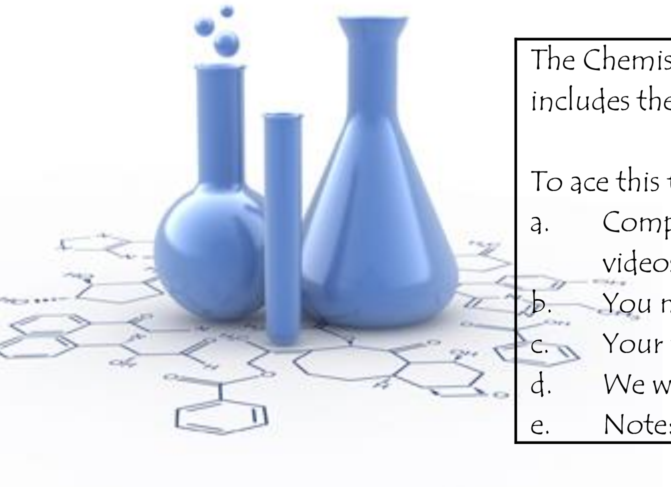
accelerated chemistry final exam review packet

Name______________________________ Period________________ final exam review packet 2017-2018

how to ace the chemistry final exam: your review packet



The Chemistry Final Exam Consists of approximately 80% multiple choice questions and 20% short answer questions based primarily on the second semester. This includes the following topics

Bonding Reactions The Mole Gases Solutions Energy Acids and Bases

To ace this test:

a. Complete this review. Use the KEY to check your answers. This packet is worth 50 points on your final exam. As you complete this packet, rewatch any videos to fill in any learning gaps.
b. You may prepare an index card that you may use during the final exam ☺
c. Your final assignment is to wear your tie-dye t- shirt to the final exam; this assignment is worth 10 points.
d. We will review hard and fast before the final- be sure to show up for it. Also, if an after school review is available it will be announced in class.
e. Note: twenty point deduction if not written neatly

Chemistry Final Exam Review
some general questions

1. What is chemistry? ______________________________

2. An oxygen molecule contains a double bond; where the two atoms of oxygen share a total of how many electrons?_____________________________

3. Reacting oxygen with a hydrocarbon is called ____________ .

4. What are the coefficients for the oxygen molecule? 4 Al + ___ O2 → 2 Al2O3

5. What happens to the amount of matter during a chemical reaction?

____

6. What is the mass of the solution when 1 pound of salt is dissolved in 20 pounds of water? ______

7. Which processes do not involve a change in chemical properties?
Rusting/Fermenting/Boiling/Burning

____

8. Are atoms neutral or charged? _________________________________

7. Bonding

In our 7th unit, we investigated bonding. We began by the basic idea of an ionic (Transfer of Electrons), covalent (Electron Sharing), or metallic ("Sea" of Electrons) bonds. We then focused on ionic bonding. We considered the ions that elements in groups 1, 2, 3, 5, 6, and 7 (+1 +2 +3 -3 -2 -1) would form, and how they could combine to form ionic compounds, We learned to write formulas and names for these compounds, and we took a look at their properties based on the nature of the ionic bond and elements involved. This included both monoatomic and polyatomic (Cl^- and NO_3^+) ions. We finished by mixing together a lot of different ionic compounds, and predicted what products would form. For some classes this included memorizing all of the mono- and polyatomic ions, and balancing the chemical reactions of ionic compounds. We then briefly looked at metallic bonds which helps explain the malleability and electrical conductivity of these materials. Finally, we investigated covalent bonding. Consisting of the bonds formed by *sharing* electrons between nonmetals, these are often organic (carbon-based) substances with a huge variety of chemical structures and properties. We spent much of our time assembling these molecules with models using bond and shorthand notation. We observed numerous examples of isomers, and learned how to name these types of molecules.

Bonding questions

1. A chemical bond is the force that holds two atoms together. Draw a reasonable substance with three chemical bonds.

2. An ionic bond is a chemical bond involving electron transferss. Which of the following substances contains an ionic chemical bond? Circle the ionic compounds:

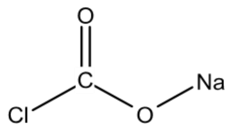
a. methanol (CH_4O) b. phosphorus trichloride, PCl_3 c. Aluminum oxide (Al_2O_3) d. iron (Fe)

3. A covalent bond is a chemical bond involving the sharing of electrons. These typically occur between non-metals. What is a *nonpolar* covalent bond?

4. A metallic bond is a bond between two metals. Based on the properties of metals, what is unique about the behavior of electrons in metals?

5. Circle the metallic bonds: a. Fe-Fe b. C-O c. Cl-Cl d. Li-Cl

6. What types of bonds are contained in the molecule below (note that a double bond counts as 2 bonds)



7. Name the following substance: CCl_4 ______________________________

8. Name the following substance: FeO_2 ______________________________

9. When naming ionic compounds that contain a metal from the s-block (groups 1 and 2) we do not need to specify the charge on the metal because____

10. The formula for calcium phosphate is________________________

11. NH_3 may be systematically named as ____________________, also known commonly as ________________

12. Ammonium nitrite is toxic to humans. What is its chemical formula? ______________________

13. Write a balanced chemical equation between silver (1) nitrate and magnesium phosphate.

14. Write the chemical reaction between iron (III) chloride reacts with magnesium carbonate

18. Write the formula for Vanadium (III) phosphate_________________________

21. Draw the electron dot structure for the war gas phosgene ($COCl_2$). Be sure to show each electron in the molecule.

8. Chemical reactions

In our eighth unit we explored the world of chemical reactions. This is the fundamental chemistry process: mixing things together and seeing what happens. We began by looking for the 4 COOL signs of a chemical reaction. We then focused on writing balanced chemical equations; this is probably the toughest part of the unit. Be sure to know how to write ionic compounds correctly (for example Al_2O_3)- we can't balance a reaction unless it is written correctly.

We then performed each of the 5 types of reactions, to get a feel for each one. In many cases we are now able to predict the products of a reaction. Ionic compounds tend to undergo replacement, combustion reactions often give the same products, and decomposition reactions often form stable common substances. Organic compounds are more complex, but involve the electron rich (nucleophilic) region of a molecule attacking (or being attracted to) the electrophilic (electron-poor) region of another.

To wrap up this unit we looked at the advantages of using solvents for chemical reactions. We took a sneak peek at acids and bases, which are really a subset of ionic compounds. We completed the unit by looking at reactions from the point of view of what really strikes your eye: precipitates.

Performing and balancing reactions shows us the ratio that substances react in, but it still doesn't tell us how much to mass out to do these reactions properly. This is the topic of the next unit: Stoichiometry and the Mole.

To ace this unit review your notes, labs, worksheets, and this how to ace it guide

Chemical reactions questions

1a. What are the 4 signs of a chemical reaction?

b. Suggest another sign of a chemical reaction:___________________________

c. Give an example of a sign of a chemical observations that is a false positive: it is a sign of a chemical reaction, but no chemical reaction is occurring.

2. Write a balanced chemical equation describing the reaction of sodium chloride and magnesium:

____

3. Write a balanced chemical equation for burning hexane (C_6H_{14}):

____.

4. Practice writing these ionic formulas: calcium hydroxide_________________ and iron(II) sulfate_________________. Ion charts will be available, as will the periodic table with the names of each element.

5. Precipitates- what is a precipitate? ____.

6. Solute and solvent. For 2 g of salt in 100 g of water, identify each___________________________

7. Consider the chemical reactions you experience each day. You get up. You are breathing… that's good. Now a biochemist could spend hours explaining it, but to us chemists you are taking in oxygen, it combines with carbon, and you exhale carbon dioxide. You're doing it right now. We can write this as a balanced chemical reaction indicating the physical states (s, l, g, or aq) of the reactants and products as: ___( ) + ___( ) → ___( ).

Ok, next you make a cup of tea. You put the kettle on, and your gas stove lights up. Yes, you have a gas stove, which uses a big propane tank next to your house. This is the combustion of propane (C_3H_8), which we can write as a balanced chemical reaction:

__________(__) + _____( ) → _____________(__) + _________________ ( ).

You drink the tea, and eat something. Your body converts this food to energy using enzymes to catalyze the reactions, and the process is complex. Time for school. You hop in the car and fire it up, and the combustion of octane (C_8H_{18}) takes place in your engine:

____________(__) + _____( ) → _____________(__) + _________________ ( ).

At school, everything is going fine, but then you start to feel sick, and you go see the nurse. You are diagnosed with a sour stomach, and the nurse suggests you add some milk to you tea next time. She gives you some antacid tablets (Na_2CO_3) which reacts with the excess hydrochloric acid in your stomach: _________(__) + _____( ) → _____________(__) + _________________ ( ). Now you feel much better. The rest of the day is yet another wonderful day of learning at Guilford High School you confidently ace your chemical reactions test.

9. The mole

In the previous unit (chemical reactions) we learned how chemicals qualitatively react.
In our 9th unit we took a quantitative look at chemical reactions by studying the mole.
We took a practical approach, emphasizing how the mole (usually referred to as molar mass) is used everyday to perform chemical reactions. The key idea is that each element has a unique characteristic mass- the average atomic mass. The mole is simply that mass in grams. Since that number is scaled to the mass of the substance it means it will always equal the same number of particles. This number is Avogadro's number: 6.2×10^{23}. Practicing chemists very rarely use Avogadro's number, and very often use molar mass. This unit emphasizes molar mass: the practical application of the mole.
Still not clear? Well, the examples are easier to follow. 12.011 grams of pure carbon (for example a 12.011 gram diamond) is a mole of carbon and would contain 6.02×10^{23} atoms of carbon. 16 grams of methane (CH_4) would be approximately one mole of methane (12 grams of carbon, and 1 gram for each hydrogen atoms) and would contain 6.02×10^{23} molecules of methane.
With the mole it is easy to create recipes that allow chemicals to react in exactly the right amounts, with nothing wasted. For example, reacting 12 grams of carbon (1mole) with 32 grams of oxygen (that's one mole of O_2) will produce exactly one mole (44 g) of carbon dioxide.
Using the mole we can perform any chemical reaction using the right amount of reactants, so that nothing is wasted. We can also predict the exact amounts of each product formed.
Using the mole, we can also calculate the percent composition of any substance, a great tool for chemical identification. Water, for example, is always 89% oxygen, and 11% hydrogen. We know this because each mole of water must consist of one mole of oxygen atoms (16g) and 2 moles of hydrogen atoms (2g)…so water is 16/18 oxygen and 2/18 hydrogen by mass.

Mole questions

Know what the mole is-both in terms of how many particles, and in terms of mass.

1. A mole of carbon dioxide has a mass of _____g and contains _________molecules

Be able to measure out one mole of any substance, or any fraction thereof.

2. To give me a mole of water you would mass out ________g.

Be able to perform a chemical reaction on any scale using the mole. This is perhaps the most common use of this handy topic.

3. To make 2 moles of water from the elements, the balanced equation is:

______ + _______ → __________

So I should mix _____g of hydrogen and _____g of oxygen

Be able to predict the percent composition of any substance.

4. Ammonia (NH_3): ______% N, ______% H

We should be familiar with all types of mole conversions: mole-mole (1 step), mole-gram or gram-mole (2 steps), and gram-gram (3 steps).

5. The combustion of one mole of hydrogen with excess oxygen will produce ________moles of water.

6. The combustion of 4 moles of hydrogen with excess oxygen will produce _______ grams of water.

7. The combustion of 4 moles of oxygen with excess hydrogen will produce ________ grams of water

We should be familiar with the concept of a limiting reactant. This is indispensable whenever any amounts of reactants are combined. Unless the stoichiometric amounts are used, this will create a situation where one reactant is in excess, and the other is limiting. We used a non-intuitive but useful method for solving these problems: we find out how much product each reactant will produce, and the lower amount "wins" – it defines the limiting reactant and therefore indicates how much product will form.

8. What will happen when one gram of hydrogen is combined with one gram of oxygen?

The limiting reactant is _______________, and this reactant will produce _____________g H_2O.

Finally, we learned to calculate the yield of a reaction; this is the (actual yield/the theoretical yield) x 100.

9. For example, if this reaction above produced 0.1 grams of water, the yield would be (_______g/_______g) x 100 =_______% yield.

10. How to balance chemical reactions- try writing a balanced chemical equation for reacting barium chloride with sodium fluoride. Put your molar masses below to help answer all the questions. Do this carefully-it will be used in several subsequent questions.

Coefficients (the numbers in front, like **2** H_2O) vs. Subscripts (the little numbers like H_2O)

11. In the equation above put circles around each subscript, and squares around each coefficient.

12. Excess reactants

In the example above, ___________________________________ is the excess reactant, and __________grams of it remain in the pot after the reaction is complete.

13. Theoretical yield (this is 100% yield).

In the example above the theoretical yield of sodium chloride is _________g.

14. Actual yield

In the example above, if 50 grams of sodium chloride was obtained, that is a _______% yield.

10. Gases

When atoms and molecules contain enough energy to break free of any attractions to other substances around them, they become gases. In many ways their physical behavior becomes much simpler, like eternally moving balls in a box. We explored how they are altered by changes in temperature, pressure, and volume, and derived formulas for each- Boyles, Charles, and Gay Lussacs. We considered how their speeds vary with size (Graham's Law), and how we can combine these behaviors using the ideal gas law.

Circle the macroscopic (visible) properties of liquids, solids, and gases.

1. A solid has/does not have a fixed volume and will/will not assume the shape of a container.

2. A liquid has/does not have a fixed volume and will/will not assume the shape of a container.

3. A gas has/does not have a fixed volume and will/will not assume the shape of a container.

Accurately predict the behaviors of molecules as they undergo changes of pressure, temperature, volume, or the number of molecules.

4. As the temperature increases in a non-expandable container like a sealed can, the ________________ increases as the _________________ remains constant.

5. As the temperature increases in an expandable container like a balloon, the _______________ increases and the _____________________ remains constant.

6. As the pressure increases in a collapsible container like a syringe, the _____________________ decreases and the ________________________ remains constant.

.

Apply the gas laws to predict how one physical change will affect another when the third remains constant.

7. If the pressure increases from 100 to 200 kilopascals in a 3 liter container and the temperature remains constant, what will the final volume be? (101.7 kilopascals = 1 atm).

8. Be able to solve gas law problems based on Avogadro's Principle.
What is the density of CO_2 gas in g/L at STP? Molar mass of CO_2 =44 g/mol

9. Be able to solve gas law problems involving multiple changes (ideal gas law).
How many moles of chlorine are contained in a 10 mL tank at 27°C and 3.5 atm?

11. Solutions

Solutions are central to chemistry because we can see them and we use them in our everyday lives. Gases are difficult to contain, or even see. Solids don't react well because of surface area issues. Solutions, on the other hand, are easy to see, react, store, and work with. It's no surprise, then, that solutions are all around us. We can find them in the grocery store, at the gas station, in our bodies.

This has been predominantly a hands-on unit, which we summarized with a laboratory experiment where we created a solution of sugar and water that had the same density of a plastic bead. Once we familiarized ourselves with the vocabulary, we learned how to prepare solutions of different concentrations, and how to change their concentration ($C_1V_1=C_2V_2$).

To help you ace this unit, we begin with a story to sharpen your language skills in this unit. Then we present some situations where solutions need to be prepared and their concentrations need to be adjusted. We finish with a road salt example of colligative properties in action.

Don't forget to review your worksheets, PowerPoints, and labs before you take the solutions test. Read the story below and fill in the blanks and answer the questions as you go. The story is designed to include all of the new vocabulary and techniques you have learned.

1. Today I decided to make rock candy I mixed 100 grams of sugar with 200 g water, so my solute is __________and the solvent is _____________, and since the resulting mixture was clear and colorless it was _________________. This solution has a concentration of ______% by mass. It took a while for the sugar to dissolve, probably because the big chunks of sugar made the *molecular* process of __________________ slow. I was surprised to see how much sugar dissolved in water, sugar is highly ________________ in water.

Concentration and dilution: For all of these questions you have 29 grams of table salt (density 2.16 g/mL) in 500g of water (density 1.0 g/mL).

2. Describe how to prepare this solution

____

3. Calculate the percent salt by mass (water has a density of 1g/mL)

____

4. Calculate the percent salt by volume (table salt has a density of 2.16g/mL)

____

5. Calculate the molarity of the solution (table salt had a molar mass of about 58g/moL)

12. Energy

In our 12^{th} unit we investigated **energy**, both from a scientific and environmental perspective. We began by considering the **primary sources of energy** we use, and their environmental consequences, as well as their abundance. This is perhaps the most important information needed to answer our essential question for this unit:
We looked at some efforts to answer this question in the form of transportation when we took a look at some **eco-friendly cars**. We then focused on the hard science that relates to energy: what it is, **types of energy**, **enthalpy**, and **exothermic** and **endothermic** chemical reactions. During this time we performed two experiments where we measured the **specific heat** of an unknown metal, and the **energy contained in a potato chip**, using a **calorimeter** of our own design.

We finished by considering the **energy required to heat water**, including the phase changes that may be involved, and we familiarized ourselves with **free energy**, which includes the esoteric concept of **entropy**.

To ace this unit be familiar with the terminology associated with energy, know how to measure it, think about the sources of energy we use, and what changes we must make for the sake of our planet. As usual, review your notes, worksheets, and lab experiments, and answer all of the questions below.

In our next unit we will ask why some reactions such as explosions are rapid, while others such as rust are slow- this is our **Rates of Reaction** unit, coming up.
Useful information to be provided on test:
$q = mc\Delta T$
where
m = mass (g), c = specific heat (J/g^{0}C; see examples for H_2O below), and DT = change in temperature (^{0}C)
$c_{water(l)}$ = 4.18 J/g^{o}C
$c_{water(s)}$ = 2.03 J/g^{o}C
$c_{water(g)}$ = 2.01 J/g^{o}C
ΔH_{vap} = 2260 J/g
ΔH_{fus} = 334 J/g
At 1 atm:
Water boils/condenses at 100^{o}C
Water melts/freezes at 0^{o}C
1 Nutritional Calorie = 4 BTU (British Thermal Units) = 1000 calories = 4184 joules = 0.0016 KWH
$\Delta G = \Delta H - T\Delta S$
Where ΔG = change in free energy (J), ΔH = change in enthalpy (J), T = temperature (K), and ΔS = change in entropy (J/K)

1. Briefly define

a. Energy:

b. Enthalpy :

c. Exothermic:

d. Endothermic:

e. Specific Heat:

l. Nutritional Calorie:

m. Joule:

n. Free Energy:

o. Entropy:

p. Bond energy:

2. Identifying a mysterious substance by finding it's specific heat:

Ten milligrams of an unknown metal at 200 ^{O}C is placed in 30,240 milliliters of water in a well insulated container, raising the water temperature from 25 to 47 ^{O}C, and cooling the metal from 200 to 47 ^{O}C. What is the specific heat of the unknown metal? (Water has a specific heat of 4.184 j/g ^{O}C and a density of 1 g/mL; 1000 milligrams = 1 gram). Remember, place all of your answers in the answer key, and don't forget your units.

Draw a labeled model of what is happening in your answer.

3. Which power plant produces carbon dioxide in large amounts when generating power?

1. The national ignition facility in Berkeley California, which is based on nuclear fusion
2. The biomass power plant in Hartford, CT
3. The nuclear fission power plant in Waterford Gas (Millstone)
4. The combined gas and oil power plant in Middletown (which exploded in 2010)
5. The 15 megawatt fuel cell power plant in Bridgeport, CT, which is based on the combustion of natural gas.

4. A 3.2 gram potato chip that can heat 10 grams of water 14 degrees hotter contains _____ nutritional calories.

a. 364,211
b. 1.89×10^{9}
c. 3.64×10^{21}
d. 0.00000000364

5. Which of the following are the main products of nuclear fusion?

a. Carbon dioxide and water
b. Barium-141 and krypton-92
c. Tritium, also known as H-3
d. helium
e. D_2O, also known as heavy water
f. unobtainium

6. Write a balanced chemical equation for the combustion of hexane ($CH_3CH_2CH_2CH_2CH_2CH_3$)

7. Given 132.8J of energy is required to heat 11.17g of aluminum from 15.73°C to 28.94°C, find the **specific heat** of aluminum.

8. If the change in enthalpy of a chemical reaction is +212 KJ, and the change in entropy is 60 J/K, at what temperature in Celsius does the reaction become spontaneous?

9. How much energy is required to heat 100 mL of methanol (C_2H_6O) from -40 to +80 degrees Celsius? Show your work on the answer sheet please.

Data for methanol (CH_4O)

molecular weight: 32 g/mol

density of methanol: 0.792 g/mL

melting point of methanol: -98 ^oC

boiling point of methanol: 65 ^oC

specific heat of methanol:

solid: 2.47 J/g ^oC

liquid: 2.51 J/g ^oC

gas: 1.91 J/g

enthalpy of fusion: 99.1 J/g^{o}C

enthalpy of vaporization: 1104 J/g

Since the final units (rates, equilibrium and acids and bases) were just covered, we can omit reviewing them😊

You should page through the unit packet and review the test you recently took, correcting any errors.

Congratulations! You have completed the final exam review packet. Bring it to the final exam. Be aware that this is not an exhaustive review- you should look at all the old tests, watch any screencasts needed, and refamiliarize yourself with all the formulas for each chapter. Fill out an index card- you can use it on the final exam. Copy this over if it is not neat (20 point deduction if not written very neatly). Good luck on the final exam.