)

Electron Affinity

- amount of ___________ ______________ when an _____________ gains an __________________
- _______________ across a period
- _______________ down a family

Ionization Energy

- amount of ___________ required to ____________ an _______________ from a neutral _____________
- _______________ across a period
- _______________ down a family

Electron Distribution – Forming Ions

- Atoms become ions by ________________ or _____________ electrons to become ___________________
- An atom is stable when it has a _____________ outer energy level (_________ valence electrons)
- Atoms that lose (give away) electrons become ___________ charged ions. Atoms that gain (steal) electrons become __________ charged ions.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Atom						Ion			
element	# of protons	# of electrons	ending electron distribution	# of valence e-	family #	to become stable	# of protons	# of electrons	ion charge
Li	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
Be	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
B	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
C	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
N	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
O	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
F	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	
Ne	___P+	_____e-				lose/gain _______ e-	____P+	_____e-	

In this activity, you will explore the trends of atomic radius vs. atomic number and ionization energy vs. atomic number. Using the table of values below, construct two graphs using Graphical Analysis showing the property vs. atomic number. Allow the program to connect the points in a dot-to-dot manner. Do NOT use the regression line or best fit curve keys on these graphs.

As an alternate graph, you could put the two graphs on one by using the GA program and ADDING A NEW DATA SET with the second set of data. If you do this, you will have to double-click on the graph after both sets of data have been entered. Then, click on AXES OPTIONS and click on the + sign next to the second data set. Open it up and click the second axis in that set. (The checked box should agree with Data Set 1.) Close this window.

Element Symbol	Atomic Number	Atomic Radius (Angstroms)	Ionization Energy (Volts)
H	1	0.79	13.6
He	2	0.49	24.587
Li	3	2.05	5.392
Be	4	1.4	9.322
B	5	1.17	8.298
C	6	0.91	11.26
N	7	0.75	14.534
O	8	0.65	13.618
F	9	0.57	17.422
Ne	10	0.51	21.564
Na	11	2.23	5.139
Mg	12	1.72	7.646
Al	13	1.82	5.986
Si	14	1.46	8.151
P	15	1.23	10.486
S	16	1.09	10.36
Cl	17	0.97	12.967
Ar	18	0.88	15.759
K	19	2.77	4.341
Ca	20	2.23	6.113
Ga	31	1.81	5.999
Ge	32	1.52	7.899
As	33	1.33	9.81
Se	34	1.22	9.752

Element Symbol	Atomic Number	Atomic Radius (Angstroms)	Ionization Energy (Volts)
Br	35	1.12	11.914
Kr	36	1.03	13.999
Rb	37	2.98	4.177
Sr	38	2.45	5.695
In	49	2.00	5.786
Sn	50	1.72	7.344
Sb	51	1.53	8.641
Te	52	1.42	9.009
I	53	1.32	10.451
Xe	54	1.24	12.13

1. Looking carefully at the data, what elements have been left out? Why do you think they have been omitted?

2. Describe the general trends in atomic radius for a period and for a group.

3. Describe the general trends in ionization energy for a period and for a group.

4. Place the one graph on top of the other and describe the differences in trends for the two properties.

5. Why is there a dip in ionization energies between the elements Be and B, between Mg and Al, and between Ca and Ga? (Think electron distributions!)

Lesson 50: Table of Electronegativities

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.1																	2 **He** –
2	3 **Li** 1.0	4 **Be** 1.5											5 **B** 2.0	6 **C** 2.5	7 **N** 3.0	8 **O** 3.5	9 **F** 4.0	10 **Ne** –
3	11 **Na** 0.9	12 **Mg** 1.2											13 **Al** 1.5	14 **Si** 1.8	15 **P** 2.1	16 **S** 2.5	17 **Cl** 3.0	18 **Ar** –
4	19 **K** 0.8	20 **Ca** 1.0	21 **Sc** 1.3	22 **Ti** 1.5	23 **V** 1.6	24 **Cr** 1.6	25 **Mn** 1.5	26 **Fe** 1.8	27 **Co** 1.8	28 **Ni** 1.8	29 **Cu** 1.9	30 **Zn** 1.6	31 **Ga** 1.6	32 **Ge** 1.8	33 **As** 2.0	34 **Se** 2.4	35 **Br** 2.8	36 **Kr** 3.0
5	37 **Rb** 0.8	38 **Sr** 1.0	39 **Y** 1.2	40 **Zr** 1.4	41 **Nb** 1.6	42 **Mo** 1.8	43 **Tc** 1.9	44 **Ru** 2.2	45 **Rh** 2.2	46 **Pd** 2.2	47 **Ag** 1.9	48 **Cd** 1.7	49 **In** 1.7	50 **Sn** 1.8	51 **Sb** 1.9	52 **Te** 2.1	53 **I** 2.5	54 **Xe** 2.6
6	55 **Cs** 0.7	56 **Ba** 0.9	*71 **Lu** 1.3	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.8	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn** 2.4
7	87 **Fr** 0.7	88 **Ra** 0.9	**103 **Lr** –															

6	* 57 **La** 1.1	58 **Ce** 1.1	59 **Pr** 1.1	60 **Nd** 1.1	61 **Pm** 1.1	62 **Sm** 1.2	63 **Eu** 1.1	64 **Gd** 1.2	65 **Tb** 1.1	66 **Dy** 1.2	67 **Ho** 1.2	68 **Er** 1.2	69 **Tm** 1.3	70 **Yb** 1.1
7	**89 **Ac** 1.1	90 **Th** 1.3	91 **Pa** 1.5	92 **U** 1.4	93 **Np** 1.4	94 **Pu** 1.3	95 **Am** 1.3	96 **Cm** 1.3	97 **Bk** 1.3	98 **Cf** 1.3	99 **Es** 1.3	100 **Fm** 1.3	101 **Md** 1.3	102 **No** 1.3

Lesson 50: Bonding Key Terms

1. **Chemical bond** – Any of several forces, especially the ionic bond, covalent bond, and metallic bond, by which atoms or ions are bound in a molecule or crystal.
2. **Ionic bond** – A chemical bond between two ions with opposite charges, characteristic of salts.
3. **Covalent bond** – A chemical bond formed by the sharing of one or more electrons, especially pairs of electrons, between atoms.
4. **Metallic bond** – The chemical bond characteristic of metals, in which mobile valence electrons are shared among atoms in a usually stable crystalline structure.
5. **Hydrogen bond** – A chemical bond in which a hydrogen atom of one molecule is attracted to an electronegative atom, especially a nitrogen, oxygen, or fluorine atom, usually of another molecule.
6. **Delocalized** – electrons in a molecule or solid metal that are not associated with a single atom or *one* covalent bond.
7. **Crystal lattice** – A geometric arrangement of the points in space at which the atoms, molecules, or ions of a crystal occur.
8. **Luster** – A radiant brightness or glow, usually due to light reflected from a smooth surface.
9. **Malleable** – Capable of being shaped or formed, as by hammering or pressure.
10. **Ductile** – Easily drawn into wire or hammered thin.
11. **Volatile** – Evaporating readily at normal temperatures and pressures.
12. **Ion** – An atom or a group of atoms that has acquired a net electric charge by gaining or losing one or more electrons.
13. **Cation** – An ion or group of ions having a positive charge and characteristically moving toward the negative electrode in electrolysis.
14. **Anion** – A negatively charged ion, especially the ion that migrates to an anode in electrolysis.
15. **Intermolecular force** – The force between two molecules; it is that negative gradient of the potential energy between the interacting molecules, if energy is a function of the distance between the centers of the molecules.

16. **London-dispersion force** – a weak intermolecular force arising from quantum induced instantaneous polarization multipoles in molecules.
17. **Polar** – Relating to or characterized by a dipole.
18. **Nonpolar** – Not characterized by a dipole.
19. **Electron dot diagram** – A structural formula in which electrons are represented by dots; two dots between atoms represent a covalent bond. Also known as electron-dot formula; Lewis formula.
20. **Lewis structure** – Can be synonymous with electron dot diagram; further depicts how electrons are shared in bonds between atoms within a molecule.

Lesson 50: Introduction to Bonding Notes

Chemistry with Lab

Chemical bond – forms when 2 or more atoms rearrange _______________ _______________ to increase ___________________.

Ionic bond – forms when valence _______________ are ________________ from one atom to another.

- cation – atom _____________ electrons to become ___________ charged
- anion – atom _____________ electrons to become ___________ charged
- In ionic compounds, the ions are arranged in a _______________ and _______________ ______________ forces hold the ions together.

Properties of ionic compounds:

- high ________________ and ____________________ points
- _________________ - not easily __________________
- ____________________ electricity when ________________ or _______________ because the ions are free to _________________.

Covalent bond - _______________ are ________________, forming ____________________.

- Covalent compounds have __________________ forces holding the ____________________ together.

Properties of covalent compounds:

- Lower _______________ and ________________ points

- Many covalent compounds are ___________________ liquids or gases.
- _______________ - easier to _______________
- are not ___________________ of electricity

Electronegativity – property that tells how strong an atom's ____________ is for ____________________.

- Since oxygen has a _____________ electronegativity than hydrogen, oxygen holds onto shared electrons ______________, giving the oxygen a _____________ negative charge and the hydrogen a partial _______________ charge.

polar covalent bonds – electrons are shared _______________, creating partially charged ends or _____________.

nonpolar covalent bonds – electrons are shared ________________ because atoms have the same electronegativities

Electronegativity difference	Type of Bond
greater than or equal to 1.7	
between 1.7 and 0.3	
less than or equal to 0.3	

Examples:

Mg and F? S and O?

Program 501, problem set 1:

<u>metallic bond</u> – electrons are __________________

(crates a "______ of ____________")

properties of metals:

1.

2.

3.

4.

<u>The Chemistry Quiz</u>

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 51: Ionic or Covalent Data Table

	Hardness test: Hard or soft?	Ionic or Covalent due to hardness?	Ionic or Covalent due to nature of formula?	Step 4 observations	Step 5 observations	Step 6 observations
Table Salt						
Baking Soda						
Butter, Margarine, or Shortening						
Solid Air Freshener						
Vegetable Oil						
Water						
Soda						
Rubbing Alcohol						
Vinegar						

Lesson 51: Ionic or Covalent Data Table (cont.)

	Step 7 observations	Step 8 observations?	Step 9 observations	Step 10 observations	Ionic or Covalent?
Table Salt					
Baking Soda					
Butter, Margarine, or Shortening					
Solid Air Freshener					
Vegetable Oil					
Water					
Soda					
Rubbing Alcohol					
Vinegar					

Lesson 52: Bonding Part II Notes

Oxygen:

symbol _________ atomic number _____________

___________ protons ___________ electrons

Oxygen has __________ valence electrons

Electron Dot Diagram – atom's _______________ surrounded by _____________ to represent its _____________ electrons

Example electron dot diagrams:

O Li

Problem Set 1:

Lewis Structure: diagram representing the arrangement of ________________ electrons in a ____________________.

Most atoms need _____ valence electrons to become stable. The exceptions are H and He which only need ______ valence electrons to be stable.

Lesson 52: Bonding Part II Notes (cont.)

Lewis structure for H_2 H - H

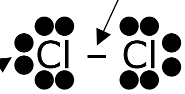
shared pair

- 2 electrons belonging to both _______
- represented by a ________ between symbols

Lewis structure for Cl_2:

Each CL atom has _____ valence electrons, giving a total of _____ valence electrons to work with.

__________ pair



unshared pair

- electrons belonging to only one ______
- represented by 2 dots

Lewis structure for HCl:

H Cl

When more than two atoms bond, you must determine which is central.

The central atom is:

- frequently _________
- never ________
- often atom with ______________ electronegativity

Lewis structure for CH_3I:

(There are a total of ________
valence electrons to work with.)

Problem Set 2:

Lewis structure of ethene, C_2H_4 (has total of ________ valence electrons)

H H

C C

H H

type of bond	pairs of electrons shared

Lesson 52: Bonding Part II Notes (cont.)

Problem Set 3:

Polyatomic Ion: _______________ bonded group of ions with a ___________

Example: NO_3^{1-} (has gained _____ electron to give a total of ________ valence electrons to work with)

O

O N

O

Problem Set 4:

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Electron Dot Diagrams

magnesium		Mg
helium	He	
krypton		Kr
aluminum	Al	
bromine		Br
potassium	K	
nitrogen		N
silicon	Si	
strontium		Sr

Draw Lewis structures for:

NH_3

CH_2Cl_2

CH_3OH

C_6H_{12}

Draw Lewis structures for:

CO_2

HCN

CH_2O

C_2HCl

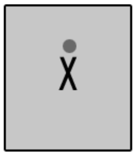
Draw Lewis structures for:

PO_4^{3-}

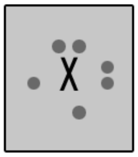
SO_4^{2-}

NH_4^{1+}

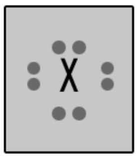
CN^{1-}



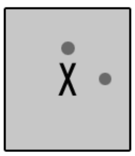
Boron



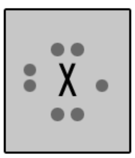
Neon



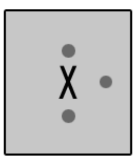
Oxygen



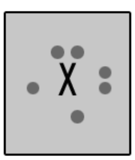
Lithium



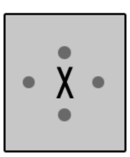
Magnesium



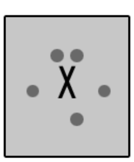
Phosphorus



Sulfur



Iodine



Carbo

Lesson 57: Ionic Charges Chart

Cations

1+		2+		3+		4+	
ammonium	NH_4^+	barium	Ba^{2+}	aluminum	Al^{3+}	tin(IV)	Sn^{4+}
cesium	Cs^+	beryllium	Be^{2+}	chromium(III)	Cr^{3+}	nickel (IV)	Ni^{4+}
gold(I)	Au^+	cadmium	Cd^{2+}	cobalt(III)	Co^{3+}	lead(IV)	Pb^{4+}
hydrogen	H^+	calcium	Ca^{2+}	gold(III)	Au^{3+}		
lead(I)	Pb^+	cobalt(II)	Co^{2+}	iron(III)	Fe^{3+}		
lithium	Li^+	copper(II)	Cu^{2+}	manganese(III)	Mn^{3+}		
potassium	K^+	iron(II)	Fe^{2+}				
silver	Ag^+	lead(II)	Pb^{2+}				
sodium	Na^+	magnesium	Mg^{2+}				
copper(I)	Cu^+	manganese(II)	Mn^{2+}				
		mercury(I)	Hg_2^{2+}				
		mercury(II)	Hg^{2+}				
		nickel(II)	Ni^{2+}				
		strontium	Sr^{2+}				
		zinc	Zn^{2+}				
		tin(II)	Sn^{2+}				

Roman numeral notation indicates charge of ion when element commonly forms more than one ion.

Anions

1-		2-		3-	
acetate	$C_2H_3O_2^-$	carbonate	CO_3^{2-}	arsenate	AsO_4^{3-}
amide	NH_2^-	chromate	CrO_4^{2-}	arsenite	AsO_3^{3-}
bicarbonate*	HCO_3^-	dichromate	$Cr_2O_7^{2-}$	citrate	$C_6H_5O_7^{3-}$
bisulfate**	HSO_4^-	oxide	O^{2-}	ferricyanide	$Fe(CN)_6^{3-}$
bisulfide	HS^-	oxalate	$C_2O_4^{2-}$	nitride	N^{3-}
bisulfite	HSO_3^-	silicate	SiO_3^{2-}	phosphate	PO_4^{3-}
bromate	BrO_3^-	sulfate	SO_4^{2-}	phosphite	PO_3^{3-}
bromide	Br^-	sulfide	S^{2-}	phosphide	p^{3-}
chlorate	ClO_3^-	sulfite	SO_3^{2-}		
chlorite	ClO_2^-	tartrate	$C_4H_4O_6^{2-}$		
chloride	Cl^-	tetraborate	$B_4O_7^{2-}$		
cyanide	CN^-	thiosulfate	$S_2O_3^{2-}$		
cyanate	OCN^-				
fluoride	F^-				
hydride	H^-				
hydroxide	OH^-				
hypochlorite	ClO^-				
iodate	IO_3^-				
iodide	I^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
perchlorate	ClO_4^-				
permanganate	MnO_4^-				
thiocyanate	SCN^-				

There are no common anions with a 4- charge.

*hydrogen carbonate

**hydrogen sulfate

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Lesson 57: Chemical Formulas Terms

1. **Ionic compound** – a chemical compound in which ions are held together in a lattice structure by ionic bonds.
2. **Molecular compound** – The smallest particle of a substance that retains the chemical and physical properties of the substance and is composed of two or more atoms bonded together by the sharing of electrons.
3. **Subscripts** – A distinguishing character or symbol written directly beneath or next to and slightly below a letter or number. In chemical formula writing, the subscript denotes how many atoms or ions of a particular element or polyatomic ion are present.
4. **Nomenclature** – A system of naming chemical compounds and for describing the science of chemistry in general. It is maintained by the International Union of Pure and Applied Chemistry (IUPAC).
5. **Polyatomic ions** – An electrically charged species formed by covalent bonding of atoms of two or more different elements, usually nonmetals, for example, the ammonium ion($NH4^{+}$).
6. **Reactant** – A substance participating in a chemical reaction, especially a directly reacting substance present at the initiation of the reaction.
7. **Product** – A substance resulting from a chemical reaction.
8. **Law of Conservation of Mass** – The notion that mass, or matter, can neither be created nor destroyed.
9. **Coefficient** – A number placed in front of a term in a chemical equation to indicate how many molecules or atoms take part in the reaction.
10. **Precipitate** – To be separated from a solution as a solid.
11. **Aqueous** – A solution dissolved in water.
12. **Synthesis reaction** – A direct combination reaction in which two or more reactants combine to form a single product. The general form is: $Ax + B \rightarrow AB$.
13. **Decomposition reaction** – A chemical reaction in which a compound is broken down into simpler compounds, or even into elements. This is the opposite of a synthesis or direct combination reaction. The general form is: $AB \rightarrow A + B$.
14. **Single replacement reaction** – A chemical reaction in which an element replaces one element in a compound. A single uncombined element replaces another in a compound. Two reactants yield two products. The general form is: $A + BC \rightarrow B + AC$.

15. **Double replacement reaction** – A molecular process involving the exchange of bonds between two reacting chemical species, which results in the creation of products with similar or identical bonding affiliations. Also known as a metathesis reaction. The general form is: $AX + BY \rightarrow BX + AY$.
16. **Combustion reaction** – The burning of any substance, in gaseous, liquid, or solid form. A chemical reaction that involves the rapid combination of a fuel with oxygen. The general form is: fuel + oxygen → heat + water + carbon dioxide.
17. **Activity series** – a series of elements that have similar properties, for example, metals, arranged in descending order of chemical activity.
18. **Hydrocarbon** – Any of numerous organic compounds, such as benzene and methane, that contain only carbon and hydrogen.

Lesson 57: Ionic Formula Notes

Chemical Formula

Tells the _____________ number of atoms of each element in a ___________.

Molecular: (___________________ bonding)

Formula represents the number of ________________ of each

________________ in a single ________________.

Ionic: (___________________ bonding)

Formula represents the ________________ _______________ of

positive and negative _________________ in one ____________________

________________.

Ionic Bond

The _________________ that binds ________________ charged _____________

together.

Metals

- _____________________ ____________________ electrons
- ions are ___________________ charged
- _____________________

Nonmetals

- _____________________ electrons
- ions are ___________________ charged
- _____________________

Lesson 57: Ionic Formula Notes (cont.)

Writing Ionic Formulas

1. Write the ________________ for the _________________.
   - The ______________ is written first.
   - The ______________ is written second.
2. Determine the _________________ on each ion.
3. Select ______________ that will make the __________ ____________ charge equal to the _____________ _____________ charge.

examples: sodium chloride   magnesium chloride

Criss-Cross Method for Formula Writing

- Write symbol of ___________ followed by symbol of _______________ along with their ________________ ______________.
- Use the _____________ ______________ of the charge of each ion as the _________________ for the other.
- If the subscripts are the same, _____________ them.
- Subscripts must be simplified in an ionic formula.

Examples:

Ca^{2+} O^{2-} K^{1+} ClO_3^{1-} Al^{3+} SO_4^{2-}

__________   __________   __________

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 59: Ionic Formula Practice

In each box, write the formula of the ionic compound consisting of the positive ion to the left of the box and the negative ion above the box.

	CL^{-}	**S^{2-}**	**F^{-}**	**N^{3-}**	**O^{2-}**	**P^{3-}**
MG^{2+}	$MGCL_2$					
CS^{+}						
CR^{3+}						
NA^{+}						
ZN^{2+}						
AL^{3+}						

Ternary Compounds

	OH^{-}	**SO_4^{2-}**	**NO_3^{-}**	**CO_3^{2-}**	**PO_4^{3-}**
MG^{2+}					
K^{+}					
AL^{3+}					
NH_4^{+}					
CA^{2+}					
FE^{3+}					

Write the formula of the compound formed from the following pairs of ions.

1. Li^{+}, O^{2-}

   ____

2. Ag^{+}, OH^{-}

   ____

3. Mg^{2+}, SO_4^{2-}

   ____

4. NH_4^{+}, CO_3^{2-}

   ____

5. Ca^{2+}, P^{3-}

   ____

6. Sn^{4+}, O^{2-}

   ____

7. Sn^{4+}, IO_3^{-}

   ____

8. Cu^{+}, NO_3^{-}

   ____

9. Mn^{2+}, S^{2-}

   ____

10. K^{+}, MnO_4^{-}

    ____

11. Al^{3+}, SO_4^{2-}

    ____

12. H^{+}, PO_4^{3-}

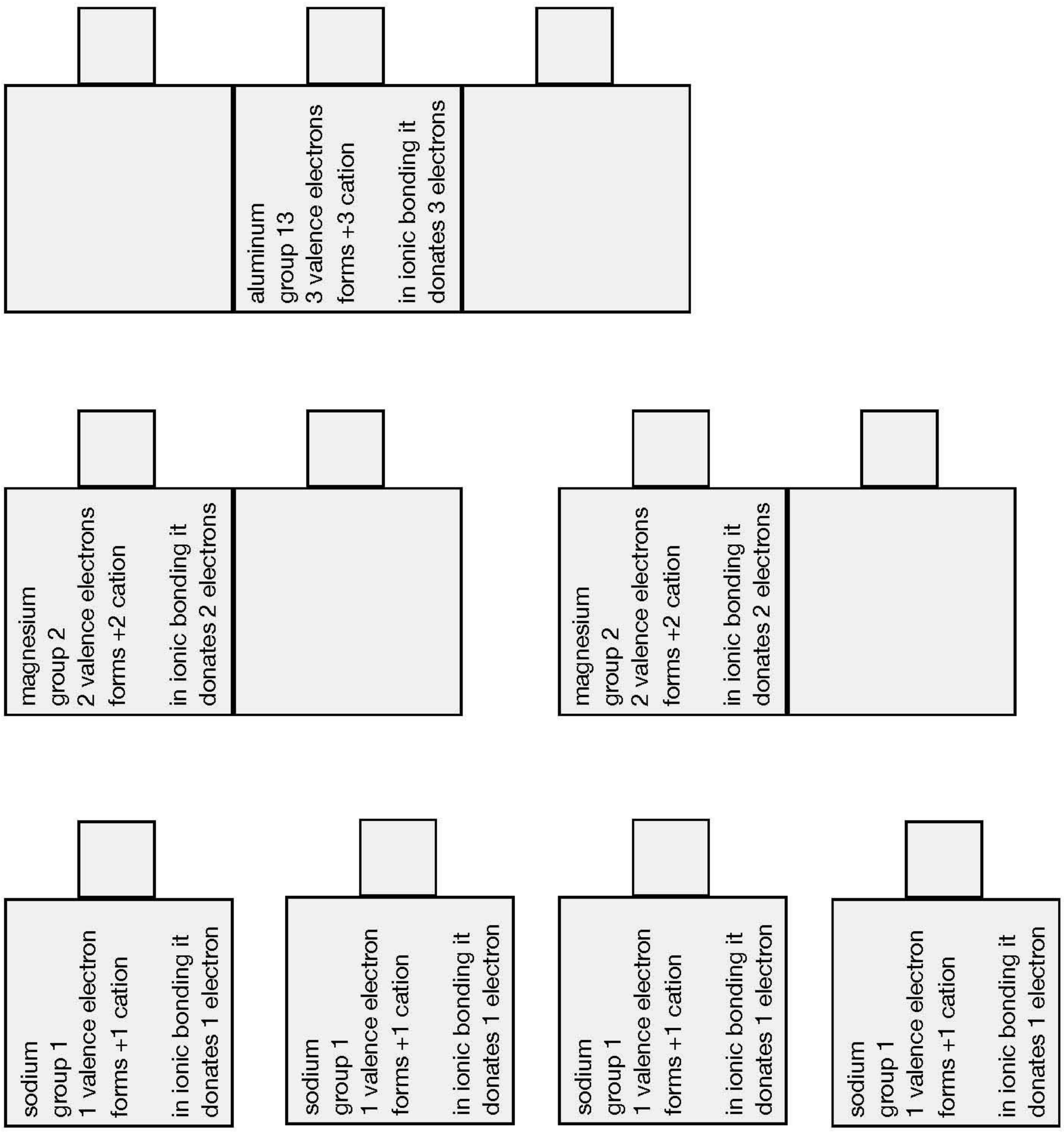
    ____

13. $Hg2^{+}$, ClO_4^{-}

    ____

14. Fe^{3+}, CN^{-}

    ____



aluminum
group 13
3 valence electrons
forms +3 cation
in ionic bonding it
donates 3 electrons
magnesium
group 2
2 valence electrons
forms +2 cation
in ionic bonding it
donates 2 electrons
magnesium
group 2
2 valence electrons
forms +2 cation
in ionic bonding it
donates 2 electrons
sodium
group 1
1 valence electron
forms +1 cation
in ionic bonding it
donates 1 electron
sodium
group 1
1 valence electron
forms +1 cation
in ionic bonding it
donates 1 electron
sodium
group 1
1 valence electron
forms +1 cation
in ionic bonding it
donates 1 electron
sodium
group 1
1 valence electron
forms +1 cation
in ionic bonding it
donates 1 electron

aluminum
group 13
3 valence electrons
forms +3 cation

in ionic bonding it
donates 3 electrons

aluminum
group 13
3 valence electrons
forms +3 cation

in ionic bonding it
donates 3 electrons

magnesium
group 2
2 valence electrons
forms +2 cation

in ionic bonding it
donates 2 electrons

magnesium
group 2
2 valence electrons
forms +2 cation

in ionic bonding it
donates 2 electrons

sodium
group 1
1 valence electron
forms +1 cation

in ionic bonding it
donates 1 electron

sodium
group 1
1 valence electron
forms +1 cation

in ionic bonding it
donates 1 electron

magnesium
group 2
2 valence electrons
forms +2 cation

in ionic bonding it
donates 2 electrons

nitrogen
group 15
5 valence electrons
forms -3 anion
in ionic bonding it
receives 3 electrons

nitrogen
group 15
5 valence electrons
forms -3 anion
in ionic bonding it
receives 3 electrons

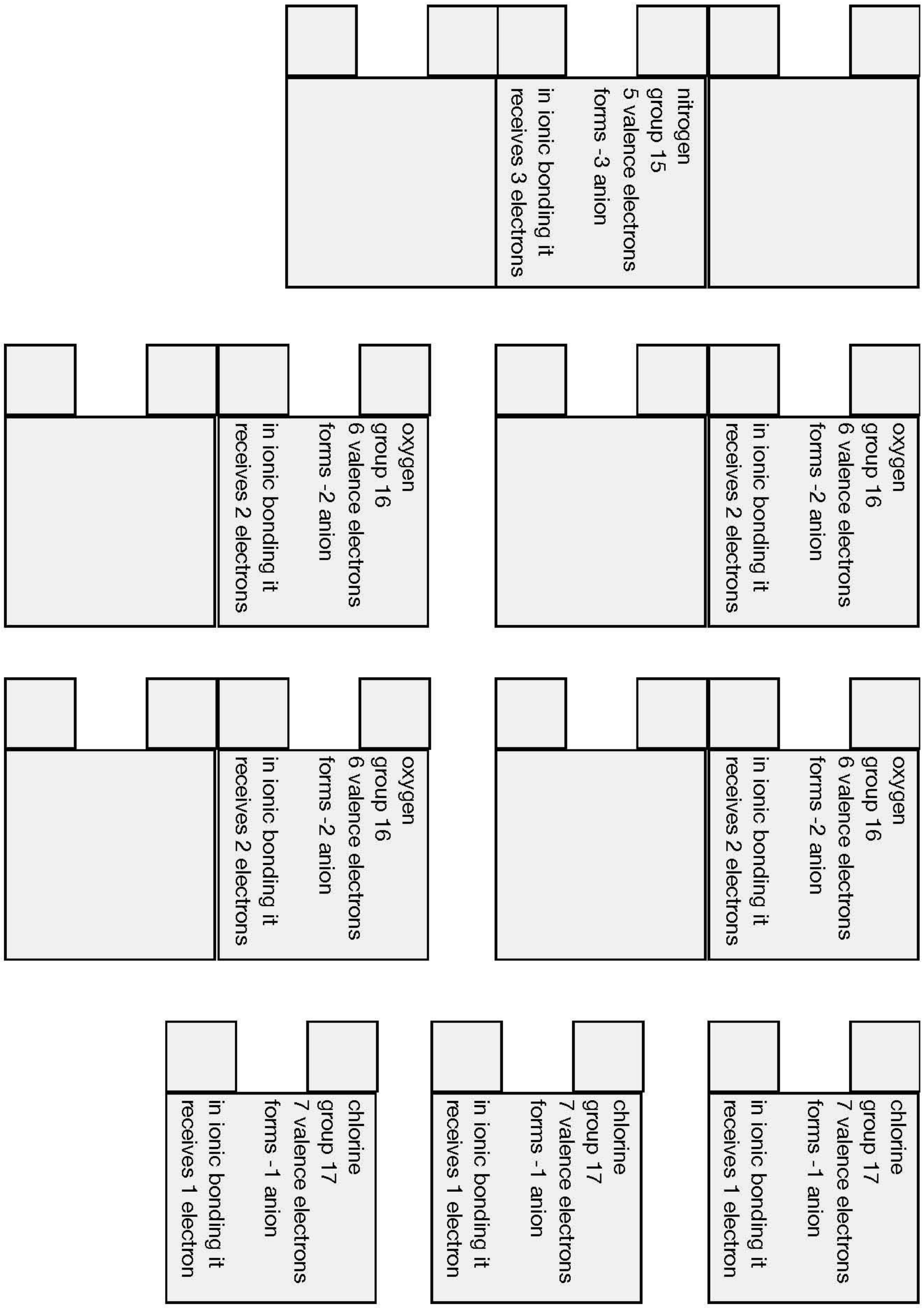
oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons

oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons

chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron

chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron

chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron



nitrogen
group 15
5 valence electrons
forms -3 anion
in ionic bonding it
receives 3 electrons
oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons
oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons
oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons
oxygen
group 16
6 valence electrons
forms -2 anion
in ionic bonding it
receives 2 electrons
chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron
chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron
chlorine
group 17
7 valence electrons
forms -1 anion
in ionic bonding it
receives 1 electron

Lesson 62: Molecular Formula Writing Notes

Molecular Compounds (_______________ bonding)

Formula represents the number of ________________ of each _____________ in a single _________________.

Prefix System of Nomenclature

Prefix	Number

Prefix	Number

examples: dinitrogen monoxide _____________

tetraphosphorus decoxide ______________

carbon dioxide _____________; carbon monoxide _______________

Organic compounds

any __________________ bonded compound containing _______________, with the exception of _______________ and ___________________.

Lesson 62: Molecular Formula Writing (cont.) Chemistry with Lab

Hydrocarbons Composed of only _______________ and _________________.

- Saturated contain only ________________ bonds (___________________)
- Unsaturated contain __________________ (____________________) and

  ________________ (____________________) carbon-carbon bonds

  Very ______________ because of _____________ and _________________

Organic Nomenclature

Prefix	**# of carbon atoms (n)**

Prefix	**# of carbon atoms (n)**

formulas: Alkane = C_nH_{2n+2} Alkenes = C_nH_{2n}

examples: propane ______________ nonane ______________

butene ______________

<u>The Chemistry Quiz</u>

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 63: Molecular Compounds

1. In molecular compounds, __________________ are used to show the number of atoms of each element per molecule.

2. Organic compounds are generally defined as compounds that contain the element __________________.

3. A special type of organic compounds is _________________ which are composed of only hydrogen and carbon atoms.

4. C_nH_{2n+2} is the general form for the specific group of organic compounds called ____________________.

5. Molecular compounds are formed between atoms of

   a. nonmetals b. metals c. metals and nonmetals

6. Compounds containing only single bonds are said to be

   a. saturated b. unsaturated

Write the formula of the following compounds.

silicon dioxide _____________ carbon tetrachloride _____________

methane __________________ octane ___________________

dinitrogen trisulfide _____________ butene __________________

tetraphosphorus triselenide _________________

Lesson 63: Molecular Compound Assessment

Molecular compounds are formed between:

a. metals   b. nonmetals   c. metals and nonmetals

In molecular compounds, the atoms bond together to form:

a. molecules   b. ions

In molecular compounds, what are used to show the number of atoms of each element per molecule?

____

Match the prefix with the meaning:

a. di	_____	one
b. octa	_____	two
c. deca	_____	three
d. tri	_____	four
e. nona	_____	five
f. hexa	_____	six
g. mono	_____	seven
h. hepta	_____	eight
i. penta	_____	nine
j. tetra	_____	ten

Fill in the blanks using the word bank.

alkane	-ane	alkenes	bond	carbon	carbon atoms	double
-ene	hydrocarbons	prefix	saturated	triple	unsaturated	

Organic compounds are now defined as compounds that contain the element _______________. The nature of the ______________ between each pair of carbon atoms in an organic compound will determine whether the compound is saturated or unsaturated. The bonds between the carbon atoms in a(n) _______________ compound are single bonds, but in a(n) _______________ compound, the bonds between neighboring carbon atoms are double or _______________ bonds. The organic compounds containing only hydrogen and carbon are called ________________. C_nH_{2n+2} is the general form for the _______________ series of hydrocarbons. The names of this series are composed of a _________________, which denotes the number of carbon atoms present, and the suffix __________. This series of hydrocarbons has only single bonds, and so, is said to be ___________. C_nH_{2n+2} is the general form for the class of hydrocarbons referred to as the _____________. Each member of this series has a pair of carbon atoms connected by a ___________bond, and so, is said to be unsaturated. Again, prefixes are used to denote the number of ____________________ present in the molecule, and all members of this series end in the suffix ___________.

Lesson 65: Naming Compounds Notes

Binary Ionic Compounds

- Consist of only __________________ elements.
- Name the _________________ ion (the __________________).
- Name the _________________ ion (the __________________), changing the ending to ___________________.
- When metals that can form _________________ than one type of _____________________ are in a compound, use a __________________ _____________________ in parentheses after the name of the _____________________ to show the ____________________.

Examples: $NaCl$ ________________ MgO ________________

Cu_2S ________________ $SnCl_4$ ________________

Ternary Ionic Compounds

- Made up of ________________ elements.
- Name the _________________ then name the _________________ _____________________ without changing the ending to "ide."

Examples: Na_2SO_4 __________________________

$FeCrO_4$ __________________________

Naming Molecular Compounds

- The elements are named in the ___________________ they appear in the _____________________.

- __________________ are used to denote the ____________ of atoms of each ________________ in the molecule. An exception is that the __________ element named is given a _________________ only if there is more than _________ atom of that element in the ________________.
- The "o" or "a" at the ______________ of a ____________________ is ____________________ when the word following the ________________ begins with a ______________.
- The ____________ element's ending is changed to _____________.

Examples: ICl_3 ____________________

As_2O_5 ____________________

Naming Hydrocarbons

Alkane: C_2H_{2n+2} Alkene: C_nH_{2n}

Examples: C_4H_{10} ____________________

C_2H_4 ____________________

<u>The Chemistry Quiz</u>

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 65: Naming Compounds

Name the following compounds.

SO_3	______________________	CO	______________________
$LiCl$	______________________	Al_2O_3	______________________
MgS	______________________	CH_4	______________________
$SnCl_4$	______________________	As_2O_5	______________________
C_3H_6	______________________	PbO_2	______________________

Lesson 68: Ionic/Covalent Compounds

Determine whether each compound is ionic or covalent. Name the compound.

1. $CoBr_2$ ______________________
2. P_2O_5 ______________________
3. $FeSO_4$ ______________________
4. Na_2CO_3 ______________________
5. P_4 ______________________
6. B_2H_4 ______________________
7. CO ______________________
8. GaCl3 ______________________
9. SiO2 ______________________
10. NH_3 ______________________

Determine whether each compound is ionic or covalent. Write the formula.

11. dinitrogen trioxide ______________________
12. lithium acetate ______________________
13. phosphorous trifluoride ______________________
14. silicon tetrafluoride ______________________
15. vanadium (V) oxide ______________________
16. silver phosphate ______________________
17. aluminum hydroxide ______________________
18. nitrogen ______________________
19. methane ______________________
20. zinc sulfide ______________________

Lesson 69: Word Equations/Balancing Notes Chemistry with Lab

Chemical Equations

- a ____________________ way of reporting the _________________ of a

  _____________________ _________________________.

Reactants

- the __________________ substances in a _________________
- placed on the ___________________ side of the _____________________

Products

- the substances __________________ during a __________________

  ____
- placed on the ___________________ side of the ________________

➔

- arrow is read as _________________

Coefficients

- used to ___________________ ____________________
- represent the ___________________ of __________________,

  __________________, or __________________ of the _________________

Law of Conservation of Mass

Mass ________________ be __________________ or __________________.

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$

$$H_2 + O_2 \rightarrow H_2O$$

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

Guidelines for Balancing Equations

- Start by ______________ an _______________ _________________
- Draw __________ around each _____________ in the ______________
- Begin with _______ molecule or formula unit of the _______________ containing the ____________ atoms.
- Balance _________________ ions that appear on ____________ sides of the ________________ as a single unit.

  Balance ______________ and ________________ atoms ___________.

$$KNO_3 \rightarrow KNO_2 + O_2$$

Hydrogen and oxygen combine to produce water.

7 Diatomic Elemental Molecules

- Diatomic: ___________ ______________
- _______________, ______________, ______________, ______________,

  _______________, ______________, ______________

Carbon disulfide reacts with oxygen to produce carbon dioxide and sulfur dioxide.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Descriptive Abbreviations

(__________) ____________________

(__________) ____________________

(__________) ____________________

(__________) ____________________

Determining State at Room Temperature

- All __________________ are solids except for _________________ which is a _________________.
- Most ________________ are _________________ with these exceptions:
  - liquid - ______________
  - solids - ______________, ______________, ______________, ________________, ______________, and ________________
- All __________________ are solids.
- ________________ compounds are ___________, unless stated otherwise.
- ________________ compounds are ___________, unless stated otherwise.

When heated, solid mercury (II) oxide yields mercury and oxygen gas.

Classifying Reactions

1. Synthesis:

- _____________________ or __________________ substances combine to form a more ______________ substance

- ____________________ → ________________
- Fe (s)+___________________ → ________________
- H_2O (l)+__________________ → ________________

2. Decomposition:

- A ___________________ substance is ______________ ______________ into ____________ or more _____________ substances.
- AB → ___________________
- $2H_2O$ (l) → ___________________ + _________________
- _______________ → _______________ + _________________

3. Single Replacement:

- A free ____________________ replaces a ____________ _____________ element in a ______________.
- A + BY → _______________ + __________________
- Zn (s)+ _____________ → _______________ + _________________
- 2Al(s)+ _____________ → _______________ + _________________
- Cu (s)+ $MgCl_2$ (aq) → ________________
    - Activity series: an ___________________ of elements in the order of their ___________________ to _________________

Activity Series of Metals

Decreasing Activity ↓

Lithium
Potassium
Barium
Calcium
Sodium
Magnesium
Aluminum
Manganese
Zinc
Chromium
Iron
Cadmium
Nickel
Tin
Lead
(Hydrogen)
Copper
Mercury
Silver
Gold

4. Double Replacement:

- The ________________ of reacting ________________ ______________ each other.
- Normally takes place in an ________________ ________________
- Also called ________________ reactions
- AX + BY $\rightarrow$ ________________ + ________________

- 2KI (*aq*)+ ________________ $\rightarrow$ _____________ + _______________
  - Precipitate
    - an _________________ ___________________
    - may be ________________ in an equation by ________________ or _______________
- NaCl (*aq*)+ _______________ $\rightarrow$ _____________ + _______________ _________________, or _________________ of the ________________

5. Combustion:

- Involves the ________________ of a substance with _______________
- Often called ___________________.
- The ________________ ___________________ of a ________________ produces ____________ and _______________.
- $C_xH_y + O_2 \rightarrow$ __________________ + __________________
- C_3H_8 (*g*)+$5O_2$ (*g*) $\rightarrow$ ________________ + __________________
- CH_4 (*g*)+ ____________ $\rightarrow$ _______________ + __________________

The Chemistry Quiz

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 77: Mass Relationships Key Terms

1. **Avogadro's number** – The number of molecules in a mole of a substance, approximately 6.0225×10^{23}.

2. **Mole** – The amount of a substance that contains as many atoms, molecules, ions, or other elementary units as the number of atoms in 0.012 kilograms of carbon 12. The number is 6.0225×10^{23}, or Avogadro's number.

3. **Molar mass** – symbol M, is a physical property characteristic of a given substance, namely its mass per amount of substance. The base SI unit for mass is the kilogram and that for the amount of substance is the mole, thus the derived unit for molar mass is kg/mol. However, for both practical and historical reasons, molar masses are almost always quoted in grams per mole (g/mol).

4. **Empirical formula** – A chemical formula that indicates the relative proportions of the elements in a molecule rather than the actual number of atoms of the elements.

5. **Molecular formula** – A chemical formula that shows the number and kinds of atoms in a molecule.

6. **Hydrate** – A solid compound containing water molecules combined in a definite ratio as an integral part of the crystal.

7. **Percentage composition** – The mass percent of each element in a compound.

Lab results: 1 dozen grains of rice = _______ g (Use this fact as a conversion factor.)

? grains of rice = 1.94 g

Avogadro's number – the _____________ = the number _________________

Molar Mass

the ____________ of one _________ of any __________ ___________________

molar mass of an element (g/mol)

equals the ________________ ______________ of the element in grams

molar mass of a compound

equals the ________ of the _________ ____________ of the ______________

making up the _________________.

Example problems

Find the molar masses of:

magnesium __________ g/mol chlorine ________ g/mol

magnesium chloride _____________ g/mol

Find the molar mass of aluminum oxide.

How many moles of atoms are there in 8.0 g of Mg?

How many atoms are in 0.7 g of He?

How many atoms are in 3.5 grams of silicon?

How many formula units are in 32.6 grams of potassium oxide?

How many molecules are in 0.25 grams of dinitrogen pentoxide?

more example problems:

What would be the mass of 5.3×10^{22} formula units of calcium iodide?

<u>Episode 701 Problems:</u> *Fill in masses from lab data and solve.*

How many atoms are in _________ grams of copper?

How many formula units are in ________ grams of salt?

How many molecules are in ______ grams of water?

<u>The Chemistry Quiz</u>

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 79: Mole Problems

What is the mass in grams of 0.250 moles of Mg?

How many moles of atoms are there in 48.096 g of sulfur?

What is the mass of 24.6 formula units of magnesium oxide?

What is the mass of 4.56 moles of sulfur dioxide?

What is the mass of 4.56 molecules of sulfur dioxide?

How many formula units are in 0.456 g of $Ca_3(PO_4)_2$?

How many moles are in 1.5 x 10^{23} atoms of fluorine?

What is the mass of 2.80 x 10^{26} molecules of CO_2?

Lesson 79: More Mole Problems

Chemistry with Lab

How many atoms are contained in 3.46 moles of magnesium?

Convert 256.3 g sodium carbonate to formula units.

What is the mass of 12.4 molecules of carbon tetrachloride?

How many moles are contained in 0.43 g Al_2O_3?

The number of silicon atoms in 42.1 g of silicon is _____________.

11.6 g of $CuCl_2$ = ______________ moles $CuCl_2$.

Convert 2.76 g Se to moles Se.

Lesson 79: More Mole Problems (cont.)

What is the mass in grams of 9.31×10^{21} atoms of carbon?

2.86 f.u. MgO = _______________ moles MgO.

How many grams of sodium is equal to 6.92×10^{21} atoms of sodium?

The mass of 3.55 moles of NaCl is _______________.

BONUS

What is the mass, in kilograms, of a 2.46×10^{20} formula units of barium chloride?

Lesson 80: Even More Mole Problems

How many grams are contained in 0.54 moles of calcium?

Convert 2.54 moles magnesium carbonate to formula units.

What is the mass of 3.1×10^{20} molecules of carbon monoxide?

How many formula units are contained in 0.67 g CaO?

The number of nitrogen molecules in 1.2 moles of N_2 is _____________.

1.23×10^{28} f.u. of Al_2O_3 = ______________ grams of Al_2O_3.

Convert 3.1 moles of water to grams of water.

Lesson 80: Even More Mole Problems (cont.)

What is the mass in grams of 4.6×10^{21} molecules of carbon tetrabromide?

3.2×10^{23} f.u. NaCl = _______________ grams NaCl.

How many moles of silver is equal to 0.31 grams of silver?

The mass of 2.60 moles of SO_2 is _______________.

BONUS

How many atoms of oxygen are contained in 12.3 kilograms of carbon dioxide?

While you watch the introduction video, note below your preliminary thoughts about which person/people you personally suspect, and why and how they might have done it.

While you interview suspects and gather evidence, make notes below about each suspect – do you personally suspect them? What reasons might they have had to commit the crime, and most importantly, how might they have done it? What were their means?

 Dr. Yervin	Do you suspect him? What is his motive? How might he have done it?
 Sam	Do you suspect him? What is his motive? How might he have done it?
 Joanna	Do you suspect her? What is her motive? How might she have done it?
 Vince	Do you suspect him? What is his motive? How might he have done it?
 Nelson	Do you suspect him? What is his motive? How might he have done it?

Lesson 81: Mixed Reception (cont.)

While you gather evidence, make notes below about how each piece of evidence might relate to a suspect, how it might have been used in the crime, etc.

 Drug Info Sheet	
 Joanna's Emails	
 Anti-Toxin Report	
 Punch	
 Coroner's Report	
 Nelson's Journal	
 Peanuts	
 Peanut Pie	
 Nelson's Pills	
 Pills from Joanna's Office	

Use this page for additional notes, calculations, and conclusions you make as you analyze the evidence and form your conclusion.

Lesson 81: Case Report Form

An unknown substance was present in Nelson's blood. What was this substance and how did it get into his blood? What evidence do you have to support this?

The molecular weight of the unknown substance in Nelson's blood is 765.82. Why?

What is the MW of the allergy drug?

What is the MW of the anti-venom?

The coroner says that Nelson died from peanut allergy, yet he was taking a drug for this. Can you explain why he would still die from peanuts?

Did Nelson take his medication that day? Was the correct concentration of medication present in his blood?

You found an abstract on Sam’s desk. Is this evidence relevant to your solution? Why or why not?

How did Sam know that giving Nelson the antivenom would kill him?

There were pills found in Joanna’s office. Is this evidence relevant to your solution? Why or why not?

Are Joanna’s emails important evidence in support of your solution?

Was any of the food left at the crime scene important evidence for your case? Why or why not?

Who did it?

Why did they do it?

How did they do it?

Lesson 84: Percent Composition Notes

Example Problem: Find the % by mass of oxygen in water.

Percentage by mass of element in a compound =
(mass of element in 1 mol of compound ÷ molar mass of compound) x 100%

(*after completing lab #1*)
Find the % of carbon in sodium bicarbonate ($NaHCO_3$).

Find the % composition of aluminum oxide. (This means to find the % of each element in the compound.)

Empirical Formula: simplest ___________ number ____________ of

_______________ in a ________________.

Example problem: Find the empirical formula for a compound containing 56.6 g of K, 8.7 g of C, and 34.7 g of O.

Step #1: Convert each mass into moles of the element.

Step #2: Divide each by the smallest to find a simple whole number ratio.

Ex. Problems: Work on separate paper.

__________ % Na

__________ % S

__________ % O

(Hint: When % are given, assume you have 100g of the compound, and the % changes to grams.)

P_xO_y

_____ g sample

_____ g P

(Hint: After step 2, if the ratio is still not whole numbers, multiply both subscripts by a number, such as 2, to get rid of fractions, such as 0.5.)

<u>The Chemistry Quiz</u>

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 84: Lab Sheet 1

I. Fill in the data table below as you watch the lab on the video.

mass of sodium bicarbonate	
mass before reaction	
mass after reaction	

II. Write a balanced equation for the reaction that took place. (Hint: the narrator on the video will help you with this.)

III. Conclusion Questions: answer each question completely. Show all work!

- Calculate the mass of carbon dioxide produced in the experiment. (Hint: think about what bubbled away.)

- Use molar masses to calculate the percent of carbon in carbon dioxide using the following formula.

$$\% \text{ C in } CO_2 = \frac{\text{mass of C}}{\text{mass of } CO_2} \times 100\%$$

- Calculate the mass of carbon in the sample of carbon dioxide using the same formula and your answers to the previous two calculations.

- Use molar masses to calculate the percent of carbon in carbon dioxide using the following formula.

$$\% \text{ C in } NaHCO_3 = \frac{\text{mass of C}}{\text{mass of } NaHCO_3} \times 100\%$$

IV. Practice Problems. Show all work!

- Calculate the percentage sodium in sodium oxide.

- Calculate the percentage aluminum in aluminum phosphate.

- Calculate the percentage hydrogen in hydrogen peroxide.

- Calculate the percentage nitrogen in dinitrogen pentoxide.

Lesson 84: Lab Sheet 2

I. Fill in the data table below as you watch the lab on the video.

mass of empty test tube	
mass of test tube with copper oxide (before reaction)	
mass of test tube with copper (after reaction)	

II. Calculations: show all work.

- Use the data to calculate the mass of copper oxide.
- Use the data to calculate the mass of copper.
- Use the two previous calculations to calculate the mass of the oxygen that was driven off.
- Using molar masses, calculate the moles of copper and oxygen.

mass of copper oxide		g
mass of copper		g
mass of oxygen		g
moles of copper		mol
moles of oxygen		mol

Lesson 84: Lab Sheet 2 (cont.)

III. Conclusion questions: answer each question completely. Show all work!

1. Within bounds of experimental error, use the mole ratio to write the formula of the compound.

2. What is the name of this compound? (Hint: you need a Roman numeral in the name.)

3. A compound of calcium and bromine is analyzed in the lab. A 20.0 g sample contains 4.00 g calcium. What is the empirical formula of the compound?

4. Find the empirical formula of a compound found to contain 26.56% potassium, 35.41% chromium, and the remainder oxygen.

Purpose: To determine the percent of sugar in soda, produce, or other grocery product.

Procedure: Choose 3 drinks, 3 fruits or vegetables, and 4 other food products of your choice. Record the 10 items in the data table below. Collect the nutrition data on the items. Record the mass of 1 serving size of each item in the data table below. Record the mass of sugar in 1 serving size of each item in the data table below.

Data:

Grocery Item	Mass (g) of 1 serving	Mass of sugar (g) in 1 serving	% Sugar

Calculations:

1. Calculate the percent composition of sugar in each substance. Use this formula: % sugar = (mass of sugar / mass of 1 serving) x 100%. Record the value in the data table, showing your work.

2. Create a bar graph and compare the percent of sugar in each substance. Include your graph as part of your lab report.

Analysis:

1. What has the highest percentage of sugar?

2. Does any category consistently have the most sugar?

3. Do you notice any correlation between % sugar and natural foods vs. processed foods?

4. Do you think more data is necessary when making health-conscious choices, or is it enough to look at % sugar alone?

Lesson 88: Molecular Formulas Notes

Molecular Formula

- represents the ____________ number of _______________ of each _______________ in the ______________
- not necessary for _______________ ______________
- necessary for _____________ ______________

The molecular formula for water is _______________, and empirical formula is also ________________.

The molecular formula for hydrogen peroxide is ______________, and empirical formula is _____________.

Example Problem
The empirical formula for glucose is CH_2O.

a) If the molar mass is 180.0 g/mole, find the molecular formula.

b) If the molar mass is ________ g/mole, find the molecular formula.

Problem Set One (work on separate paper)

empirical formula	molar mass
CH	____ g/mol
NO_2	____ g/mol
C_3H_8	____ g/mol

Ex. Problem: Find the molecular formula for a compound with:

__________ g N __________ g O molar mass __________ g/mol

Hydrates

- _______________ with _______________ molecules adhering to the _______________ or ____________
- $Na_2CO_3 \bullet$ _______ H_2O

  indicates _______ ______________ molecules adhering to each _______________ _________ of sodium carbonate
- mass of _____________ = mass of _______________ compound minus mass of __________________ compound

Example Problems

Determine the formula of hydrated barium chloride from this data:

initial mass of hydrated compound = 1.373 g
mass after heating – 1.175 g

Determine the formula for the hydrate that is ___________% $CaSO_3$ and

__________% H_2O.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Mass of empty pipet	
Mass of hydrate and pipet	
Mass after 1st heating	
Mass after 2nd heating	

Calculations:

1. Using the data, calculate the mass of the water driven off.

2. Using your answer to number 1, calculate the number of moles of water driven off.

3. Using the data, calculate the mass of the “anhydrous” salt.

4. Using your answer to number 3 and the fact that your "anhydrous" salt is copper (II) sulfate, calculate the number of moles of "anhydrous" salt.

5. Using your answers to numbers 2 and 4, determine the formula of the hydrate.

Conclusions:

1. A hydrate is a crystal with ____________ molecules adhering to the ions or molecules.

2. Define "anhydrous."

3. Write the formula for hydrated zinc sulfate with 7 molecules of water adhering to each formula unit of zinc sulfate.

Lesson 91: Molecular Formulas and Hydrates Chemistry with Lab

1. The empirical formula of a compound is found to be P_2O_5. Experiments show that the molar mass of the compound is 283.9 g/mol. What is the molecular formula of the compound?

2. A compound has the following % composition—76.54 % C, 12.13 % H, and 11.33 % O. If its molar mass is 282.5 g/mol, what is its molecular formula?

3. What is the formula for a hydrate which consists of 90.7 % SrC2O4 and 9.30 % H2O?

Lesson 93: Stoichiometry Key Terms

1. **Stoichiometry** – Calculation of the quantities of reactants and products in a chemical reaction.
2. **Mole ratio** – Tells you the relative amounts of reactants and products in a chemical reaction.
3. **Actual yield** – Mass of product obtained from a chemical reaction.
4. **Theoretical yield** – Amount of product stoichiometrically predicted to be present when a limiting reactant is used up.
5. **Percent yield** – Actual yield of a chemical reaction, expressed as a percentage of the theoretical yield.
6. **Limiting reactant** – The reactant that is used up completely in a reaction, limiting further chemical change.
7. **Excess reactant** – The reactant that is left over in a chemical reaction.

Lesson 93: Mole-Mole and Mole-Mass Notes Chemistry with Lab

Stoichiometry

- study of the _______________ relationships in a ____________________ _____________________
- based on __________________ equations

$2\ Mg + O_2 \longrightarrow 2\ MgO$

The__________________ in a ________________ ______________ give the _____________________ ____________________ for the ___________________ involved in the __________________.

Ex. Problem:

When elemental aluminum reacts with elemental iodine, aluminum iodide is produced.

mole ratios: _____Al: _______I_2

_____Al: _______AlI_3

_____I_2: _______AlI_3

If you start with 4 moles of Al, how many moles of AlI_3 will be produced?

Problem Set One

How many moles of water will be produced when ________ grams of hydrogen gas react with the oxygen in the air?

(Hint: to “make the switch” between different substances in a reaction, us the

__________ ratio from the _________________ equation.)

Problem Set Two

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Can You Find the Mole Ratio?

In this experiment, baking soda ($NaHCO_3$) reacts with hydrochloric acid (HCl) to produce sodium chloride, carbon dioxide, and water. You will determine the moles of reactant used and moles of product produced using both the lab data and stoichiometry and compare the two.

Data table:

mass of test tube and baking soda	
mass of empty test tube	
mass of baking soda	
mass of test tube and sodium chloride	
mass of empty test tube	
mass of sodium chloride	

Conclusion Questions:

1. Calculate the number of moles of baking soda used in the lab.

2. Calculate the number of moles of sodium chloride produced in the lab.

3. What is the experimental mole ratio of baking soda ($NaHCO_3$) to sodium chloride (HCl)?

4. Write a balanced equation for the reaction that took place in the experiment.

5. According to the balanced equation, what is the theoretical mole ratio of baking soda to sodium chloride?

6. Was the experimental mole ratio exactly the same as the theoretical mole ratio? If not, give some possible reasons.

Lesson 96: Stoichiometry Notes

Ex. Problems:

Sodium metal reacts with oxygen to produce solid sodium oxide. How many grams of sodium must react to produce ______ g of sodium oxide?

When _____ g of hydrogen reacts with oxygen, how many grams of water are produced?

Actual yield: amount of _____________ produced when the _____________ is performed in a ________________

Theoretical yield: amount of ____________ expected to be __________________ based on the _______________ _______________ and the amount of _________________

Percent yield: (_______________ yield / _______________ yield) x 100%

Percent Yield in Lab

actual yield of CO_2 = ________________ (from lab data)

Calculate theoretical yield from balanced equation:

$NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$

?g CO_2 = 0.23g $NaHCO_3$

% yield =

What is the % yield of carbon dioxide when ____________ grams of propane are burned and ___________ grams of carbon dioxide are collected?

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 102: Limiting Reactants Notes

Stoichiometry Problem Guidelines:

1. _______________ the equation
2. convert to ________________ of _________________ substance
3. MAKE THE ________________ USING ______________ _______________

   FROM BALANCED EQUATION
4. convert to ________________

Problem Set One:

Limiting Reactant

_______________ used up ____________ in a ________________ reaction.

Excess Reactant

_______________ that is not used up on a chemical ________________.

Ex. Problem:

When $FeCl_3$ reacts with O_2, Fe_2O_3 and Cl_2 are produced. If 4.0 moles of $FeCl_3$ and 4.0 moles of O_2 are mixed, how many grams of Fe_2O_3 will be produced?

$\rightarrow$

(Hint: Work two separate problems, using one reactant at a time.)

(Hint: Answer will be the [smaller, larger] amount of product.)

What is the limiting reactant?

What is the excess reactant?

Problem Set Two:

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Complete the following using the candy provided. **Do not unwrap any candy.**

1. One butterscotch reacts with one peppermint to form a pepperscotch according to the following BALANCED equation. 1 Bs + 1 P • 1 PS

   a. Use the candy in the bag to illustrate this reaction and answer the following questions. How many pepperscotches can be formed and what is the limiting reactant? What reactant is in excess and how much of it is left over?

   Number of pepperscotches formed ________________

   Limiting Reactant _____________

   Excess Reactant _____________

   Amount of excess left over ______

   b. Use the balanced equation to answer the following question. One butterscotch has a mass of 5.0 grams and one peppermint has a mass of 4.0 grams. How many pepperscotches can be made with 50.0 grams of butterscotch and 48.0 grams of peppermints?

2. One tootsie roll reacts with four gummy bears to form a tootsie bear according to the following BALANCED equation. 1 Tr + 4 Gb • 1 TB

   a. How many tootsie bears can be formed and what is the limiting reactant? What is the excess reactant and how much is left over?

   Number of tootsie bears formed ________________

   Limiting Reactant _____________

Excess Reactant ______________

Amount of excess left over ______

b. Use the balanced equation to answer the following question. One tootsie roll has a mass of 2.0 grams and one gummy bear has a mass of 1.5 g. How many tootsie bears can be made with 12.5 grams of tootsie rolls and 15.0 grams of gummy bears?

3. Two starburst fruit react with six skittles to from a skitburst according to the following BALANCED equation. 2 Sb + 6 Sk • 1 SB

a. How many skitbursts can be formed and what is the limiting reactant? What is the excess reactant and how much is left over?

Number of skitbursts formed _________________

Limiting Reactant ______________

Excess Reactant ______________

Amount of excess left over ______

b. Use the balanced equation to answer the following question. One starburst has a mass of 5.0 grams and one skittle has a mass of 1.0 gram. How many skitburts can be made from 40.0 grams of starburst and 26.0 grams of skittles?

4. One caramel reacts with three candy corn to form a caramel corn according to the following BALANCED equation. 1 C + 3 Cc • 1 CC

   a. How many caramel corns can be formed and what is the limiting reactant? What is the excess reactant and how much is left over?

      Number of caramel corns formed ________________

      Limiting Reactant _____________

      Excess Reactant _____________

      Amount of excess left over ______

   b. Use the balanced equation to answer the following question. One candy corn has a mass of 1.5 grams and one caramel has a mass of 11.0 g. How many caramel corns can be made with 60.0 grams of candy corn and 66.0 grams of caramel?

$$P_1V_1 = P_2V_2 \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$PV = nRT \qquad P_T = P_1 + P_2 + P_3 \ldots$$

$$\frac{v_1}{v_2} = \sqrt{\frac{d_2}{d_1}} \qquad \frac{v_1}{v_2} = \sqrt{\frac{m_2}{m_1}}$$

$$R = 8.314 \frac{L \cdot kPa}{mol \cdot K} \qquad R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$$

Water-Vapor Pressure

Temp (°C)	Pressure (mm Hg)	Pressure (kPa)
0.0	4.6	0.61
5.0	6.2	0.87
10.0	9.2	1.23
15.0	12.8	1.71
16.0	13.6	1.82
17.0	14.5	1.94
18.0	15.5	2.06
19.0	16.5	2.19
20.0	17.5	2.34
21.0	18.5	2.49
22.0	19.8	2.64
23.0	21.1	2.81

Temp (°C)	Pressure (mm Hg)	Pressure (kPa)
24.0	22.4	2.98
25.0	23.8	3.17
26.0	25.2	3.36
27.0	26.7	3.57
28.0	28.3	3.78
29.0	30.0	4.01
30.0	31.8	4.25
35.0	42.2	5.63
40.0	55.3	7.39
50.0	92.5	12.23
60.0	149.4	19.93
70.0	233.7	31.18

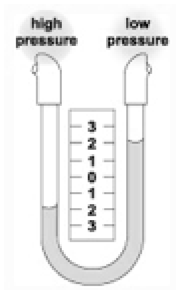
Kinetic Theory

1. **Kinetic theory** – physical theory that explains the behavior of gases on the basis of the following assumptions: (1) Any gas is composed of a very large number of very tiny particles called molecules; (2) The molecules are very far apart compared to their sizes, so that they can be considered as points; (3) The molecules exert no forces on one another except during rare collisions, and these collisions are perfectly elastic, i.e., they take place within a negligible span of time and in accordance with the laws of mechanics.
2. **Ideal gas** – A gas that, when kept at a constant temperature, would obey the gas laws exactly. No known gas is an ideal gas.

Pressure

3. **Pressure** – Force applied uniformly over a surface, measured as force per unit of area.
4. **Atmosphere** – A unit of pressure equal to the air pressure at sea level. It equals the amount of pressure that will support a column of mercury 760 millimeters high at 0 degrees Celsius under standard gravity.
5. **Kilopascal** – One thousand (10^3) pascals; the metric unit of pressure; one pound per square inch is equal to 6.89 kilopascals; abbreviated kPa.
6. **mm Hg**– A unit of pressure that is equal to approximately 1.316×10^{-3} atmospheres or 133.3 pascals.
7. **Atmospheric pressure** – Pressure caused by the weight of the atmosphere. At sea level it has a mean value of one atmosphere but reduces with increasing altitude.
8. **Standard temperature** – A temperature of 0°C or 273 Kelvin.
9. **Standard pressure** – A pressure of 1 atmosphere (101,325 newtons per square meter), to which measurements of quantities dependent on pressure, such as the volume of a gas, are often referred.
10. **STP** – standard temperature and pressure
11. **Barometer** – An absolute pressure gage specifically designed to measure atmospheric pressure. This instrument is a type of manometer with one leg at zero pressure absolute.

12. **Manometer** – A double-leg liquid-column gage used to measure the difference between two fluid pressures.



Gas Laws

13. **Boyle's Law** – The principle that at a constant temperature the volume of a confined ideal gas varies inversely with its pressure. $P_1V_1 = P_2V_2$

14. **Charles' Law** – The physical law that the volume of a fixed mass of gas held at a constant pressure varies directly with the absolute temperature.
$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad \frac{V_2}{V_1} = \frac{T_2}{T_1} \quad \text{or} \quad V_1T_2 = V_2T_1$$

15. **Combined Gas Law** – A gas law which combines Charles' law, Boyle's law, and Gay-Lussac's law.
$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

16. **Avogadro's Law** – The principle that equal volumes of all gases under identical conditions of pressure and temperature contain the same number of molecules.

17. **Dalton's Law of Partial Pressure** - The law that the pressure of a gas mixture is equal to the sum of the partial pressures of the gases composing it.
   P total = P1+P2+P3+…

18. **Ideal Gas Law** - The equation of state of an ideal gas which is a good approximation to real gases at sufficiently high temperatures and low pressures; that is, $PV = nRT$, where P is the pressure, V is the volume, n is the number of moles, T is the temperature, and R is the gas constant.

Lesson 108: Kinetic Theory Notes

Kinetic Theory

- Gases are composed of _______________, ____________________ particles called __________________.
- Gas molecules are in __________________ ____________________.
- All __________________ between particles are _________________ ____________________.
- The _________________ of a gas display no ___________________ or __________________________ for one another.
- The _________________ ____________________ ____________________ of the molecules is ______________ ____________________ to the _____________________ temperature of the gas.

Ideal Gas

- Gas whose ___________________ conforms to the _________________ __________________ - it is _____________________.

Gas Pressure:

Pressure =

Atmospheric Pressure – the _______________ the earth's __________________ exerts due to its ________________.

Barometer:

- Instrument used to measure _______________ __________________.
- Invented by __________________

Normal Atmospheric Pressure

- Also called ________________ ____________________
- _____________________
- _____________________
- _____________________
- _____________________

STP:

- ________________ _______________ and _________________
- ________________
- ________________

Manometer:

- ________________ used to measure ________________ _______________
- U-shaped tube __________________ filled with _________________
- One end _____________ to __________________ _______________
- One end _____________ to __________________

<u>The Chemistry Quiz</u>

CR1. ______ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 110: Boyle's Law Notes

Chemistry with Lab

<u>Boyle's Law</u>

- The _________________ of a fixed ___________ of gas varies _____________________ with the __________________ at constant _____________________.
- _____ = ______
- __________ = ________

<u>Kinetic Theory and Boyle's Law</u>

❖ ________________ of a gas is caused by the ______________ of the gas ________________ the walls of the _________________.

❖ If the gas is ____________________ to __________ the volume it had, ________________ as many __________________ are present in any ________________ ____________________.

  ❖ _____________ as many __________________ per _________________ on the walls of the _________________

  ❖ __________________ of the gas will __________________

<u>Ex 1</u>: A balloon filled with Helium has a volume of 457 mL at standard atmospheric pressure. After the balloon is released, it reaches an altitude of 6.3 km where the pressure is only 65.5 kPa. What is the volume of the balloon at this altitude?

Ex 2: Under pressure of __________ mm Hg, a confined gas has a volume of

_____ mL. If the pressure is increased until the volume is ______ mL, what is the new pressure, assuming the temperature remains constant?

Charles's Law

- For a __________ ___________ of gas, as long as the _______________ is held _______________, the _________________ varies _______________ with the _____________ ____________________.
- _____ = ______
- __________ = ________

The Kelvin Temperature Scale

- _____________________ zero
  - ❖ ___________________ possible ___________________
  - ❖ ___________________ been reached
- _____ = absolute zero
- _____ = ______________
- K = ____________

Lesson 110: Boyle's Law Notes (cont.)

Ex 1: A quantity of gas occupies a volume of 506 cm^3 at a temperature of 147°C. Assuming the pressure stays constant, at what temperature will the volume of the gas be 604 cm^3?

Kinetic Molecular Theory and Charles's Law

- _______________ the ___________ ___________________ of a gas _______________ the average ______________ ______________ of its _______________.
- _______________ moving molecules
  - strike the walls of the ________________more __________
  - strike the walls of the ________________ with more __________
- From __________ law we derive that the _______________ would have to _____________ if the _______________ ___________________ is _______________ so that ______________ would remain _____________.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lab: Boyle's Law—Datasheet

Data and Calculations:

• Pressure is defined as force per unit area: P = Force / Area

• When a book rests on top of the plunger, the pressure it exerts equals the weight of the book divided by the area of the bottom of the plunger.

1. Diameter of syringe =______cm = _______m
   Show your work here.

2. Weight (force) of 1 book =________N
   Show your work here.

3. Calculate the area of the circle in m^2 using the formula πr^2:

4. Using the fact that pressure is defined as force ÷ area, what is the pressure exerted by 1 book in N/m^2?

5. If 1 N/m^2 = 1 Pascal (Pa), what is the pressure exerted by 1 book in Pa?

Using conversion facts, convert Pa to kPa. Show your work here.

Exerting Pressure	Pressure (kPa)	Volume of gas (cm^3)	P/V (kPa/cm^3)	P x V ($kPa \times cm^3$)
atmosphere				
atmosphere + 1 book				
atmosphere + 2 books				
atmosphere + 3 books				
atmosphere + 4 books				
atmosphere + 5 books				

Use graphical analysis to make a graph of "pressure vs. volume". Use proper technique for scientific graphing and be sure to draw a best-fit line.

Conclusions:
As the pressure exerted on the air inside the syringe increased, the volume (increased, decreased).

Volume and pressure are (directly, inversely) proportional.

Look at the last two columns of your data table. Within bounds of experimental error, the relationship between pressure and volume of a gas is (P/V = k , P•V = k). This relationship is known as: (Look at the title of this lab!)

____________________ __________________.

Lesson 112: Behavior of Gases Notes

The Combined Gas Law

- Expresses the relationship between the ______________, _____________, and _________________ of a ___________ amount of ___________.
- ________________ or _______________________

Ex: A sample of gas has a volume of ____L when its temperature is _____K and its pressure is _____mm Hg. What volume will the gas occupy at STP?

V_1 = ________________	V_2 = ________________
T_1 = ________________	T_2 = ________________
P_1 = ________________	P_2 = ________________

Diffusion

- The _________________ spreading of a ____________

Graham's Law of Diffusion

- Under the same conditions of _______________ and ________________, gases _______________ at a rate ____________ proportional to the ___________ __________ of their _____________ (or molar ___________).

- ________________ or ________________________

Ideal Gas Equation

- ______________________

- New variables:
  - n = ____________ of gas in ________________
  - R = _______________ __________ _______________
    - ❖ ____________________constant
    - ❖ value depends on ______ used for __________ and ____________
    - ❖ value of R when using ________________ and ____________:

      R = ___________________________

Ex: The average lung capacity for a female student is 3.9L. At normal body temperature, 37°C and 110kPa, how many moles of air could her lungs hold?

P = ____________ V = ____________ T = ____________

n = _______ R = ____________

Avogadro's Law

- Equal _______________ of different __________ under the ___________ conditions have the ________ number of _______________.
- Conversely, if samples of _______________ ________ at the same _______________ and ________________ contain the __________ number of ___________, then the ___________ of all the ____________ must be ________.
- At ________, one _______ of any gas occupies a ___________ of ______L
- ______ is the _____________ ______________ of a gas.

Ex. 3.2 moles of KNO_3 are heated, producing O_2 and KNO_2. Calculate the volume of O_2 in liters that could be obtained at STP.

Dalton's Law of Partial Pressures

- The _________________ of a gas __________________ is the __________ of the _______________ ______________ of each gas _______________.
- __________________________

Ex: Oxygen gas has been collected over water at a total pressure of 95.0 kPa and a temperature of 25°C. What is the pressure of the dry oxygen gas?

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Introduction

The diffusion rates (velocities) of HCl and NH_3 gases will be compared. Hydrogen chloride fumes will come from hydrochloric acid and ammonia fumes will come from aqueous ammonia. Both will be simultaneously introduced into opposite ends of a glass tube. When the gases meet, they will form a white precipitate, NH_4Cl, which will form a ring in the tube.

According to the ________________ theory, gas molecules are in constant motion, hitting each other and the sides of their container with perfectly

________________ collisions. The temperature of a gas is a measure of the

average __________________ energy of the molecules. The equation for calculating this energy is KE = ½mv^2.

If two gases are the same temperature, the molecules have the same average

kinetic energy. This makes KE a (constant, variable). This means that m and v^2

are _________________ proportional. Heavier molecules move (slower, faster)

than light molecules at the same temperature. Mathematically, the relationship can be stated as:

$$m_1 v_1^2 = m_2 v_2^2 \quad \textit{which equals} \quad \frac{v_1^2}{v_2^2} = \frac{m_2}{m_1} \quad \textit{which equals} \quad \boxed{\frac{v_1}{v_2} = \sqrt{\frac{m_2}{m_1}}}$$

The last equation is known as **Graham's Law of Diffusion**.

Procedure:

1. A drop of concentrated hydrochloric acid (a source of HCl fumes) was placed on a cotton swab. A drop of concentrated aqueous ammonia was placed on another cotton swab.

2. The swabs were simultaneously inserted into opposite ends of a glass tube.

3. The glass tube was left undisturbed for two minutes.

4. After two minutes, a white ring was located and the center of the ring was marked.

5. The distance from each end of the tube to the mark was measured.

   HCl: d_1= _______________ NH_3: d_2 = _________________

6. Calculate the ratio d_1/d_2 = _____________
   This is also the ratio of the velocities of the molecules, v_1/v_2.

   $$\boxed{\frac{v_1}{v_2} = \_\_\_\_\_\_}$$

7. Calculate the molar masses of the molecules:

   HCl: m_1= ____________ NH_3: m_2 = ________________

8. Calculate the ratio:

   $$\boxed{\sqrt{\frac{m_2}{m_1}} = \_\_\_\_\_\_}$$

9. Within bounds of experimental error, does $\frac{v_1}{v_2} = \sqrt{\frac{m_2}{m_1}}$? ______________

1. Which of these is a qualitative observation?

    a. The candle is 3.0 cm in length.

    b. The wick of the candle is curled at the top.

    c. The wax contains carbon and hydrogen.

2. Numbers expressed in scientific notation must be expressed in this form where: $Mx10^n$

    a. M is one and n is greater than or equal to one but less than ten.

    b. M is greater than or equal to one but less than ten, and n is one.

    c. n is greater than or equal to one but less than ten.

    d. M is greater than or equal to one but less than ten.

3. The process used to find the volume of an object by dropping it in water is called:

    a. water displacement   b. water subtracting   c. irregular volume

4. Which of the following could be used as a conversion factor?

    a. 1000 m = 1 mm   b. 1000 mm = 1 m

    c. 100 m = 1 cm   d. 1000 cm = 1 m

5. What do you do when you don't know a conversion factor between the two units in the problem?

    a. Make something up.   b. Give up and take a nap.

    c. Use multiple conversion factors.   d. Change the problem.

6. The instrument used to measure mass is the:

   a. graduated cylinder b. ruler

   c. balance d. stopwatch

7. Which of the following is NOT a physical property of the chair?

   a. The chair is blue b. The chair is made of wood.

   d. The chair will burn. d. The chair has four legs.

8. Which of the following is NOT a symbol for an element?

   a. Na b. CO c. S d. He

9. Which of the following is an example of a heterogeneous mixture?

   a. tea b. oxygen c. pizza d. water

10. A(n) ______ is made of two or more elements chemically combined.

    a. element b. mixture c. compound d. alloy

11. What is the formula for solving a density problem?

    a. D=mxv b. D=m/v c. D=v/m

12. Dalton's atomic theory stated atoms unite in small, whole number ratios to form compounds.

    True False

13. What particles are responsible for nearly all the mass of the atom?

    a. The protons and electrons b. The protons and neutrons

    c. The neutrons and electrons d. The electrons

14. The principal quantum number:

    a. is represented by n.

    b. refers to the main energy level of an electron.

    c. is a whole number.

    d. All of these

15. What is the electron distribution of carbon, which has an atomic number of six?

    a. $1s^1 2s^1 3s^1 3p^3$ b. $1s^2 2s^2 3s^2$ c. $1s^2 2s^2 2p^2$

16. The word "atom" comes from the Greek word "atomos" which means:

    a. divisible   b. indivisible   c. invisible

17. Mendeleev arranged the elements by increasing atomic:

    a. number   b. mass   c. theory

18. What phase of matter is oxygen at room temperature?

    a. solid   b. liquid   c. gas

19. How many energy levels does an atom of oxygen have?

    a. 1   b. 2   c. 3   d. 4

20. Which of the following elements has the largest atomic mass?

    a. K   b. Ca   c. Sc

Solutions

1. **Homogeneous** – Uniform in structure or composition throughout.
2. **Heterogeneous** – Consisting of dissimilar elements or parts; not homogeneous.
3. **Solution** – A homogeneous mixture of two or more substances, which may be solids, liquids, gases, or a combination of these.
4. **Mixture** – A composition of two or more substances that are not chemically combined with each other and are capable of being separated.
5. **Solute** – A substance dissolved in another substance, usually the component of a solution present in the lesser amount.
6. **Solvent**– A substance in which another substance is dissolved, forming a solution.
7. **Miscible** – Can be mixed in all proportions.
8. **Immiscible** – Cannot undergo mixing or blending.
9. **Alloy** – A homogeneous mixture or solid solution of two or more metals, the atoms of one replacing or occupying interstitial positions between the atoms of the other.
10. **Filtration** – The act or process of separating or (completely or partially) removing selected components of a mixture by means of a filter.
11. **Suspension** – A system in which microscopically visible particles are dispersed throughout a less dense liquid or gas from which they are easily filtered but not easily settled because of system viscosity or molecular interactions.
12. **Tyndall effect** – Visible scattering of light along the path of a beam of light as it passes through a system containing discontinuities, such as the surfaces of colloidal particles in a colloidal solution.
13. **Colloid** – A system in which finely divided particles, which are approximately 10 to 10,000 angstroms in size, are dispersed within a continuous medium in a manner that prevents them from being filtered easily or settled rapidly.
14. **Rate of Solution** – How quickly a solute dissolves in a solvent. Factors determining the rate of solution are: surface area, stirring, amount of solute already dissolved, and temperature.

Solubility

15. **Electrolytes** – A chemical compound that ionizes when dissolved or molten to produce an electrically conductive medium.
16. **Nonelectrolytes** – A substance whose molecules in solution do not dissociate to ions and thus do not conduct an electric current.
17. **Concentrated** – Having a high concentration of the solute.
18. **Dilute** – Describing a solution that has a relatively low concentration of solute. solubility - The amount of a substance that can be dissolved in a given amount of solvent.
19. **Aqueous** – Dissolved in water.
20. **Tincture** – A solution with alcohol as the solvent.
21. **Emulsion** – A suspension of small globules of one liquid in a second liquid with which the first will not mix: an emulsion of oil in vinegar.
22. **Saturated** – Combined with or containing all the solute that can normally be dissolved at a given temperature.
23. **Supersaturated** – To cause (a chemical solution) to be more highly concentrated than is normally possible under given conditions of temperature and pressure.

Molarity and Colligative Properties

24. **Molarity** – The molar concentration of a solution, usually expressed as the number of moles of solute per liter of solution.
25. **Molality** – The molal concentration of a solute, usually expressed as the number of moles of solute per kilograms of solvent.
26. **Colligative Property** – Properties dependent on the number of molecules but not their nature.

Lesson 120: A Special Type of Mixture Notes
Chemistry with Lab

Solutions

- Formed when substances ______________ in other _______________
- ___________________ mixtures
- ______________ phase
- remain __________________; particles do not ______________ out
- cannot be separated by ___________________
    - solvent:
        - present in _______________ amount
        - _______________ the ___________ to make the solution
    - solute:
        - present in _______________ amount
        - _______________ in the _____________

Examples of Types of Solutions

Liquid - ____________ solvent in which a ________, _____________, or ___________________ is _________________

- _____ dissolved in _____________: _______________ ____________
- __________ in ____________: __________________ in water
    - _______________: the two liquids mix
    - ___________________: the two liquids __________ mix
- ______ dissolved in _____________: _______ water

Lesson 120: Special Type of Mixture (cont.)

Solid

- __________: solid mixtures of _________: (_____________ is a mixture of _____________ and ___________)

Gas

- gases dissolved in __________ other: (______ is most common example)

Aqueous: ____________ is the ____________________

Tincture: ____________ is the ____________________

Suspension

- a ________________________ mixture
- particles in the ____________ are thousands of times _______________ than ________________ and ___________
- particles will ______________ out upon _________________
- can be separated by __________________
- exhibit the ________________ ____________________ - the ____________________ of ____________ in all directions

Colloid

- particles are ______________ in size between those of ______________ and true _____________
- particles do not ____________ out upon _________________

Lesson 120: Special Type of Mixture (cont.)

- can not be separated by ____________________
- exhibit the ________________ ______________

Emulsion

- ________________ dispersion of ________________ in ________________
- ___________________ agent is necessary for maintaining _____________

(____________ is an example)

Electrolyte: dissolves in water to form a ______________ that ______________

____

Nonelectrolyte: dissolves in water to form a _____________ that does ______

conduct _____________

Factors Affecting the Rate of Solution

1) ________________ _________:

increasing the surface area of the _____________ by ________________

speeds up _____________ by increasing the number of ______________

between the _______________ and the _____________ surface.

2) _____________________:

________________ or _____________ helps to disperse solute particles,

__________________ the number of ________________ between the

________________ and the _____________ surface.

3) ______________:

increases the average _______________ ___________ of the solvent molecules so that ______________ between the solvent molecules and the ___________ are more _______________.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 122: Solubility Notes

Saturated: solution containing all the ___________________ ______________ possible at given conditions of ___________________ and ________________.

Unsaturated: solution containing ________ dissolved _______________ than the _________________ amount that can be _________________ at given conditions of ___________________ and ________________.

Supersaturated: unusual solution containing _______ dissolved ___________ than is normally ________________ at given conditions of ________________ and _________________.

Solubility: a _______________ of how much __________ can ______________ in a given amount of ____________ at a specific ___________________.

Dilute Solution: The amount of ______________ dissolved is _________ in relation to the amount of ______________ present.

Concentrated Solution: The amount of ____________ dissolved is _________ in relation to the amount of ___________ present.

Factors Affecting Solubility

1) _______________ of ____________ and ________________: "________ dissolves ________."

2) ____________________

- Generally, increasing the ___________________ of the solution _______________ the solubility of a ______________ solute

3) _______________

- Only affects the ______________ of a _________ ____________
- ____________________ pressure _________________ solubility
- ____________________ pressure _________________ solubility
- Henry's Law: The __________ of a _______ dissolved in a given _______________ of ____________ is _____________ proportional to the _______________ of the _________.

The Chemistry Quiz

CR1. _____ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Procedure:

1) Each lab group was assigned a different temperature, calculating the solubility of KCl in water at that temperature. Each group conducted the lab as seen on the video, and their data follows. You are to complete columns #5 and #6, using the provided data from each lab group.

Data:

#1	#2	#3	#4	#5	#6	#7
Assigned Temperature °C	Mass of evaporating dish + cover	Mass of evaporating dish + cover + KCl solution	Mass of evaporating dish + cover + dry KCl	Mass of KCl (#4 - #2)	Mass of water (#3 - #4)	Mass of KCl per 100 g of water
10	37.81 g	48.75 g	40.54 g			
20	37.65 g	50.42 g	40.99 g			
30	36.95 g	47.82 g	39.81 g			
40	37.80 g	48.19 g	40.63 g			
50	36.50 g	47.49 g	39.75 g			
60	37.75 g	50.08 g	41.52 g			
70	35.98 g	48.29 g	39.86 g			
80	37.81 g	48.64 g	41.37 g			
90	36.99 g	48.96 g	40.98 g			

2) Column #7 is to be calculated as to express the solubility of KCl in grams per 100 grams of water. Following the equation given, complete column #7 for each group's data.

$$? \text{ g KCl} = 100 \text{ g } H_2O \times \frac{\text{Mass of KCl}}{\text{Mass of water}} =$$

3) Construct a graph using the vertical axis for grams of solute per 100 g of solvent and the horizontal axis for temperature. Staple your graph to this paper. BE SURE TO USE PROPER SCIENTIFIC GRAPHING TECHNIQUES.

<u>Conclusion Questions</u>:

1. The solubility of a solute is the maximum mass of the solute that will

   dissolve in a certain amount of water at a certain _________________.
   This is the same as saying that solubility is the concentration of a/n

   (unsaturated, saturated) solution of the solute.

2. From your graph, what mass of KCl can be dissolved in 100 g of water at these temperatures? (*Use dotted lines on your graph to show how you used your graph to determine your answers*):

   a) 25 °C ____________ b) 55 °C ____________

3. For each of the following, tell whether the solution would be saturated, unsaturated, or *crystalizing. (Hint: Plot the point and see whether it lies above, below, or on the best-fit line. Remember that the line represents

   a _____________ solution.)

   a) 40.0 g of KCl in 100 g of water at 75 °C _____________

   b) 34.0 g of KCl in 100 g of water at 55 °C _____________

   c) 45.0 g of KCl in 100 g of water at 25 °C _____________

   * Normally, crystallization (rather than supersaturation) occurs when more solute is present than what can be dissolved in a given amount of solvent at a given temperature. Crystallization simply refers to the excess solute “crystallizing” and settling out of the solution. Supersaturation is rare.

Lesson 125: Molarity Notes

<u>Molarity (M)</u>

- expresses ________________ ___________________
- M = ____________________
- ________________, ____________, ____, and ______/________ all represent the same ratio.

Ex #1. A saline solution contains ________g of NaCl per ________mL of solution. What is its molarity?

M = _______________

Ex #2 How many moles of solute are contained in ______L of _____M $CaCl_2$?

Colligative Properties

- Any of the ________________ of a ________________ that change when the ______________________ of the ______________ changes
- Depend on the _________________ of _________________ dissolved in a given __________ of _________________
- Examples of Colligative Properties:
    - Vapor _______________ Depression
        - The _________________ of the _________ _______________ of a liquid that occurs when substances are ______________ in the _______________
        - Vapor Pressure: the ________________ of a _____________ in _______________________ with its _____________
    - Freezing _____________ Depression
        - the _________________ of the ______________ __________ of a liquid that occurs when substances are ______________ in the _______________ (ex. using ________________ in car radiators and ________ on icy roads)
    - Boiling ______________ Elevation
        - the _________________ of the ______________ __________ of a liquid that occurs when substances are ______________ in the _______________

    - boiling occurs when __________ ________________ equals ______________ ______________ (ex. _______________ in a car acts as a coolant in the summer.)

- More Examples of Colligative Properties:

  ___________, ____________, _____________

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 130: Molarity Practice Problems

1. How many grams of potassium carbonate are needed to make 280 mL of a 2.5 M solution?

2. How many liters of water are needed to make a 4.00 M solution using 75.0 grams of lithium bromide?

3. What is the concentration if I have 450 mL of iron (II) chloride solution that contains 9.6 grams of iron (II) chloride solute?

4. How many grams of ammonium sulfate are needed to make a 0.25 L of solution at a concentration of 6.0 M?

5. If I have 2.5 L of a solution that contains 660 grams of calcium phosphate, what is this solution's concentration?

6. Extra Credit: How many grams of copper (II) fluoride are needed to make 6.7 liters of a 1.2 M solution?

Lesson 130: Molarity Practice Problems (cont.) Chemistry with Lab

1. How many liters of 0.88 M LiF solution can be made with 25.5 grams of solute?

2. What is the concentration of a solution that has a volume of 660 mL and contains 33.4 grams of aluminum acetate?

3. How many liters of 0.75 M solution can be made using 75 grams of lead (II) oxide?

4. How many kilograms of manganese (IV) oxide are needed to make 5.6 liters of a 2.1 M solution?

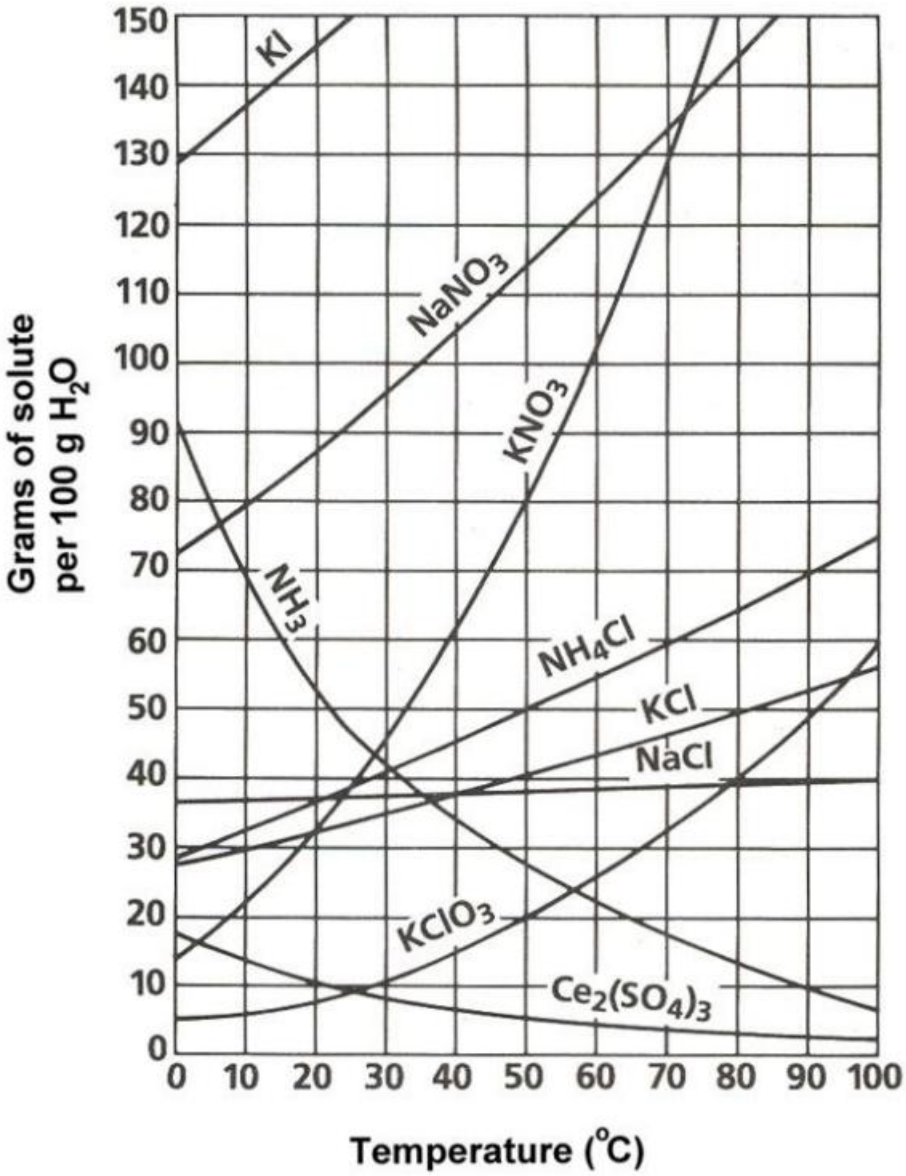
5. What is the concentration of a solution with a volume of 9.00 mL that contains 0.025 grams of iron (III) hydroxide?

6. Extra Credit: What is the concentration of a solution containing 3.3 mL of solvent and 12 grams of ammonium sulfite?

Lesson 130: Solubility Curve Worksheet

Chemistry with Lab

Use this solubility graph curve to answer the questions.



1. What are the customary units of solubility on solubility curves?

2. Define solubility.

3. According to the graph, the solubility of any substance changes as ____________________ changes.
4. List the substances whose solubility decreases as temperature increases.
5. Which substance is least affected by temperature changes?
6. How many grams of ammonium chloride (NH_4Cl) at 50°C?
7. _______________ and ___________ have the same solubility at approximately 78°C.
8. Which compound is least soluble in water at 10°C?
9. How many grams of KNO_3 can be dissolved at 50°C?
10. Are the following solutions unsaturated, saturated, or supersaturated?
    a. 45g of NaNO3 in 100 g of water at 30°C ____________________
    b. 60g of KClO3 in 100 g of water at 90°C _____________________
11. How many grams of sodium chloride, NaCl are required to saturate 100 grams of water at 100°C?
12. How many grams of $NaNO_3$ are required to saturate 100 grams of water at 75°C?
13. How many grams of KCl will saturate water at 20°C?
14. At what temperature would 25g of potassium chlorate ($KClO_3$) dissolve?

15. At what temperature would 60g of NH_4Cl dissolve?

16. 89g NaNO3 is prepared at 30°C.

    a. Will all of the salt dissolve? ___________

    b. What mass of NaNO3 will dissolve at this temperature? ________

17. If 50g of NH_4Cl is dissolved at 50°C, how many additional grams of NH_4Cl would be needed to make the solution saturated at 80°C?

18. At 50°C, how many grams of KNO_3 will dissolve?

19. At 70°C, how many grams of cerium (III) sulfate ($Ce_2(SO_4)_3$) dissolve?

20. Determine if each of the following is unsaturated, saturated, or supersaturated.

    a. 55g of NH_3 at 20°C _______________________

    b. 10g of $Ce_2(SO_4)_3$ at 10°C __________________

    c. 110g of KNO_3 at 60°C _____________________

    d. 65g of NH_4Cl at 80°C _____________________

    e. 12g of NH_3 at 90°C _______________________

    f. 78g of $NaNO_3$ at 10°C _____________________

    g. 145g of $NaNO_3$ at 80°C ____________________

    h. 35g of NaCl at 100°C _____________________

Lesson 132: Acids, Bases, and Salts Terms

Introduction to Acids, Bases, and Salts

1. **Acid** – Any of a class of substances whose aqueous solutions are characterized by a sour taste, the ability to turn blue litmus red, and the ability to react with bases and certain metals to form salts.
2. **Base** – Any of a class of compounds whose aqueous solutions are characterized by a bitter taste, a slippery feel, the ability to turn litmus blue, and the ability to react with acids to form salts.
3. **Salt** – Chemical compound formed when the hydrogen of an acid is replaced by a metal or its equivalent.
4. **Arrhenius acid** – A substance which releases hydrogen ions (H^+) in solution.
5. **Arrhenius base** – A substance which releases hydroxide ions (OH^-) in solution.
6. **Operational definition of an acid** – Any substance that increases the concentration of the H^+ ion with dissolved in water.
7. **Operational definition of a base** – Any substance that increases the concentration of the OH^- ion when dissolved in water.
8. **Brønsted-Lowry acid** – Any substance that can "donate" hydrogen ions (H^+) or protons to bases which "accept" them.
9. **Brønsted-Lowry base** – A homogeneous mixture or solid solution of two or more metals, the atoms of one replacing or occupying interstitial positions between the atoms of the other.

Indicators and the pH Scale

10. **pH** – A measure of the acidity or alkalinity of a solution, numerically equal to 7 for neutral solutions, increasing with increasing alkalinity and decreasing with increasing acidity. The pH scale commonly in use ranges from 0 to 14.
11. **pOH** – A logarithmic scale measuring the acidity (or basicness) of a solution. It is based on the amount of OH^- (hydroxide ions) in the solution.
12. **Strong acid** – An acid with a high degree of dissociation in solution; for example, HCl or H_2SO_4.
13. **Weak acid** – An acid that does not ionize greatly; for example, acetic acid or carbonic acid.

14. **Strong base** – A base with a high degree of dissociation in solution; for example NaOH or KOH.
15. **Weak base** – A chemical base that does not ionize fully in an aqueous solution; for example, ammonia.
16. **Electrolyte** – A chemical compound that ionizes when dissolved or molten to produce an electrically conductive medium.
17. **Indicator** - Any of various substances, such as litmus or phenolphthalein, that indicate the presence, absence, or concentration of another substance or the degree of reaction between two or more substances by means of a characteristic change, especially in color.
18. **Litmus paper** - An unsized white paper impregnated with litmus and used as a pH or acid-base indicator.
19. **Phenolphthalein** – A white or pale yellow crystalline powder, $C_{20}H_{14}O_4$, used as an acid-base indicator, in making dyes, and formerly in medicine as a laxative. Because of its toxicity, it is no longer used in over-the-counter laxatives.

<u>Neutralization Reactions</u>

20. **Neutralization reaction** – A reaction between an acid and a base that yields a salt and water.
21. **Titration** – The process, operation, or method of determining the concentration of a substance in solution by adding to it a standard reagent of known concentration in carefully measured amounts until a reaction of definite and known proportion is completed, as shown by a color change or by electrical measurement, and then calculating the unknown concentration.
22. **End point** – The point in a titration at which no more titrant should be added. It is determined, for example, by a color change in an indicator or by the appearance of a precipitate.
23. **Equivalence point** – The point in a titration where the amounts of titrant and material being titrated are equivalent chemically.
24. **Hydronium** – - A hydrated hydrogen ion, H_3O^+.

Lesson 133: Acids, Bases, Salts Notes

Arrhenius definitions: When reacting with _______________,

- ___________ release ___________________ ions (ex _____________)
- ___________ release ___________________ ions (ex _____________)
- ___________ are ____________ compounds that release neither

  ________________ ions nor __________________ ions

Practice: Write “A” for acid, “B” for base, and “S” for salt.

________________ ______________ ______________

________________ ______________ ______________

Solution	**Acid (A), Base (B), or Salt (S)**	**Reacts with Mg (R or NR)**	**Conductivity (C or NC)**
HNO_3			
$NaOH$			
$NaCl$			
H_2SO_4			
HCl			
NH_4OH			
KNO_3			
KOH			
$HC_2H_3O_2$			

Operational Definitions:

ACIDS:

- __________ taste
- react with some ___________ to produce _______________ gas
- ____________________
- turn __________-___________ indicators different _____________
- react with ___________ to produce a _____________ and water

BASES:

- _______________ taste
- _______________
- ____________________
- turn __________-___________ indicators different _____________
- react with ___________ to produce a _____________ and water

Naming Bases and Salts

To name ____________ and ___________, follow the standard rules for naming _______________ compounds.

NaOH ______________________________

$CaCl_2$ ______________________________

Naming Acids

BINARY ACID

- only ______ elements
- first element is _________________
- named: hydro________ic acid
- HCl ______________________

TERNARY ACID

- __________ elements
- first element is _________________
- other elements part of a ________________________ ion
- naming does NOT require a ________________
- __________ → _____
- __________ → _____
- H_2SO_4 - _____________ acid
- H_2SO_3 - _____________ acid

Examples:

- H_3PO_4 _________________________
- HF ____________________________
- $HClO_2$ _________________________

Name the following acids:

_______________ ______________________________

_______________ ______________________________

_______________ ______________________________

_______________ ______________________________

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 137: Indicators and the pH Scale Notes Chemistry with Lab

Acid/Base indicator: __________________ that changes _____________ in the presence of an ____________ or a __________.

Solution	Litmus paper ($R \rightarrow B$, $B \rightarrow R$, or NC)	Phenolphthalein (color or NC)	Bromothymol Blue (color or NC)	Cabbage Juice (color or NC)
HNO_3				
NaOH				
KOH				
H_2SO_4				

Conclusion Questions:

1. Litmus turns ______________ in an acid and ____________ in a base.
2. Phenolphthalein (phth) turns __________ in an acid and ____________ in a base.
3. Bromothymol blue turns ______________ in an acid and ____________ in a base.
4. Cabbage juice turns _____________ in an acid and ____________ in a base.

Strong Acids: dissociate ______________ in ____________ (ex: __________)

Weak Acids: dissociate ________________ in ____________ (ex: __________ or __________________)

Strong Bases: dissociate ______________ in ____________ (ex: __________)

Weak Bases: dissociate _______________ in ____________ (ex: __________ or _____________) pH = __________

0______________ 7 _____________ 14

Determine the pH of a solution of HCl that has a molarity of 1×10^{-4} M.

Calculate the pH for a solution of HNO_3 with a molarity of ________________.

Calculate the pH for a solution of H_2SO_4 with a molarity of ___________.

$[H^+][OH^-]$ = ___________________

Calculate the pH of a solution of NaOH with a molarity of 3.0×10^{-2} M.

Find the pH for a solution of $Ca(OH)_2$ with a molarity of _____________.

Calculate both the hydrogen ion concentration and the hydroxide ion concentration for an aqueous solution that has a pH of _____________.

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 142: Neutralization Reactions Notes

- A ___________________ reaction is the reaction between an __________ and a __________ to produce a _________ plus ________________.
- A __________ is any compound containing the ________________ from a base and the ______________ from an acid.

Write the neutralization reaction when H_2SO_4 reacts with KOH. Label the acid, the base, and the salt.

H_2SO_4 + KOH $\rightarrow$

Write the neutralization reaction when ____________ acid reacts with ____________________ hydroxide.

- A ___________________ is a laboratory method used to determine the ____________________ of an acid or base in _________________ by performing a ___________________ reaction with a _________________ solution.
- In a _________________ solution, the __________ of ________________ ions must equal the __________ of ________________ ions.

$$moles\ \_\_ = \frac{moles\ \_\_}{1\ mole_A}(M_A)(V_A)$$

$$moles\ \_\_ = \frac{moles\ \_\_}{1\ mole_B}(M_B)(V_B)$$

$$\frac{moles\ \_\_}{1\ mole_A}(M_A)(V_A) = \frac{moles\ \_\_}{1\ mole_B}(M_B)(V_B)$$

Example Titration Problem:

Find the molarity of this sample of hydrochloric acid (HCl) by neutralizing it with 0.5M sodium hydroxide (NaOH).

Volume of HCl	Volume of NaOH

- The ________________ of a titration is the point at which the indicator changes _____________ indicating that _______________________ has been reached so the ___________ of ____________________ ions and the ____________ of ________________ ions are _____________.

$$\frac{moles\ H^+}{1\ mole_A}(M_A)(V_A) = \frac{moles\ OH^-}{1\ mole_B}(M_B)(V_B)$$

In a titration of ______ with ______, _______mL of the base were required to neutralize 10.0 mL of a ______M _______. What is the molarity of the KOH?

60.0 mL of ______ molar ________ were needed to neutralize 30.0 mL of __________. What is the molarity of the acid?

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 146: Thermochemistry Terms

1. **Kinetic energy** – The energy possessed by a body because of its motion, equal to one half the mass of the body times the square of its speed.
2. **Potential energy** – The energy of a particle or system of particles derived from position, or condition, rather than motion. A raised weight, coiled spring, or charged battery has potential energy.
3. **Heat** – A form of energy associated with the motion of atoms or molecules and capable of being transmitted through solid and fluid media by conduction, through fluid media by convection, and through empty space by radiation.
4. **Temperature** – A measure of the average kinetic energy of the particles in a sample of matter, expressed in terms of units or degrees designated on a standard scale.
5. **System** – Any set of interrelated parts. An *open system* allows mass and energy to circulate into and out of it; a *closed system* gives and receives energy but not mass.
6. **Surroundings** – Everything outside the thermodynamic system.
7. **Specific heat capacity** – Quantity of heat required to increase temperature of a body by one degree Celsius.
8. **Phase diagram** – A graph showing the pressures at which phase transitions between different states of a pure compound occur, as a function of temperature.
9. **Heat of fusion** – The amount of heat required to convert a unit mass of a solid at its melting point into a liquid without an increase in temperature.
10. **Heat of vaporization** – The amount of heat required to convert a unit mass of a liquid at its boiling point into vapor without an increase in temperature.
11. **Endothermic** – Characterized by or causing the absorption of heat.
12. **Exothermic** – Denoting a chemical reaction that releases heat into its surroundings.
13. **Activated complex** – An energetically excited state which is intermediate between reactants and products in a chemical reaction. Also known as transition state.
14. **Activation energy** – The energy, in excess over the ground state, which must be added to an atomic or molecular system to allow a particular process to take place.

15. **Heating curve** – A heating curve is a graph showing how a substance's phases (gas, liquid or solid) change while being heated.
16. **Absolute zero** – The theoretical temperature at which substances possess no thermal energy, equal to -273.15°C or -459.67°F or 0 K.

Lesson 147: Thermochemistry Notes

Temperature:

- a measure of the _______________ ______________ energy of the particles in a sample of matter
- does not depend on the amount of ________________ in the sample
- symbol is ______; unit is _________

heat:

- __________ amount of _______________ energy that flows because of a difference in ___________________.
- depends on _________ of sample
- symbol is ______; unit is ____(_________) (1 cal = 4.18 __)

Kinetic energy is _______________ in _________________

Potential energy is ______________ ______________

- Potential energy is hiding and cannot be _______________.
- Only _____________ in P.E. can be measured.

specific heat capacity:

- amount of ________ required to raise the ________________ of 1 ___ of substance 1 _______
- symbol is ______; unit is _________

Lesson 147: Thermochemistry Notes (cont.)

____

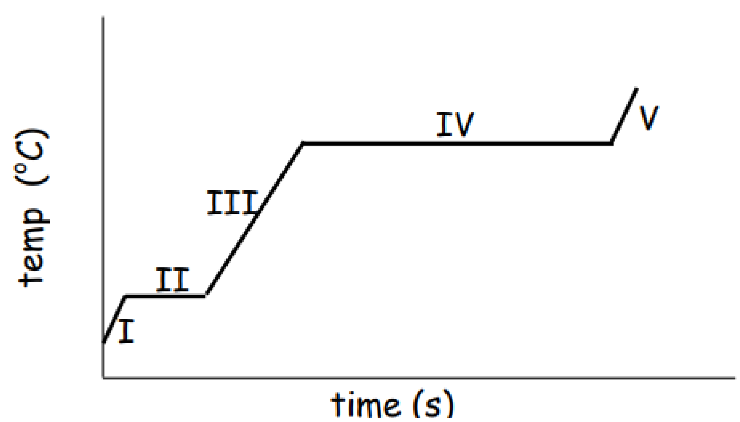
$$Q = m \times C \times \Delta t$$

______________ ______________ ______________________________

When heat (Q) is absorbed by a system, part of it (C) goes into storage as ____________________ energy and part of it is used to make the molecules move around _____________, raising the ___________________ (Δt).

**Why does sand get hotter in the day and colder at night than the water?

<u>Heating Curve for Water</u>



I:

Heat is being used to raise the ___________________ of the _________.

Q = ______ x ________ x ________

II:

Heat is being used to turn solid to ________________. (phase change)

Q = ______ x ________

heat of fusion - _______________ required to change 1 g of

____________ to ________________

III:

Heat is being used to raise the ___________________ of the __________.

Q = ______ x ________ x ________

IV:

Heat is being used to turn liquid to _____. (phase change)

Q = ______ x ________

heat of vaporization - __________ required to change 1g of

_______________ to _______

endothermic change: (________________ is an example.)

- _______________ or _______________ change in which a ____________ absorbs ________ from its __________________
- _____ $\rightarrow$ _____ (Heat seems to ______________________.)
- _____ of system ________________ and it becomes less _____________.

  (____________________ is another example.)

exothermic change:

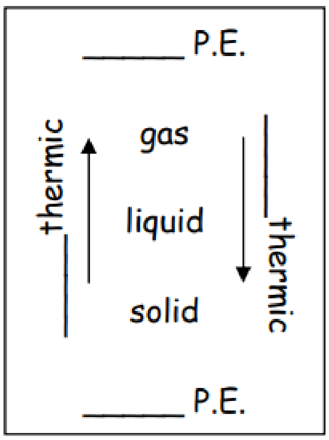
- physical or chemical _____________ in which a system ______________ ________ heat to its _______________________

Lesson 147: Thermochemistry Notes (cont.)

- _____ → _____ (Heat seems to ___________ out of ________________)
- _____ of system ________________ and it becomes ___________ stable.

Ex. – Why does your skin feel cool when you get out of the pool?

Think about these steps to answer the question:



Identify the system - ____________________

goes from liquid (____ P.E.) to _____ (_____ P.E.)

This is an ____________________ change. In this

type of change, the system (the water) __________

heat from the surroundings.

Identify the surroundings - __________

Your skin feels ____________ because it ___________ heat. The heat was used

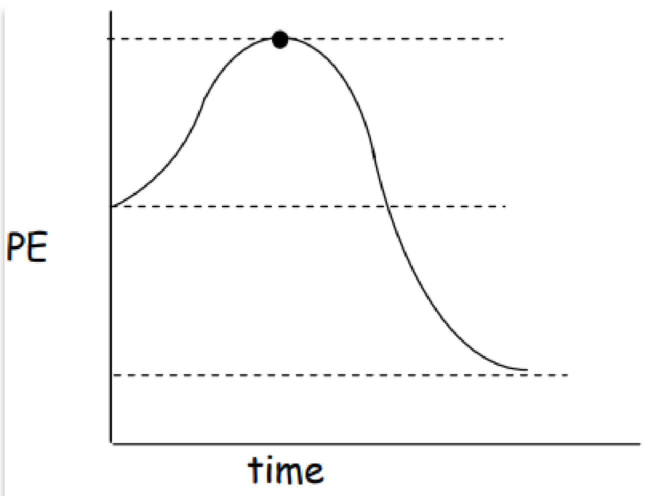
to ___________________ the water.

Why do farmers spray fruit on trees with water when the temperature is going to drop below freezing? *Identify the system and surroundings and make the statements about them (as done above)*.

Lesson 147: Thermochemistry Notes (cont.)

Energy Diagram of Chemical Change:

Label the chart:



As molecules get closer, their electron clouds ___________ each other, and their P.E. (increases, decreases)

The ________________ complex is highest point in P.E.

The energy required to reach the complex is called the ____________ energy.

Products are (higher, lower) in P.E. than reactants and are (more, less) stable.

This reaction is _______thermic.

Problem Set #1: Draw the P.E. diagram shown and label the following: *reactants, products, activation energy, activated complex, ΔH_r (+ or -)*

Products are (higher, lower) in P.E. than reactants and (more, less) stable.

This reaction is _______thermic.

When Act E is high, the reaction is (slow, fast)

Lesson 147: Thermochemistry Notes (cont.)

Sketch a diagram for these reactions:

slow, exothermic | faster, endothermic | faster, exothermic

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

Table I.

	Ag	Water
Mass		
Initial Temperature		
Final Temperature		
Change in Temperature		
Heat Content		

Record your observations as you run the experiment.

Data Analysis and Interpretation

1. Which substance (Ag or water) loses heat when they are combined? Which substance gains heat when they are combined? Which process is endothermic and which is exothermic?

2. Calculate the heat transferred to or from Ag. Use the equation $q = mC_s\Delta t$ (q is heat in Joules, m is mass, C_s is the heat content, and Δt is the change in temperature).

3. Calculate the heat transferred to or from water ($q = mC_s\Delta t$).

4. Compare the heats associated with the Ag and water. Make a generalization concerning these heats.

5. How would your results have been different if you had used different amounts of Ag and water starting at different temperatures? Try this out by doing new experiments and record your data and conclusions below.

6. Draw a picture of the microscopic view of a piece of solid silver metal in liquid water.

7. Describe how heat was transferred between the two substances. What did you observe happening during the period just after the silver and water were added to each other? How about when the final temperature was reached? How does this help explain how heat is transferred?

	Al	Water	Cu	Water	Fe	Water
m						
t_i						
t_f						
Δt						
C_s						
q						
molar heat capacity						

Data Analysis and Interpretation

8. Calculate the heat lost or gained by each metal. Show your work for one of the calculations.

9. Compare the results for all four metals. How are they different?

10. Which metal would make the best cookware? Why?

11. Calculate the molar heat capacity for Al, Cu, and Fe in units of $\frac{J}{mol\ ^{\circ}C}$. Record the value on Table II.

	Unknown I	Water	Unknown II	Water
m				
t_i				
t_f				
Δt				

12. Calculate values for the specific heats for the two unknown metals.

13. Based on the values in question 11, estimate the molar heat capacity for each unknown metal.

14. Calculate the molar mass of each unknown metal.

15. Assuming the unknown metals are pure substances, identify them.

Lesson 158: Nuclear Chemistry Key Terms

1. **Alpha particle** – A positively charged particle, indistinguishable from a helium atom nucleus and consisting of two protons and two neutrons.

2. **Beta particle** – A high-speed electron or positron, especially one emitted in radioactive decay.

3. **Gamma ray** – Electromagnetic radiation emitted by radioactive decay and having energies in a range from ten thousand (104) to ten million (107) electron volts.

4. **Fission** – A nuclear reaction in which an atomic nucleus, especially a heavy nucleus such as an isotope of uranium, splits into fragments, usually two fragments of comparable mass, releasing from 100 million to several hundred million electron volts of energy.

5. **Fusion** – A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.

6. **Isotope** – One of two or more atoms having the same atomic number but different mass numbers, due to a different number of neutrons in the nucleus.

7. **Nuclear reactor** – Any of several devices in which a chain reaction is initiated and controlled, with the resulting heat typically used for power generation and the neutrons and fission products used for military, experimental, and medical purposes.

8. **Radioactivity** – Spontaneous emission of radiation, either directly from unstable atomic nuclei or as a consequence of a nuclear reaction.

9. **Artificial transmutation** – An artificially induced nuclear reaction caused by the bombardment of a nucleus with subatomic particles or small nuclei.

10. **Bombardment** – - the act (or an instance) of subjecting a body or substance to the impact of high-energy particles.

11. **Nucleon** – A proton or a neutron, especially as part of an atomic nucleus.

12. **Half-life** – The time required for half the nuclei in a sample of a specific isotopic species to undergo radioactive decay.

13. **Particle accelerator** – A device, such as a cyclotron or linear accelerator, that accelerates charged subatomic particles or nuclei to high energies.

Lesson 159: Nuclear Science Notes

Nucleons:

- are particles occupying the ________________
- consist of ___ charged ___________ and ___________ ____________
- have almost 2000 times the __________ of ________________
- are made up of ___________ and _______________

Nuclear Notation: atomic _________ (__)

= ____________ + ____________

$$^{27}_{13}\mathrm{Al}$$

Atomic ______________ (__)

= no. of ____________ (=____________)

$^{27}_{13}$Al has ____ protons and ____ neutrons, for a total of ______ nucleons.

$^{28}_{13}$Al has ____ protons and ____ neutrons, for a total of ______ nucleons.

<u>isotopes</u>: atoms of ________ element with different numbers of

_______________ (different ___________)

Forces acting on nucleons:

strong forces –

- forces of _______________ between nucleons
- are independent of the _____________ of the nucleon
- are short range (exist only between ____________ neighbors)

electrical force –

- force of ____________ between _____________ charged protons
- are long range

When are nuclei **unstable**?

a. large nuclei (Z>82) – electrical forces of _____________ are greater than strong forces of _______________
b. wrong neutron: proton _____________

A radioactive isotope:

- has an ______________ nucleus
- spontaneously emits a ________________ and ___________ into another _____________________

transmutation – changing into another ____________ through radioactive _______________

Types of Radioactive Emission:

	Symbol	Composition	Stopped by -
alpha			
beta			
gamma			

Nuclear Equations:

alpha decay:

$$^{214}_{84}Po \rightarrow \quad He +$$

beta decay:

$$^{209}_{82}Pb \rightarrow \quad e +$$

During beta decay, a neutron changes into a proton and an ______________

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

half-life:

- the ________ it takes for _________ the ___________ of a radioactive sample to ___________
- ranges from a fraction of a _____________ to billions of _______
- is ________ affected by _______________ conditions

The Chemistry Quiz

CR1. _____ CR2. _________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 162: Types of Nuclear Reactions Notes

natural transmutation – Uranium spontaneously decays.

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

artificial transmutation – bombardment of a stable isotope to force it to decay.

$$^{14}_{7}N + \, ^{4}_{2}He \longrightarrow \, ^{1}_{1}p + \_\_\_\_\_\_\_\_$$

nuclear "bullet"

When the bullets are ____________ charged, they are ______________ by the nucleus they are bombarding. To overcome the repulsions, they must be __________________ to very high speeds by _______________ accelerators.

nuclear fission – Heavy nuclei are bombarded with neutrons and split.

$$^{1}_{0}n + \, ^{235}_{92}U \longrightarrow \underbrace{^{142}_{56}Ba + \, ^{91}_{36}Kr + 3(^{1}_{0}n) + E}$$

Mass of particles produced is slightly _________ than the mass of the reactants.

This mass is converted into ___________ (E = _____)

- critical mass: ___________________ mass of __________________ material required for a ___________ ______________
- nuclear reactors: control fission __________ reactions to produce energy

dangers:

Lesson 162: Types of Nuclear Reactions (cont.) Chemistry with Lab

nuclear fusion – combination of _______ nuclei into ______ with release of

____

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

Mass of particles produced is much ______ than the mass of the reactants.

This mass is converted into ___________ (E = _____)

List advantages and problems with using fusion as an energy source:

Lesson 167: Nuclear Decay

1. Fill in the table below and then use it to figure out what is happening during each type of decay – alpha, beta, and gamma.

Parent Isotope	Particle Emitted	New, Daughter Isotope	Alpha, Beta, or Gamma Decay?	# of protons lost or gained by "parent"	Change in mass number
$^{226}_{88}Ra \longrightarrow$	$^{4}_{2}He\ +$	$^{222}_{86}Rn$	*Alpha*	*Lost 2*	*Minus 4*
$^{214}_{84}Po \longrightarrow$	$^{4}_{2}He\ +$	$^{210}_{82}Pb$			
$^{47}_{20}Ca \longrightarrow$	$^{0}_{-1}e-\ +$	$^{47}_{21}Sc$			
$^{148}_{64}Ra \longrightarrow$	$^{4}_{2}He\ +$	$^{144}_{62}Sm$			
$^{14}_{6}C \longrightarrow$	$^{0}_{-1}e-\ +$	$^{14}_{7}N$			
$^{148}_{64}Gd \longrightarrow$	$^{0}_{0}Y\ +$	$^{148}_{64}Gd$			

2. What changes take place in the nucleus when an alpha particle is emitted?

3. What is the identity of an alpha particle?

4. What changes take place in the nucleus when a beta particle is emitted?

5. Which particle is associated with beta decay?

6. Fill in the missing numbers and elements from these nuclear reactions:

$^{40}_{\_\_}\_\_\_\_\_ \longrightarrow {}^{0}_{-1}e + {}^{40}_{20}Ca$	$\_\_\_\_\_ \longrightarrow {}^{4}_{2}He + {}^{226}_{88}Ra$
$^{35}_{14}Si \longrightarrow {}^{0}_{-1}e + \_\_\_\_\_$	$^{238}_{92}U \longrightarrow + {}^{4}_{2}He + \_\_\_\_\_$

$^{110}_{53}I \longrightarrow \_\_\_\_ + ^{106}_{51}Sb + ^{0}_{0}Y$	$^{140}_{56}Ba \longrightarrow \_\_\_\_ + ^{140}_{57}La$

7. a) Show the alpha decay of radon-198

   b) Show the beta decay of uranium-237

8. Plutonium-244 undergoes gamma decay. What are the products of this reaction?

9. Does the identity of an atom change during radioactive decay? Why or why not?

10. List the 3 types of radiation (alpha, beta, gamma) in order from least penetrating to most penetrating.

11. What is the difference between nuclear fusion and nuclear fission?

1. **Kinetics** – The study of the factors that govern how rapidly reactions occur.
2. **Collision theory** – The rate of a reaction is proportional to the number of collisions that occur each second between the reactants.
3. **Effective collision** – A collision in which the colliding particles approach each other at the proper angle and with the proper amount of energy.
4. **Catalyst** – A substance, usually used in small amounts relative to the reactants, that modifies and increases the rate of a reaction without being consumed in the process.
5. **Inhibitor** – One that inhibits, as a substance that retards or stops a chemical reaction.
6. **Enzyme**– Any of numerous proteins or conjugated proteins produced by living organisms and functioning as biochemical catalysts.
7. **Activation energy** – The minimum kinetic energy that must be possessed by the reactants in order to give effective collision (one that produces products).
8. **Potential Energy Diagram** – A plot of the change in potential energy that occurs during a chemical reaction.
9. **Transition state** – The moment during a reaction when the species involved have acquired the minimum amount of potential energy needed for a successful reaction.
10. **Reaction rate** – How quickly the reactants disappear and the products form.
11. **Reversible reaction** – a reaction in which the equilibrium constant is such that the reaction can be made to proceed at a detectable rate in either direction under appropriate conditions. A reversible reaction is a chemical reaction that results in an equilibrium mixture of reactants and products.
    - For a reaction involving two reactants and two products, this can be expressed symbolically as: $aA + bB \rightleftharpoons cC + dD$
    - A and B can react to form C and D or, in the reverse action, C and D can react to form A and B.
12. **Le Châtelier's principle** – When a system that is in dynamic equilibrium is subjected to a disturbance that upsets the equilibrium, the system undergoes a change that counteracts the disturbance and, if possible, restores the equilibrium.

13. **Stress** – An applied force or system of forces that tends to shift a reaction.
14. **Endothermic** – Characterized by or causing the absorption of heat.
15. **Exothermic** – Releasing heat.

Lesson 169: Reaction Rates Notes

Reaction rate:

- the change in ________________________ of the ____________________ per unit of time
- speed at which __________________ disappear and _____________ appear

The study of reaction rates is called _______________ ______________.

✓ In order for a reaction to occur, reacting ________________ must ________________ effectively.

✓ In an ________________ collision, the ______________ particles collide with the correct _________________ and the proper amount of ________________.

Factors affecting reaction rates:

- ___________ of reactants –
  - The _________________ of the reactants, types of _____________ involved, and state of ____________ all affect reaction rates.
- _____________________ of reactants –
  - Increasing the concentration of reactants → more _____________ collisions → increases ________ of reaction
  - Increasing the _____________ of the reactants (__________ only) → increases the _________________ → increases _______________ ________________ → increases _______________ __________

- _______________ of reactants –
    - Increasing ________________ of reactants → more ____________ _______________ → increased _________ of reaction.
- _______________ -
    - A substance that ___________ the rate of a reaction without being permanently ______________
    - Increases the rate of reaction by ____________ the required amount of ___________ to achieve ________________ collisions
- ____________ area of reactants –
    - Increasing surface _________ → more __________________ _______________ → increases _________ of reaction.

The Chemistry Quiz

CR1. _____ CR2. __________ 1. _______ 2. _______ 3. _______

4. ________ 5. _______

Lesson 172: Le Chatelier's Principle/Keg Notes Chemistry with Lab

A reaction in which the ____________________ can react to form the

____________________ is called a ____________________ reaction.

$$A + B \boxed{} C + D$$

Chemical ______________________ occurs when the _______________ in a

____________________ reaction form the ___________________ at the same

_______________ that _______________ form ____________________.

At equilibrium:

- the ______________________ of the reactants and products does not _______________
- the concentration of reactants can be ________ to, _______, or _______ the concentration of the products.

$$\_\_\_\_\_\_A + \_\_\_\_\_B \rightleftharpoons \_\_\_\_\_\_\_C + \_\_\_\_\_\_\_D$$

$$K_{eq} = \frac{[\quad]^c[\quad]^d}{[\quad]^a[\quad]^b}$$

If K_{eq} is…	… reaction is favored
= 1	neither
<1	
>1	

Ex. #1: Calculate K_{eq} for the following reaction using the given equilibrium concentrations. Then determine whether the forward or reverse reaction is favored.

$N_2O_4 \rightleftharpoons 2NO_2$ $[N_2O_4] = 0.23M$ $[NO_2] = 0.037M$

Ex. #2: Calculate K_{eq} for the following reaction using the given equilibrium concentrations.

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

$[N_2]$ = _________ $[H_2]$ = _________ $[NH_3]$ = ________

The ___________ reaction is favored.

LeChatelier's Principle:

When a _____________ is applied to a system in a state of equilibrium, the system reacts in a way to __________________ the stress.

Stress	System Will Shift
addition of a chemical	
removal of a chemical	
increase in temperature	
decrease in temperature	
*increase in pressure	
*decrease in pressure	
addition of a catalyst	

*applies to reactions involving gases only

examples: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ + heat

- When H_2 is added to the system, the reaction shifts to the ___________ to use up the extra H_2. The amount of _____________ produced will increase.
- What if ______________ is removed? The reaction shifts to the _______ trying to replenish the N_2.
- When the temperature increases, the reaction shifts to the ___________.
- When the pressure of the system increases, the reaction shifts to the ___________, toward the side with ____________ particles.
- When a catalyst is added, there will be _______ _______________.

When a system in equilibrium…

shifts:	[products]	[reactants]
to the right		
to the left		

The Chemistry Quiz

CR1. _____ CR2. _________ 1. ______ 2. ______ 3. ______

4. _______ 5. ______

Lesson 172: Le Chatelier's Principle Datasheet

Table I:

Ion	Color
K^+	
Cl^-	
SCN^-	
Fe^{3+}	
$Fe(SCN)^{2+}$	

Use the table on the left to determine the colors of reactants and products. Write the color in the blanks above the equation below.

Table II:

colors: ____________________ ________________________

$$Fe^{3+} + SCN \rightleftharpoons Fe(SCN)^{2+}$$

Chemical Added	Color Change	Direction of Shift
$FeCl_3$		
NaOH		
KSCN		

Conclusions:

According to LeChatelier's Principle, when a___________________ is applied to a system in equilibrium, the system will readjust to ______________________ the stress, restoring a state of equilibrium.

For each procedure in Table II, identify the stress (ex. – addition of a reactant, removal of a product, etc.) and the reason for the shift in equilibrium (ex. – shift to the right uses up reactants):

Stress	**Reason for Shift**
______________________	______________________________
______________________	______________________________
______________________	______________________________

Hint: NaOH reacts with Fe^{3+} to form solid $Fe(OH)_3$

Lesson 174: Exploring Equilibrium Lab

Look at the screen and observe everything you can find out about the reaction pictured:

$A + BC \rightleftharpoons AB + C$

1. What type of reaction is it? Is it endo or exothermic?

2. Predict what will happen when 50 A's are added to the box and 50 BC's are added.

3. In the box labeled "current amounts," enter 50 for A and 50 for BC.

    a. Was your prediction correct? Describe and explain any differences.

    b. Describe the nature of dynamic equilibrium when small numbers of particles (such as 50, as compared to 6.02×10^{23}) are present.

4. Predict what will happen when the temperature is raised so it is NOT above the activation energy max but IS above the energy level of the products.

5. Raise the temperature as described. Did your prediction come true? Describe and explain any differences.

6. Predict what will happen when the temperature is raised so it is above the activation energy max.

7. Raise the temperature as described. Again, was your prediction correct? Describe and explain any differences.

8. What did you notice about the rate at which reactants/products fluctuated between the three different temperatures? If you didn't notice anything, hit "reset all" and test it again.

9. Did temperature affect equilibrium position? Did it affect it in the way you expected? Explain.

10. Did temperature affect reaction rate? Did it affect it in the way you expected? Explain.

11. Write the equilibrium expression for this reaction. Because of the way the numbers fluctuate, it would be hard to calculate the equilibrium constant, but how would you do it if the numbers ever settled down?

1. Match the unit of measure to the metric equivalent:

    _____ gram a. length

    _____ Celsius b. mass

    _____ liter c. volume

    _____ meter d. temperature

2. Number the steps of the scientific method in the correct order:

    _____ Construct hypothesis

    _____ Ask a question

    _____ Report results

    _____ Do background research

    _____ Test with experiment

3. Write 0.0056 in scientific notation.

4. Of what does an atom consist?

5. What part of the periodic table is a group?

6. A unit of magnesium carbonate contains one magnesium atom, one carbon atom, and three oxygen atoms. What is its chemical formula?

7. Who postulated the term "quanta" to explain light?

8. What term describes the average of the masses of all of an element's isotopes (often displayed as a decimal number on the periodic table of the elements)?

9. A proton has what electrical charge?

10. Who arranged the elements on the periodc table by increasing atomic number, like we see it today?

11. What is the term for the energy required to remove an electron from a neutral atom?

12. ________ bonding produces substances that are hard, or difficult to crush.

13. When an atom gains electrons, an ion with a ________________ charge is formed.

14. Determine whether the compound NH^3 is ionic or covalent and name it appropriately.

15. Determine whether the compound zinc sulfide is ionic or covalent and write the appropriate formula for it.

16. Use the activity series to write balanced checmical equations for single replacement reactions. If no reaction will occur, write "NR."

    $Cl_2(g) + NaBr(aq)$

17. Solid lithium hydroxide reacts with carbon dioxide to form solid lithium carbonate and liquid water.

18. What is the molecular formula for a compound with an empirical formula of NO_2 and a molar mass of 92.0 g/mol?

19. Convert 45g Se to moles of Se.

20. Lithium combines with oxygen according to the following equation. How many grams of Li_2O can be formed when 2.00g of lithium reacts with 2.00g of oxygen? (Atomic mass: Li = 6.94, O = 16.00)

$$4Li + O_2 \rightarrow 2Li_2O$$

a. 3.63 b. 3.74 c. 4.00 d. 3.12

21. Zinc and sulfur react to form zinc sulfide. How many grams of ZnS can be formed when 12.0g of Zn reacts with 6.0g of S? (Atomic mass: Zn = 65.38, S = 32.06)

$$Zn + S \rightarrow ZnS$$

a. 19.7 b. 17.9 c. 13.8 d. 18.5

22. When atmospheric pressure increases, how does the height of the column of mercury change?

23. A quantity of gas is collected over water at 15°C. The manometer indicated a pressure of 24.0 kPa. What would be the pressure of the dry gas? (The constant needed is 1.7 kPa.)

24. How many liters of 4M solution can be made using 100 grams of lithium bromide?

25. What is the concentration of a solution with a volume 3.3mL that contains 12 grams of ammonium sulfite?

26. Define solubility.

27. (Refer to the linked graph for numbers 27 and 28.) At 50°C, how many grams of KNO_3 will dissolve?

28. _______ and __________ have the same solubility at approximately 78°C.

29. Name this acid: H_2SO_3.

30. Name this base or salt: $MgSO_4$.

31. Calculate the amount of heat released when 25.0g of water at 25.0°C cools to 0.0°C.

32. Calculate the amount of heat released when the same sample from number 31 freezes.

33. What changes take place in the nucleus when an alpha particle is emitted?

34. Which particle is associated with beta decay?

35. In the reaction $2NO(g) + H_2(g) \rightleftharpoons N_2O(g) + H_2O(g) + 36\ kJ$, which direction will the equilibrium shift if the pressure is increased?

a. left b. right

36. How would an increase in pressure affect the $[H_2]$ in the following reaction?

$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$

a. increase b. decrease c. no change

37. List the 3 types of radiation (α, β, γ) in order from least penetrating to most penatrating.

38. Does the identity of an atom change during radioactive decay? Why or why not?

39. What is the word equation for the neutralization of a strong acid and strong base?

40. Which answer best describes a substance whose formula is O_2?

a. a compound b. an atom c. an element

41. Calculate the molar mass for ethane, C_2H_6.

42. When solid potassium nitrate is heated, it decomposes to solid potassium nitrite, and oxygen gas is evolved. The type of bonding that results in volatile liquids is:

a. ionic b. covalent c. metallic

43. Which element is a member of the Transition Metal family?

a. Zn b. Mg c. At d. Ca

44. What is the term for the region of space where an electron is likely to be found?

a. Orbital b. Home c. Sector d. Section