Recitation Worksheet (Optional Extra Practice)

Name: Key UGA ID:

Textbook:

Chemistry & Chemical Reactivity
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Instructions:

- This recitation worksheet is optional extra practice for Ch. 7.5-7.6, 8.1-8.2, lattice energy.
- You **do not** need to submit it to Gradescope.
- The answer key has been posted with this worksheet to eLC.
- A periodic table and formula sheet are attached to the end of this worksheet.

1. Which of the options provided below represents the second ionization energy of Ca?

$\boxed{C}$

A. Ca^{+} (aq) + e^{-} → Ca (aq)
B. Ca^{+} (aq) → Ca^{2+} (aq) + e^{-}
C. Ca^{+} (g) → Ca^{2+} (g) + e^{-}
D. Ca^{+} (g) + e^{-} → Ca (g)
E. Ca^{2+} (aq) → Ca^{+} (aq) + e^{-}
F. Ca^{2+} (g) → Ca^{+} (g) + e^{-}

2. Which of the options provided below accurately rank the atoms in order of increasing first ionization energies (i.e. the lowest first ionization energy given first)?

$\boxed{D}$

A. Br < Se < As
B. As < Se < Br
C. Br < As < Se
D. Se < As < Br
E. As < Br < Se

Group 16 exception – Se lower than As

3. Which of the following correctly ranks the atoms in terms of increasing ionization energy?

$\boxed{B}$

A. Al < Si < N < O
B. Al < Si < O < N
C. Si < Al < O < N
D. Al < O < Si < N
E. Si < Al < N < O
F. N < O < Si < Al
G. O < N < Si < Al
H. O < N < Al < Si

Group 16 exception – O lower than N

4. Consider a hypothetical element with the following ionization energies given below. Based on this information, what is the most probable, common oxidation state of the hypothetical element? Answer with an integer and sign (e.g. +4, -2).

IE_1 = 123 kJ/mol
IE_2 = 344 kJ/mol
IE_3 = 566 kJ/mol
IE_4 = 12,300 kJ/mol

3 valence e^-s

core e^-

$\boxed{+3}$

5. Why does boron have a lower ionization energy than beryllium?

$\boxed{E}$

A. Beryllium has a full 2s subshell, making it very unstable
B. Boron has an electron in an additional sublevel, making the electron harder to remove
C. The 2s sublevel is higher in energy than the 2p sublevel
D. Boron is to the left of beryllium; ionization energy decreases across a period
E. Boron has a single 2p electron, which is easier to remove than one of beryllium's electrons in a full 2s subshell
F. The ionization energy of boron is expected to have a higher ionization energy, not lower

6. Complete the following statements with (A) larger than, (B) smaller than, or (C) equal to. You may need to use an option more than once. Write the corresponding letters in the boxes below.

I. The *first* ionization energy of sodium is $\boxed{B}$ the *first* ionization energy of magnesium.

→ periodic trends

II. The *second* ionization energy of sodium is $\boxed{A}$ the *second* ionization energy of magnesium.

→ losing a core e^-

7. Which of the options provided below represents the first electron attachment enthalpy of sulfur?

$\boxed{E}$

A. $S^+ (g) + e^- \rightarrow S (g)$
B. $S (g) + 2e^- \rightarrow S^{2-} (g)$
C. $S (g) \rightarrow S^- (g) + e^-$
D. $S (g) \rightarrow S^+ (g) + e^-$
E. $S (g) + e^- \rightarrow S^- (g)$

8. Which of the following elements given below would have the **most positive (i.e. least favorable)** electron attachment enthalpy?

$\boxed{B}$

A. K
B. Ca
C. Ga
D. Ge

Group 2 exception

9. Which of the following correctly ranks the atoms in order of increasing (least to most favorable) electron attachment enthalpy?

$\boxed{A}$

A. Mg < P < Si < Cl
B. Mg < Si < P < Cl
C. Cl < Si < P < Mg
D. Cl < P < Si < Mg

Group 15 exception

10. Which of the following is **false** about periodic trends? Select any that apply and answer using capital letters with no spaces (e.g. ABCDE).

BCF

A. Ionization energy increases with increasing effective nuclear charge
B. Ionization energy increases with increasing atomic radii
C. Electronegativity decreases with increasing effective nuclear charge
D. Electronegativity decreases with increasing atomic radii
E. The electron attachment enthalpy magnitude increases (becomes more favorable) with increasing effective nuclear charge
F. Electron attachment enthalpy becomes less favorable with increasing effective nuclear charge

11. Which of the following statements are **false**? Select any that apply and answer using capital letters with no spaces (e.g. ABCDE).

AD

A. The addition of an electron to nitrogen is more favorable than adding an electron to carbon because nitrogen will be closer to a stable noble gas configuration
B. The electron attachment enthalpy of arsenic is less favorable than for germanium because of the subsequent electron repulsions in adding an electron to arsenic's half-filled 4p subshell
C. Chlorine's more negative (i.e. more favorable) electron attachment enthalpy (-347.8 kJ/mol) compared to fluorine's (-327.8 kJ/mol) can be explained by chlorine's larger atomic radius leading to less electron repulsions
D. While effective nuclear charge (Z_{eff}) increases across a row, this trend does not correlate to electron attachment enthalpy because electron-electron repulsions are the only determining factor in the energy released upon electron addition

12. When considering general periodic trends, elements with a ________ ionization energy are more likely to form cations, while elements with a _______ electron attachment enthalpy are more likely to form anions.

B

A. more positive; less negative
B. less positive; more negative
C. more positive; more negative
D. less positive; less negative
E. None of the above are true

13. It is observed that effective nuclear charge (Z_{eff}) increases across a row. Which of the following options best explain this trend? Select any that apply and answer using capital letters with no spaces (e.g. ABCDE).

AC

A. Electrons don't effectively shield other electrons in the same principle energy level
B. p electrons more effectively shield s electrons
C. The atomic number increases across a row
D. Larger atoms have a greater effective nuclear charge
E. Electron attachment becomes more negative across a row

14. What is **false** about effective nuclear charge (Z_{eff}) and shielding?

D

A. Z_{eff} is the charge electrons actually feel based on atomic number and shielding
B. Increasing effective nuclear charge across a row causes atoms to be smaller
C. It takes more energy to remove an electron (ionization energy is less favorable) across a row because of increased effective nuclear charge
D. In a phosphorus atom, the 3s electrons experience less shielding than the 2p electrons
E. All of the above are true

15. Which of the following atoms has the **smallest** atomic radius?

D

A. Pb
B. Sb
C. Se
D. Cl
E. Based on periodic trends, especially Z_{eff}, it is not possible to conclude which atom has the smallest atomic radius

16. Which of the following correctly ranks the atoms in terms of *increasing* atomic radius?

E

A. Ge < Ne < Br < Kr
B. Ge < Ne < Kr < Br
C. Ne < Ge < Br < Kr
D. Kr < Br < Ge < Ne
E. Ne < Kr < Br < Ge
F. Kr < Br < Ne < Ge

17. Which of the following has the **smallest** atomic radius?

C

A. N
B. N^{3+}
C. N^{5+}
D. N^{5-}
E. N^{3-}
F. They would all have the same atomic radii

18. Consider the following ions and atom below. Which of the following statements is/are **true**? Select any that apply and answer using capital letters with no spaces (e.g. ABCDE).

I^-, Cs^+, Xe

E

54

A. All of the atoms and ions have ~~53~~ electrons
B. All of the atoms and ions are isoelectronic and have the same size
C. The I^- ion has the smallest atomic size (largest)
D. The Xe atom has the highest number of protons because it is neutrally charged
E. None of the above are true

19. Choose the option below that correctly arranges the following ions and atoms in order from largest to smallest size (ionic radii).

Ca^{2+}, Cl^- , Ar, K^+, and S^{2-}

(isoelectronic series)

D

A. Ar > Ca^{2+} > K^+ > Cl^- > S^{2-}
B. K^+ > Ca^{2+} > Ar > Cl^- > S^{2-}
C. Cl^- > S^{2-} > Ar > K^+ > Ca^{2+}
D. S^{2-} > Cl^- > Ar > K^+ > Ca^{2+}
E. S^{2-} > Cl^- > K^+ > Ca^{2+} > Ar
F. All of the ions have the same size because they are all isoelectronic

20. Which of the following atoms would have the *greatest* polarizability?

A

A. Magnesium → larger atomic radii
B. Aluminum
C. Silicon
D. Phosphorus
E. Sulfur

21. Based on periodic trends **solely**, rank the following elements in order of increasing electronegativity (i.e. the least electronegative element given first).

C

A. Si < In < Cl < P
B. Si < P < In < Cl
C. In < Si < P < Cl
D. In < P < Cl < Si
E. P < Si < In < Cl
F. P < In < Si < Cl

22. Based on periodic trends **solely**, would you expect nitrogen or sulfur to be more electronegative?

D

A. Nitrogen would be more electronegative because it is farther up in the periodic table
B. Sulfur would be more electronegative because it is farther right in the periodic table
C. Sulfur would be more electronegative based on its placement in the periodic table and because it contains more protons than nitrogen
D. Based on periodic trends, it is inconclusive whether nitrogen or sulfur is more electronegative

23. Which of the following bonds would best be considered non-polar covalent?

C

A. C—F
B. O—H
C. C—C
D. C—N

24. Which of the following has the bonds correctly arranged in order of **decreasing** polarity (i.e. most polar bond written first)?

$\boxed{B}$

A. Sr-I > P-S > Si-Cl > F-F
B. Sr-I > Si-Cl > P-S > F-F
C. F-F > Si-Cl > P-S > Sr-I
D. F-F > P-S > Si-Cl > Sr-I
E. P-S > Sr-I > Si-Cl > F-F
F. Si-Cl > Sr-I > P-S > F-F

25. Which of the following has the partial charges correctly assigned?

$\boxed{A}$

A. $Cl^{\delta-}$ — $C^{\delta+}$
B. $Cl^{\delta-}$ — $F^{\delta+}$
C. $C^{\delta-}$ — $F^{\delta+}$
D. $N^{\delta-}$ — $O^{\delta+}$

26. Which of the following reactions represents the lattice energy in the formation of solid barium selenide?

$\boxed{A}$

A. Ba^{2+} (g) + Se^{2-} (g) → BaSe (s)
B. Ba^{2+} (aq) + Se^{2-} (aq) → BaSe (s)
C. Ba^{2+} (g) + 2 Se^{-} (g) → $BaSe_2$ (s)
D. Ba^{2+} (aq) + 2 Se^{-} (aq) → $BaSe_2$ (s)

27. Rank the compounds given below in order of increasing lattice energy released upon formation (i.e. the least negative, or weakest, associated lattice energy given first).

+1 +1 +2 +2
NaI, NaBr, CaO, SrO
-1 -1 -2 -2

greater magnitude of charge, greater lattice energy

$\boxed{D}$

A. CaO < SrO < NaI < NaBr
B. CaO < SrO < NaBr < NaI
C. NaI < NaBr < CaO < SrO
D. NaI < NaBr < SrO < CaO
E. CaO < NaI < SrO < NaBr
F. SrO < NaBr < CaO < NaI

if charges equal, then look at ionic radii
↳ smaller ionic radii, greater lattice energy

28. Which of the following ionic compounds would be expected to have the greatest (most favorable) lattice energy?

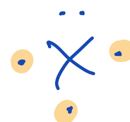
$\boxed{C}$

A. Calcium sulfide
B. Sodium sulfide
C. Magnesium nitride → largest charges; Mg_3N_2 has smaller ionic radii
D. Magnesium phosphide
E. Sodium nitride

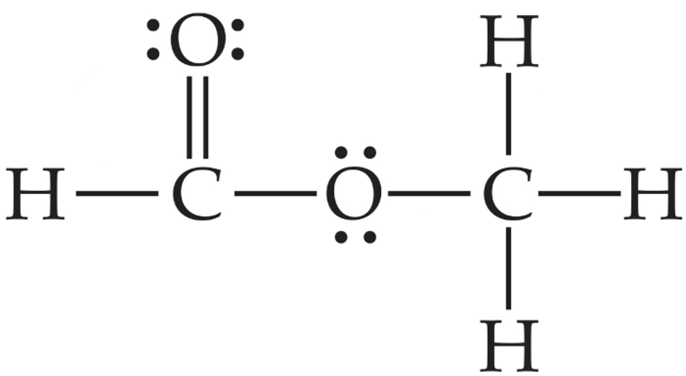
$\overset{+2}{Ca}\overset{-2}{S}$ $\overset{+1}{Na_2}\overset{-2}{S}$ $\overset{+2}{Mg_3}\overset{-3}{N_2}$ $\overset{+2}{Mg_3}\overset{-3}{P_2}$ $\overset{+1}{Na_3}\overset{-3}{N}$

29. The Lewis symbol of an unknown atom has 2 paired electrons and 3 unpaired electrons. How many bonds is this atom likely to form in a covalent compound? Answer by using an integer (e.g. 0, 1, etc.).

3



30. Consider the Lewis structure given below. How many bonding pairs are present? Single bonds? Double bonds? Non-bonding (lone) pairs?



I. Non-bonding (lone) pairs: 4

II. Bonding pairs: 8

III. Single bonds: 6

III. Double bonds: 1

Additional Practice Questions:

1. Which of the options provided below accurately rank the atoms in order of increasing first ionization energies (i.e. the lowest first ionization energy given first)?

$\boxed{B}$

Group 2 exception

Mg > Al

A. Na < Mg < Al
B. Na < Al < Mg
C. Mg < Al < Na
D. Al < Na < Mg
E. Al < Mg < Na

2. Which of the following would have the *highest* *second* ionization energy?

$\boxed{D}$

A. Kr
B. Ge
C. Se
D. K → loses core e^-
E. Ca

3. Which elements are correctly compared based on their first ionization energies?

$\boxed{E}$

A. Na > Mg
B. S > P
C. Br > Cl
D. Cs > Li
E. N > O

Group 16 exception

4. Which of the following elements has the **largest** atomic radius?

$\boxed{A}$

A. As
B. O
C. Br
D. S

5. Which of the following elements has the **smallest** atomic radius?

$\boxed{A}$

A. Mg
B. Ca
C. Sr
D. Ba
E. All of these elements have the same atomic radii because they are all group 2 elements

6. Which of the following species has the **smallest** ionic radius?

$\boxed{A}$

A. Al^{3+}
B. Na^{+}
C. K^{+}
D. S^{2-}
E. Cl^{-}

7. Which of the following ions would be expected to have the **smallest** ionic radius?

$\boxed{D}$

- Periodic trends
- cations smaller than anions

$\rightarrow Sr^{2+} < Ba^{2+}$

A. Se^{2-}
B. Br^{-}
C. Rb^{+}
D. Sr^{2+}
E. Ba^{2+}

8. Barium is predicted to have a higher polarizability than strontium, calcium, or magnesium. Which of the following options below best supports this?

$\boxed{C}$

A. Barium's larger atomic mass increases its polarizability
B. Barium's smaller atomic mass increases its polarizability
C. Barium's larger atomic radius increases its polarizability
D. Barium's smaller atomic radius increases its polarizability

9. Effective nuclear charge increases across a period because…

$\boxed{B}$

A. The number of electrons increases across a period
B. The number of protons increases across a period, but the number of core electrons remains the same
C. The size of atoms decreases across a period
D. The number of possible covalent bonds an atom may make increases across a period due to increasing non-metal character

10. Consider a number of hypothetical ions: M^+, R^{5+}, X^{2-}, and Z^{4-} that can form various ionic compounds. Which of the following compounds would have the largest lattice energy released upon formation (i.e. the most negative, or strongest, associated lattice energy)?

$\boxed{D}$

A. M_2X
B. M_4Z
C. R_2X_5
D. R_4Z_5 → greater magnitude of charge, greater lattice energy

11. Rank the compounds given below in order of increasing lattice energy released upon formation (i.e. the least negative, or weakest, associated lattice energy given first).

$\overset{+2\ -2}{SrSe}$, $\underset{+2\ -2}{MgO}$, $\overset{+2\ -2}{CaS}$

$\boxed{B}$

magnitude of charges equal, now we look at ionic radii

↑ ionic radii = ↓ lattice energy

A. SrSe < MgO < CaS
B. SrSe < CaS < MgO
C. MgO < CaS < SrSe
D. MgO < SrSe < CaS
E. CaS < MgO < SrSe
F. CaS < SrSe < MgO

12. Which of the following has the elements ranked in order of *increasing* electronegativity (the least electronegative to the most electronegative element)?

$\boxed{A}$

A. C < N < O < F
B. C < O < N < F
C. N < C < O < F
D. F < O < N < C
E. F < N < O < C

13. An electrostatic potential map is collected for an unknown bond. There is very little to no difference in color in the map. Which of the following bonds likely correspond to this map?

$\boxed{C}$

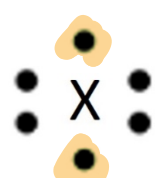
A. N—Br bond
B. As—F bond
C. C—H bond very little polarity
D. More than one of the options above
E. None of the above

14. The phosphorus-chlorine bonds in phosphorus trichloride are…

$\boxed{B}$

A. Ionic
B. Polar covalent
C. Covalent
D. Cannot tell based on electronegativity values

15. How many bonds would the following atom make, according to its Lewis Symbol? Answer by using an integer (e.g. 0, 1, etc.).



2 unpaired e⁻s

$\boxed{2}$

Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc [97]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
37 Cs 132.91	56 Ba 137.33		72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]		104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [277]	110 Ds [281]	111 Rg [282]	112 Cn [285]	113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac [227]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]

Formula Sheet

Length
1 kilometer = 0.62137 mile
1 inch = 2.54 centimeters (exactly)
1 Ångstrom = 1 × 10^{-10} meter

Energy
1 joule = 1 $kg \cdot m^2/s^2$
1 calorie = 4.184 joules
1 Calorie = 1 kilocalorie = 1000 calories
1 L·atm = 101.325 joules

Pressure
1 pascal = 1 N/m^3 = 1 $kg/m \cdot s^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in^2
1 bar = 1 × 10^5 Pa (exactly)

Temperature
0 K = −273.15°C
K = °C + 273.15
°C = (5/9)(°F − 32)

Mass
1 kg = 2.205 lbs

Volume
1 mL = 1 cm^3 = 1 cc

Constants
c = 2.998 × 10^8 m/sec
h = 6.626 × 10^{-34} J·sec
R = 0.08206 L·atm/mol·K = 8.314 J/mol·K
Specific heat of water = 4.184 J/g·K
Mass of an electron: 9.109 x 10^{-31} kg
Mass of a proton: 1.673 x 10^{-27} kg
RH = 2.18 x 10^{-18} J
Specific heat of water = 4.184 J/g·K
Avogadro's number: 6.022 × 10^{23}
F = 96485 J/(V·mol e^-)
K_w = 1.0 × 10^{-14} at 25 °C
k_b = 1.381 × 10^{-23} J/K

Equations
$(P + a(n^2/V^2)) \cdot (V - nb) = nRT$
molar mass (M) = mRT/PV
density (d) = MP/RT

$$KE = \frac{3}{2}RT$$

$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{Rate\ of\ effusion\ A}{Rate\ of\ effusion\ B} = \sqrt{\frac{MW_B}{MW_A}}$$

$$\Delta E = -2.18 \times 10^{-18} J(\frac{1}{n_f^2} - \frac{1}{n_i^2})$$

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$C_g = kP_g$
$P_{solution} = P_{solvent}\, X_{solvent}$
$P_{solution} = \sum P_j = \sum P_j X_j$
$\pi = MRTi$

Thermodynamic and Electrochemistry

$S = k_b \times \ln(W)$
$\Delta S = q_{rev}/T$
$\Delta G = \Delta G° + RT \cdot \ln Q$
R=8.314 J/mol.K
$\Delta G° = -RT \cdot \ln K$
$\Delta G = -nFE_{cell}$
$E°_{cell} = RT/nF \ln K$
$E°_{cell} = (0.0257/n) \ln K = (0.0592/n) \log K$
$E_{cell} = E°_{cell} - (RT/nF) \ln Q$
$E_{cell} = E°_{cell} - (0.0257/n) \ln Q$
Electrolysis: Q (total charge) = I × t = n × F

Integrated Rate Laws & half-life

$$\ln\frac{[A]}{[A]_0} = -kt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

$$[A] = -kt + [A]_0$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$ln\frac{k_2}{k_1} = -\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})$$

Equilibrium and Acid / Base

$K_P = K_c \times (RT)^{\Delta n}$

$$ln\frac{K_2}{K_1} = \frac{\Delta H^{\circ}_{rxn}}{R}(\frac{1}{T_1} - \frac{1}{T_2})$$